Topic 3: Chemical Reactions

Appendices

Appendix 3.1: Calculating Average Atomic Mass

1. Identify the numbers of protons, neutrons, and electrons in a neutral atom of each of the following:

   a) ${}^{235}_{92}U$

   b) ${}^{226}_{88}Ra$

2. Complete the following table to calculate the average atomic mass of chlorine (Cl).

Isotope	**Mass of Each Atom**	**Number of Atoms**	**Total Mass**
Cl-35	34.969 μ	758	
Cl-37	36.966 μ	242	
Totals		1000	
Average			

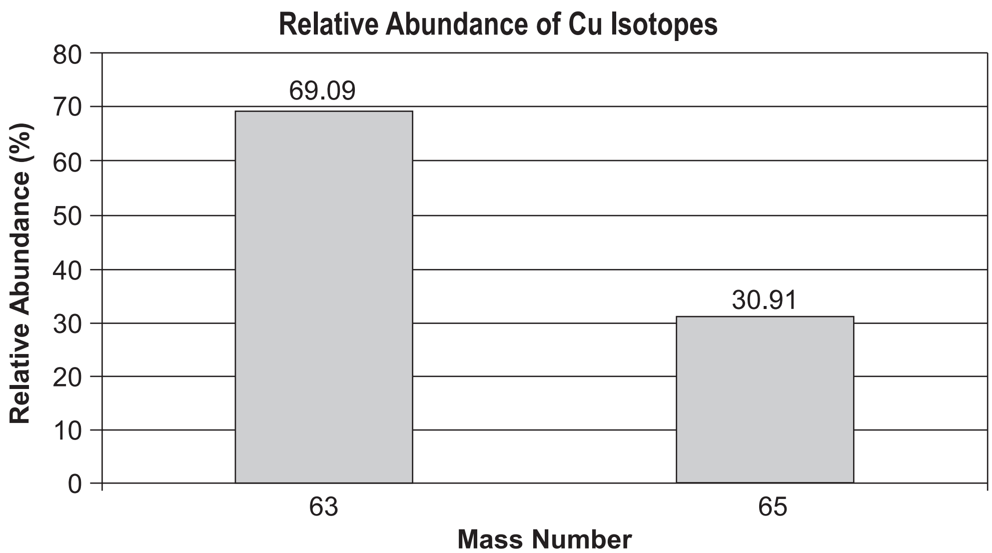
3. Complete the following table to calculate the average atomic mass of each element.

Element	**Symbol**	**Mass Number**	**Mass (μ)**	**Relative Abundance (%)**	**Average Atomic Mass (μ)**
Carbon (C)	C-12	12	12 (exactly)	98.98	
	C-13	13	13.003	1.11	
Silicon (S)	Si-28	28	27.977	92.21	
	Si-29	29	28.976	4.70	
	Si-30	30	29.974	3.09	

Appendix 3.1: Calculating Average Atomic Mass *(continued)*

4. Define the term *isotope*. Explain how an element's atomic mass is related to the abundances of its different isotopes.

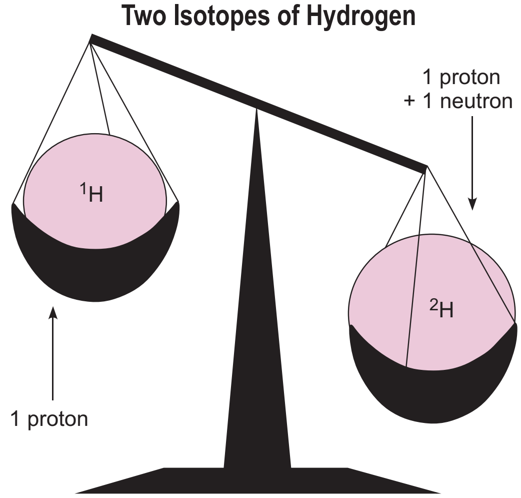
5. Using the graph below, calculate the average atomic mass of copper (Cu).



Appendix 3.2: Don't Be an Isodope: Get the Facts on Isotopes

Any single atomic element has a fixed number of protons. However, nearly all elements are capable of possessing more than one fixed number of neutrons. For example, hydrogen is defined as an atom with only one proton. Hydrogen commonly has zero neutrons, giving this type of atom an atomic mass of 1. This means that 6.0225×10^{23} (one mole) of these common hydrogen atoms weighs 1 gram. However, every once in a while a hydrogen atom will also have a neutron, which has basically the same mass as a proton. These heavier hydrogen atoms are referred to as ^{2}H, or deuterium, and have an atomic mass of 2. Even more rarely, a hydrogen atom will have two neutrons. These atoms have an atomic mass of 3 and are referred to as ^{3}H, or tritium. ^{1}H, ^{2}H, and ^{3}H are all *isotopes of hydrogen*.

Two Isotopes of Hydrogen



Some isotopes are classified as stable isotopes and others are classified as unstable, or radioactive, isotopes. Stable isotopes maintain constant concentrations on Earth over time. Unstable isotopes are atoms that disintegrate at predictable and measurable rates to form other isotopes by emitting a nuclear electron or a helium nucleus and radiation. These isotopes continue to decay until they reach stability. As a rule, the heavier an isotope is than the most common isotope of a particular element, the more unstable it is and the faster it will decay.

Because the rates of radioactive decay are measurable, unstable isotopes are useful tools in determining age. For example, nuclear bombs put large and detectable amounts of certain radioactive isotopes into the atmosphere. After the near-elimination of nuclear bomb testing due to the Limited Test Ban Treaty in 1963, the carbon-14 concentration in the atmosphere began decreasing immediately. Therefore, someone born in 1965 has a significantly lower concentration of carbon-14 in the bones than a person who was born in 1960. Anybody born after

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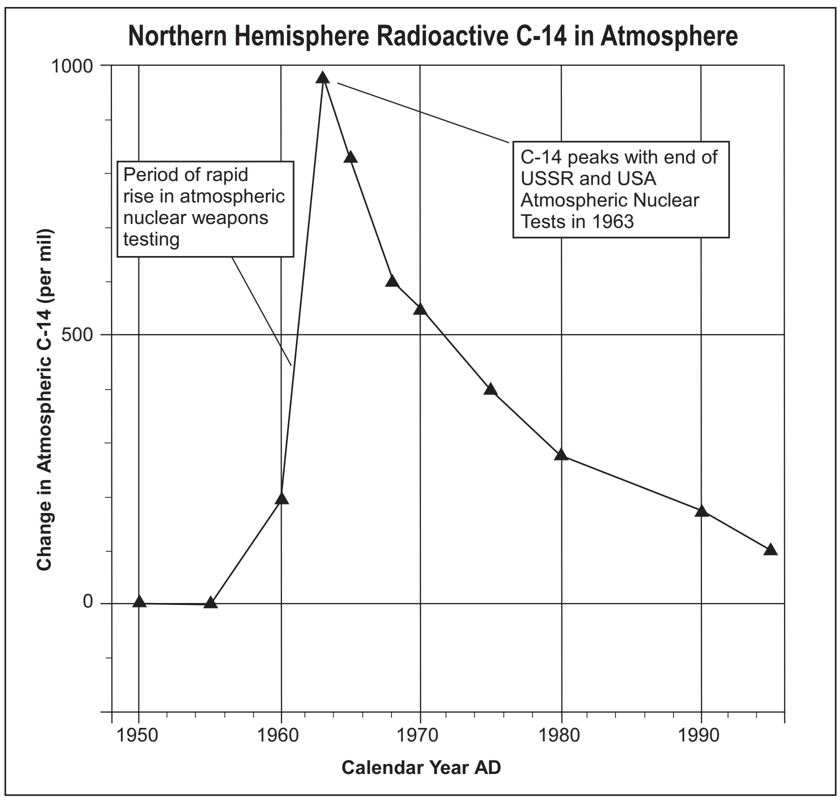
Appendix 3.2: Don't Be an Isodope: Get the Facts on Isotopes *(continued)*

1965 or so possesses a significantly lower concentration of carbon-14 than someone born before atmospheric nuclear testing came to a close. Thus, we can tell how old many living organisms are based on the recent history of carbon-14 in the atmosphere from nuclear tests! Because carbon-14 decays slowly (decreases to half of its original concentration every 5,370 years), we can also tell how long ago much older organisms died, based upon what we know the pre-industrial carbon-14 concentration in the atmosphere was and how much radioactive decay of this isotope has occurred since the organism's death. This is a very useful tool in dating petrified wood, bones from ancient civilizations, and shells in ocean-floor sediments. For very old specimens such as dinosaur bones, more stable radioactive isotopes must be used because of their slower decay rates.



Teacher Notes

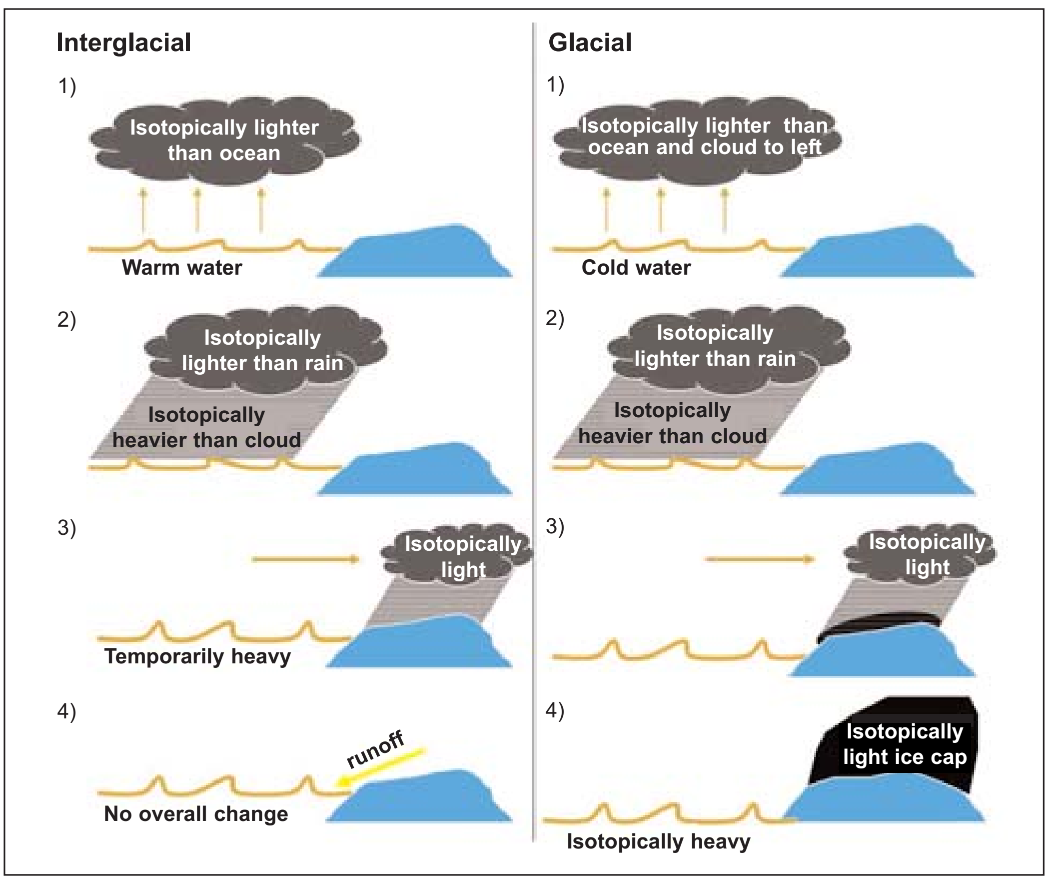
During the Cold War, the USA and USSR conducted numerous atmospheric atomic weapons tests. Those events are recorded in the history of radioactive C-14 that was contained in Earth's atmosphere.

The year 1963 was the end of atmospheric testing, as the Nuclear Test Ban Treaty was signed by both competing nations. Even after 40 years, some residual radioactive carbon is being scrubbed out of the atmosphere in precipitation.

Appendix 3.2: Don't Be an Isodope: Get the Facts on Isotopes *(continued)*

While radioactive isotopes are very cool, you may be looking for a little more stability in your life, and are therefore interested in *stable isotopes*. Because stable isotopes don't decay, they remain in the environment at a constant concentration. However, the distribution of stable isotopes throughout the environment constantly changes as a result of changing environmental preferences. For example, much like salinity, the concentrations of heavy hydrogen and oxygen isotopes in the ocean increase significantly during glacial periods because cold air is not as good at absorbing and holding onto heavy water molecules as warm air is. This means that water molecules made of light hydrogen and oxygen isotopes evaporate from the ocean more easily than heavy molecules.

During glacial times, much more water is trapped on land as ice than during interglacial times, redistributing the light water molecules that preferentially evaporated from the ocean onto land. Trapping water on land decreases the volume of the ocean and therefore increases the concentration of heavy hydrogen and oxygen isotopes in the ocean. Because the concentration of these heavy isotopes in the ocean is a function of temperature, the environment has been keeping a record of sea surface temperature for millions of years by the constant piling up of dead organic matter on the ocean floor (sedimentation). For example, Foraminifera living



Appendix 3.2: Don't Be an Isodope: Get the Facts on Isotopes *(continued)*

at the ocean's surface make their calcium carbonate, $CaCO_3$, shells out of the chemistry they pull out of the ocean. If the ocean is changing its isotopic composition over time, and if these shells are constantly falling to the ocean floor as new shells are created, then measuring concentrations of heavy oxygen isotopes in shells going down in depth from the ocean floor is like travelling back in time and measuring sea surface temperature over millions of years.

There are many other processes that influence the preferential distribution of other heavy and light isotopes as well. For example, the concentration of nitrogen-15 (^{15}N) in an animal indicates whether or not that animal is starving. The combination of nitrogen-15 and carbon-13 in a human indicates how much meat that person eats, and comparing these measurements around the world can tell us much about dietary differences from country to country. Changes in carbon-13 concentration throughout the hoof of a herbivore, such as a deer, can indicate drought because of changes in chemistry of the plants that the deer eats, and changes in carbon-13 of bones from 10 million to 3 million years ago indicate a significant change from woody plants (C3) to grasses (C4) about 5 million years ago.

Appendix 3.3: Isotopes Used in Medicine and Climatology

Isotope	Application	Use	Radiation	Half-Life
Sodium-24	Medical radioactive tracer	• to detect blood flow constrictions and obstructions	Beta emitter	14.8 h
Iodine-131	Medical radioactive tracer	• to test the activity of the thyroid gland	Beta emitter	8 d
Technetium-99	Medical radioactive tracer	• to image organs such as heart, liver, and lungs • to do 3-phase bone scans	Gamma emitter	6 h
Cobalt-48	Medical radioactive tracer	• to determine intake of vitamin B12 that contains non-radioactive cobalt		71.3 d
Iron-59		• to determine the rate of red blood cell formation (they contain iron)		45.6 d
Chromium-51		• to determine blood volume and lifespan of red blood cells		27.8 d
Hydrogen-3 Tritium		• to determine volume of water in person's body • to determine the use of (labelled) vitamin D in body • to conduct cellular chemistry research		12.3 y
Strontium-85		• to do bone scans		64 d
Gold-198		• to do liver scans		2.7 d
Phosphorus-32		• to determine eye disorders, liver tumours		14.3 d

Radioisotope tracers are used for diagnosis in medicine. One advantage of using radioactive isotopes is that their particle emissions are straightforward to detect. Photographic imaging techniques or the use of devices known as scintillometers (counters) can detect their presence even in small amounts.

Appendix 3.4: The Importance and Application of Isotopes

During today's class, you will find information about the following isotopes:

Aluminum-26
Americium-241
Bismuth-213
Calcium-42
Calcium-44
Californium-252
Carbon-13
Carbon-14
Cesium-137
Cobalt-60
Deuterium (Hydrogen-2)
Fluorine-18
Gold-198
Iodine-125
Iodine-131
Iridium-192
Krypton-85
Lead-206
Nickel-62
Nitrogen-14
Nitrogen-15
Oxygen-15
Oxygen-16
Oxygen-18
Phosphorus-32
Plutonium-238
Polonium-210
Promethium-147
Rhenium-188
Rubidium-82
Silicon-32
Sodium-24
Strontium-90
Technetium-99m
Thorium-229
Tritium (Hydrogen-3)
Uranium-238
Vanadium-5

Isotopes are made in particle accelerators or nuclear reactors. As they decay over time (some quite quickly), parent isotopes are manufactured so that the desired isotope is made through decay during storage or transport.

The isotope products can be used in

1. medical diagnosis
2. medical treatment
3. agriculture
4. science
5. industry and business
6. consumer products and services

In the computer lab, use WebQuests to research the importance and applications of isotopes. The following two websites will be especially useful:

- Environment Canada. "Isotopes Link Birds to Breeding and Moulting Areas." *Science and the Environment Bulletin*. (May/June 2001): <www.ec.gc.ca/science/sandemay01/article7_e.html>
- Nuclear Energy Institute (NEI): <www.nei.org/index.asp?catnum=1&catid=1>

Based on your research on the importance and applications of the isotopes listed above, complete the left column of the charts on the following pages. Please note that some isotopes may be used more than once.

Appendix 3.4: The Importance and Application of Isotopes *(continued)*

1. **Medical Diagnosis**

   Isotopes can identify abnormal bodily processes, because some natural elements tend to concentrate in certain internal organs. After a patient has been injected with an isotope, a special camera can take pictures of the internal workings of the patient's body. Nuclear medicine was developed in the 1950s.

Isotope	Use
• •	PET (positron emission tomography) scans use these two isotopes to measure energy metabolism in the brain.
•	This isotope is used to image the thyroid, heart, lungs, and liver. It also measures blood iodine levels.
•	Injected into the bloodstream as a salt solution, this isotope can be monitored to trace blood flow in order to detect constrictions and obstructions in the circulatory system (i.e. prevent a heart attack).
•	This isotope is used in cardiac (heart) imaging because its chemical reactivity is similar to that of potassium (which is used in muscles, such as the heart). Once the isotope reaches the heart, a PET image can be made. The isotope decays away within a day.
•	This isotope helps diagnose bone infection in children, as well as brain tumours.
•	This isotope is used to detect *helicobacter pylori* (which can damage the stomach lining and lead to ulcer formation) in the human stomach. If the bacteria are there, they will ingest the isotope and produce $^{13}CO_2$ that will be exhaled at a higher concentration than normal.

Appendix 3.4: The Importance and Application of Isotopes *(continued)*

2. Medical Treatment

The direct internal or external application of isotopes is superior to chemotherapy because the isotopes are specific to the tumour or cancer involved, and cause less damage to healthy tissue.

Isotope	Use
•	This isotope fights lung cancer and leukemia.
•	Implants of this isotope, or mixtures of Strontium-90 and Ytterbium-90, are used to destroy pituitary and breast cancer tumours.
•	Gamma rays from this isotope are used to destroy brain tumours.
•	This isotope is used as a therapeutic seed implant for prostate cancer.
•	Used to treat Graves' disease (a thyroid condition), this isotope concentrates in the thyroid and destroys the diseased portion. It is also used to treat thyroid cancer.
•	This isotope treats cervical cancer.
•	This isotope treats bone pain.

3. Agriculture

Isotope	Use
• • •	These three isotopes help farmers determine the amount of nutrients and water that plants take from the soil. This information gives farmers a better indication of how much fertilizer and water they should use to avoid excess. It helps them to protect the environment from runoff, to preserve the water supply, and to save money.
Other Uses of Isotopes in Agriculture • To breed disease-resistant livestock, scientists use radioactive material to pinpoint where illnesses strike animals. • Scientists use isotopes to control pests. Small levels of radiation sterilize disease-laden insects and pests, replacing the undesirable use of pesticides. • Isotopes are used to develop new strains of virus-resistant plants.	

Appendix 3.4: The Importance and Application of Isotopes *(continued)*

4. Science

Isotope	Use
•	This isotope links birds and butterflies to their breeding and moulting areas. It enters feather and wing tissue (where the feathers are grown or the wing tissue is developed) as a result of the water that the organisms drink and the plants they eat. Concentrations of this isotope are higher in organisms from southeast regions of North America, and lower in the northwest regions. When water evaporates at the equator, the water with this isotope evaporates faster because it is heavier.
•	An isotope of this element was used to find a possible explanation of why some mallard ducklings born and raised in agricultural areas cannot fly.
• •	Criminal investigators use the ratio of these two isotopes in hair, tissues, and fluids to determine where people have lived.
•	This isotope measures the silicon uptake by ocean phytoplankton in order to monitor possible trends in global warming.
• •	One of these isotopes is given intravenously to a patient, and the other is ingested. The isotope given intravenously passes quickly through the body, while the ingested one must be metabolized. The ratio between the two isotopes indicates how long calcium is retained in the body before it is eliminated.
• •	These two isotopes are used to date objects such as rocks. They can be used to date objects up to 4.5×10^9 years old.
•	This isotope is used to study the effects of acid rain. Its path to waterways can be modelled using the isotope as a tracer. It is also being used to see whether there is a relationship between aluminum and Alzheimer's disease.
• •	The ratio between these two isotopes is used to indicate global temperatures. Rain near the tropical ocean has more of the heavier isotope, while precipitation near the poles has more of the lighter isotope. During ice ages, more of the lighter isotope is locked into polar ice, and rain everywhere on the planet has a higher concentration of the heavier isotope.
•	Large quantities of this isotope are required for nuclear weapons and nuclear reactions. It is also burned in brown dwarfs (failed stars).

Appendix 3.4: The Importance and Application of Isotopes *(continued)*

5. Industry and Business

Isotope	Use
•	This isotope tracks leakage in buried sections of pipe.
•	This isotope gauges the moisture content of the soil at road- and building-construction sites.
• •	These two isotopes determine the efficiency of filtration systems.
•	This isotope is used to identify when mixing is complete in molten liquids. This ensures maximum strength for alloys.
• • •	These three isotopes are used to recognize welding defects in pipelines, boilers, and aircraft parts. The radiation is passed through the object, striking photographic film on the other side. The greater the number of cracks, breaks, or flaws in the object, the more radiation is recorded on the film.
•	This isotope is used in industrial devices to eliminate static electricity. It is also used to treat bottles before they are filled and to reduce static charge in the production of photographic film.
•	Museums rely on this radioactive material to verify the authenticity of paintings and art objects, while archeologists use it to determine the age of bones.
• •	These two isotopes can be used to detect explosives.
•	This isotope provides National Aeronautics and Space Administration (NASA) spacecraft with power.
•	This isotope is a fuel for nuclear power plants.
•	This isotope monitors paper thickness. If too few particles of this isotope pass through to the detector, the paper is too thick. If too many particles pass through, the paper is too thin.

(continued)

Appendix 3.4: The Importance and Application of Isotopes *(continued)*

5. Industry and Business (continued)

<table>
<tr><th>Isotope</th><th>Use</th></tr>
<tr><td>•
•</td><td>Food is exposed to either of these two isotopes for 15 to 45 minutes in order to kill bacteria and parasites that can cause disease. (Note: This does not make the food radioactive.) This exposure reduces spoilage and increases the shelf life of foods in areas without refrigeration. While the process reduces levels of vitamin A, vitamin C, vitamin E, and thiamine, the levels are not reduced as much as when foods are cooked, canned, or frozen.

The same isotopes are used to sterilize cosmetics, hair products, medical devices, bandages, condoms, tampons, and contact-lens solutions.</td></tr>
<tr><td>•</td><td>This isotope is used to date vintage wines.</td></tr>
<tr><td colspan="2">Other Uses of Isotopes in Industry and Business
• Some isotopes can determine the flow rates of piping systems.
• The automobile industry uses isotopes to measure the rate of engine wear and the quality of steel used to make cars.
• Manufacturers use isotopes to monitor the corrosion of their processing equipment.
• Isotopes are used to identify the densities of various materials.
• The proper thickness of aluminum and tin cans can be maintained with the help of isotopes.
• Some isotopes are used to identify the geological structure of a site. This would show the likelihood of finding oil, natural gas, or minerals.
• Isotopes used in photocopiers help eliminate static and prevent papers from sticking together.</td></tr>
</table>

Appendix 3.4: The Importance and Application of Isotopes *(continued)*

6. Consumer Products and Services

Isotope	Use
•	Smoke detectors rely on this radioactive source to sound an alarm when sensing smoke from a fire. This isotope emits particles that ionize the air particles around the alarm. These charged air molecules conduct electricity so that current flows through the device. Smoke will interrupt the flow of electricity and set off the alarm.
•	This gas can extend the life of a computer chip eight times longer than a chip created with hydrogen gas.
•	This isotope is used in indicator lights in washing machines, dryers, stereos, and coffee makers.
• •	These two isotopes act as voltage regulators and current surge protectors.
•	This isotope is used in electric blanket thermostats.
•	This isotope prolongs the life of fluorescent lights.
•	Together with minute amounts of phosphor, this isotope creates the luminescence that is used for emergency lighting, self-luminous aircraft, exit signs, and luminous dials, gauges, watches, and paint. Because no electricity is required, emergency lighting can be provided where sparks can be dangerous or where no wiring exists.
•	This isotope is used in dental fixtures (i.e., crowns and dentures) to provide natural colour and brightness.
•	This isotope powers pacemakers, reducing the risk of repeated surgery.
Other Uses of Isotopes in Consumer Products and Services • Computer disks remember data better when they are treated with isotopes. • Non-stick pans are treated with isotopes so that the coating sticks to the surface better.	

Appendix 3.4: The Importance and Application of Isotopes *(continued)*

Answer Key

1. **Medical Diagnosis**
   - Oxygen-15/Fluorine-18
   - Iodine-131
   - Sodium-24
   - Rubidium-82
   - Technetium-99m
   - Carbon-13

2. **Medical Treatment**
   - Bismuth-213
   - Gold-198
   - Cobalt-60
   - Iodine-125
   - Iodine-131
   - Californium-252
   - Rhenium-188

3. **Agriculture**
   - Phosphorus-32/Nitrogen-15/ Carbon-14

4. **Science**
   - Deuterium (Hydrogen-2)
   - Nitrogen-14
   - Oxygen-18/Deuterium (Hydrogen-2)
   - Silicon-32
   - Calcium-42/Calcium-44
   - Uranium-238/Lead-206
   - Aluminium-26
   - Oxygen-16/Oxygen-18
   - Deuterium (Hydrogen-2)

5. **Industry and Business**
   - Iodine-131
   - Californium-252
   - Deuterium (Hydrogen-2)/Oxygen-18
   - Vanadium-52
   - Iridium-192/Sodium-24/Cobalt-60
   - Polonium-210
   - Carbon-14
   - Nickel-62/Californium-252
   - Plutonium-238
   - Uranium-238
   - Strontium-90
   - Cobalt-60/Cesium-137
   - Tritium (Hydrogen-3)

6. **Consumer Products and Services**
   - Americium-241
   - Deuterium (Hydrogen-2)
   - Krypton-85
   - Nickel-62/Californium-252
   - Promethium-147
   - Thorium-229
   - Tritium (Hydrogen-3)
   - Uranium-238
   - Plutonium-238

Appendix 3.5: Names, Formulas, and Charges of Some Common Ions

Positive Ions		Negative Ions	
Ammonium	NH_4^+	Acetate, Ethanoate	CH_3COO^-
Cadmium	Cd^{2+}	Arsenate	AsO_4^-
Chromium(II)	Cr^{2+}	Bromate	BrO_3^-
Chromium(III)	Cr^{3+}	Carbonate	CO_3^{2-}
Cobalt(II)	Co^{2+}	Hydrogen Carbonate, Bicarbonate	HCO_3^-
Cobalt(III)	Co^{3+}	Chlorate	ClO_3^-
Copper(I)	Cu^+	Chlorite	ClO_2^-
Copper(II)	Cu^{2+}	Chromate	CrO_4^{2-}
Hydrogen, Hydronium	H^+, H_3O^+	Cyanide	CN^-
Iron(II)	Fe^{2+}	Dichromate	$Cr_2O_7^{2-}$
Iron(III)	Fe^{3+}	Hydride	H^-
Lead(II)	Pb^{2+}	Hydroxide	OH^-
Lead(IV)	Pb^{4+}	Hypochlorite	ClO^-
Manganese(II)	Mn^{2+}	Iodate	IO_3^-
Manganese(III)	Mn^{3+}	Nitrate	NO_3^-
Mercury(I)	Hg_2^{2+}	Nitrite	NO_2^-
Mercury(II)	Hg^{2+}	Oxalate	$C_2O_4^{2-}$
Nickel	Ni^{2+}	Oxide	O^{2-}
Scandium	Sc^{3+}	Perchlorate	ClO_4^{1-}
Silver	Ag^+	Permanganate	MnO_4^-
Tin(II)	Sn^{2+}	Phosphate	PO_4^{3-}
Tin(IV)	Sn^{4+}	Phosphite	PO_3^{3-}
Zinc	Zn^{2+}	Monohydrogen Phosphate	HPO_4^{2-}
		Dihydrogen Phosphate	$H_2PO_4^-$
		Sulphate	SO_4^{2-}
		Hydrogen Sulphate, Bisulphate	HSO_4^-
		Sulphite	SO_3^{2-}
		Hydrogen Sulphite, Bisulphite	HSO_3^-
		Hydrogen Sulphide, Bisulphide	HS^-

Appendix 3.6: Ionic Name Game

1. Remove one positive ion (blue piece of paper) and one negative ion (yellow piece of paper) from the container presented by your teacher.
2. List the cation and anion in the appropriate columns of the table below.
3. Use this information to determine the formula and the name of the ionic compound formed when these two ions combine.
4. List the formula and the name in the table.
5. Return the cation and anion to the bowl, mix the papers, and repeat steps 1 through 4 until you have named 12 different ionic compounds.

Positive Ion (Cation)	Negative Ion (Anion)	Formula of the Compound	Name of the Compound

Appendix 3.7: Stoichiometry: The Formula of a Precipitate (Student Experiment)

Problem

What is the formula of cobalt hydroxide?

Materials

- well plate
- ruler
- 2 polyethylene pipettes
- 13 culture tubes (6 x 50 mm)
- toothpick (or sealed capillary tube)

Aqueous Solutions

- cobalt(II) chloride, $CoCl_2$, 0.0159 g/mL
- sodium hydroxide, NaOH, 0.0160 g/mL

Safety Precautions

- Sodium hydroxide is caustic and corrosive. Use caution and wash any spills with plenty of water.
- Cobalt compounds are toxic; do not ingest. Wash hands thoroughly before leaving the laboratory. See Material Safety Data Sheet.

Procedure

1. Place the 13 small culture tubes (test tubes) in the wells of the well plate.
2. Add 24 drops of the cobalt(II) chloride solution to the 1st test tube. Add 22 drops of cobalt(II) chloride solution to the 2nd test tube. Continue to decrease the number of drops by 2 until you have placed 2 drops in the 11th test tube. The 12th test tube remains empty; it acts as a control. Check test tubes 1 through 11 to see whether each successive test tube has decreased by the same amount—it should look like a staircase. If one or more tubes seem to have too little or too much, adjust accordingly by adding or withdrawing some cobalt(II) chloride solution.
3. Repeat this procedure with the sodium hydroxide solution, but start in the reverse order (i.e., add 24 drops of sodium hydroxide solution to the 12th test tube, 22 drops to the 11th tube, and so on). The 1st test tube will have no sodium hydroxide solution added; it also acts as a control. The total volume in every test tube should be the same.
4. Using a toothpick or a sealed capillary tube, stir the solutions with an up-and-down motion. This helps the precipitate settle to the bottom of the test tube.
5. Allow the test tubes to sit undisturbed in the well plate for 10 minutes.
6. Take each test tube out of the well plate in turn, and measure the height of the precipitate in the test tube.

Appendix 3.7: Stoichiometry: The Formula of a Precipitate (Student Experiment) *(continued)*

7. Record your results in a table similar to the one below.

Test Tube	Drops of $CoCl_{2(aq)}$	Drops of $NaOH_{(aq)}$	$\frac{\text{Drops of } CoCl_{2(aq)}}{\text{Drops of } NaOH_{(aq)}}$	Height of Precipitate
#1	24	0	...	0
#2	22	2	11	...
#3	20	4	5	...

8. Transfer the contents from all the test tubes to the appropriate waste container.

Conclusion

1. What can you conclude from this experiment?

2. What is the mass to volume ratio of the precipitate?

Follow-up Questions

1. What is the significance of measuring the volume?

2. How could you relate the volume measured to the mass of the reacting substances?

3. What is the significance of measuring the height? What does the height represent (approximately), or what is the height proportional to?

Appendix 3.7: Stoichiometry: The Formula of a Precipitate (Discussion/Extensions—Teacher Notes)

Safety Precautions

- Lead and cobalt compounds are toxic and accumulate in body tissues. Wear gloves when preparing these solutions.

Alternatives

- This experiment could be done with solutions of
  - lead(II) nitrate, $Pb(NO_3)_2$, and potassium iodide, KI
  - calcium nitrate, $Ca(NO_3)_2$, and sodium oxalate, $Na_2C_2O_4$
- Use only freshly distilled or boiled distilled water to make up these solutions, as dissolved carbon dioxide interferes with the results.

Disposal

Cobalt compounds are toxic. Collect all waste material in a waste container specified for this experiment. Acidify the solution with a minimum amount of 6.0 mol/L hydrochloric acid, and then precipitate out cobalt(II) sulphide by adding 3.0 mol/L solution of sodium sulphide. Allow the precipitate to settle. Decant the excess solution and flush it down the drain with excess water. Allow the precipitate to dry and arrange to have it disposed of in an appropriate manner.

The lead iodide formed must also be disposed of carefully. Collect all waste material in a waste container specified for this experiment. It should be converted into its most insoluble form. Then arrange to have it buried at an approved hazardous waste landfill site. Any waste solution should be treated with a *threefold* excess of sodium sulphide or thioacetamide and stirred occasionally for about one hour. Adjust the pH to neutral with 3 M sodium hydroxide, NaOH, solution to complete the precipitation of the lead compound. Separate the solid lead sulphide by filtration and allow it to dry. Place the lead sulphide in a plastic container of appropriate size, and have it buried at an appropriate landfill site. The filtrate should be added slowly with stirring to an excess of ferric chloride. A precipitate will form. Neutralize the remaining solution with sodium carbonate (some evolution of CO_2 will occur). Allow the precipitate to settle. Flush the neutral solution down the drain with excess water. Allow the precipitate to dry and have it buried at an appropriate landfill site.

Further Examination

Extend the experiment by showing which solutions have an excess of reagent. Have students remove a small amount of liquid from above the precipitate (supernatant liquid) in the 2nd test tube using a clean pipette, and transfer to either the well or another small test tube. Show that there is unreacted cobalt(II) chloride present by adding a drop or two of sodium hydroxide solution. Also, if an ammonium thiocyanate solution (5% NH_4SCN) is added, the turquoise tetrahedral cobalt(II) tetrathiocyanato complex, $Co(SCN)_4^{2-}$, should be observed.

Appendix 3.7: Stoichiometry: The Formula of a Precipitate (Discussion/Extensions—Teacher Notes) *(continued)*

Similarly, you can show unreacted sodium hydroxide in solutions by adding a drop or two of cobalt(II) chloride solution or a few drops of phenolphthalein indicator solution to a sample of the supernatant liquid.

Stoichiometry: Concept Extensions

A. By reacting solutions with a variety of known proportions of two substances, we can find the proportion giving the greatest amount of precipitate, and thereby determine the definite proportion for the two substances in that compound. Another way to determine the definite proportion of two substances reacting is to vary the proportion of two substances, and measure the amount of heat given off. The correct proportion will give off the most heat.

1. Prepare a solution of 0.50 M sodium hypochlorite (NaOCl, the active ingredient in bleach), and a solution of 0.50 M sodium thiosulphate ($Na_2S_2O_3$, the substance that photographers call "fixer").
2. Measure increasing quantities of NaOCl (see the following chart) into a foam cup. Measure decreasing quantities of $Na_2S_2O_3$ (see the following chart) into a beaker. Place a thermometer (with 1/10 degree graduations) in the foam cup. Pour the solution from the beaker into the cup and swirl it for two seconds. Immediately note the temperature and record the highest temperature reached after the solutions have been mixed.

Station #	NaOCl (mL)	$Na_2S_2O_3$ (mL)	Starting Temperature (°C)	Highest Temperature (°C)
1	10	90		
2	20	80		
3	30	70		
4	40	60		
5	50	50		
6	60	40		
7	70	30		
8	80	20		
9	90	10		

3. Ask students to analyze the data in the table to determine the definite proportions for these two substances. (The proportion of solutions producing the largest change in temperature is the correct proportion of the two compounds needed to react and form the product.)

Appendix 3.7: Stoichiometry: The Formula of a Precipitate (Discussion/Extensions—Teacher Notes) *(continued)*

B. An important chemical principle is the Law of Definite Proportions, which states that compounds will always have the same proportion of elements by mass regardless of where or how the compounds were formed. In other words, water has the same proportion of oxygen and hydrogen in Russia, Canada, and Africa. Furthermore, the active ingredient or compound in different brands will have the same proportion of elements in the compound no matter where they were produced.

Epsom salts is the common name for magnesium sulphate heptahydrate. If some Epsom salts are heated, water is given off. If all brands of Epsom salts have the same proportion of water, we should be able to verify this experimentally using the Law of Conservation of Mass.

1. Purchase three different brands of Epsom salts. Most pharmacies have their own generic brand. Mass different amounts of the three brands and assign them to different groups.
2. Place the Epsom salts into a large test tube and heat slowly but thoroughly for about 10 to 15 minutes to drive off the water. A Bunsen burner or other heat source could be used (e.g., Epsom salts can also be heated in an oven for a few hours at about 350°F [176°C].)
3. It is important to drive off all the water, regardless of the heating method used. If the Bunsen burner is used, weigh the test tube before and after the initial heating and again after each subsequent heating until two subsequent weights are within 0.2 g. Allow the test tube to cool to room temperature before weighing. If the oven is used, a second heating of about 30 minutes is needed to ensure that all the water has been driven off.
4. Ask groups to determine the percentage of water in each brand of Epsom salt using class results.

Appendix 3.7: Stoichiometry: The Formula of a Precipitate (Discussion/Extensions—Teacher Notes) *(continued)*

C. The table below presents some typical results using the method described in Experiment 14 in *Microscale Chemistry Laboratory Manual* (Slater and Rayner-Canham 39). An attempt was made to speed up the procedure by rinsing the copper metal with five small portions of propanone (acetone) using filter paper and a funnel. The propanone was allowed to drain thoroughly between rinses. The filter paper was then removed from the funnel and allowed to air-dry 15 minutes before weighing. While these results are reasonable, better results can be obtained by allowing the sample to dry overnight.

Take the opportunity to emphasize the importance of observation and interpretation in all experiments. Students might need to be led with questions such as these:

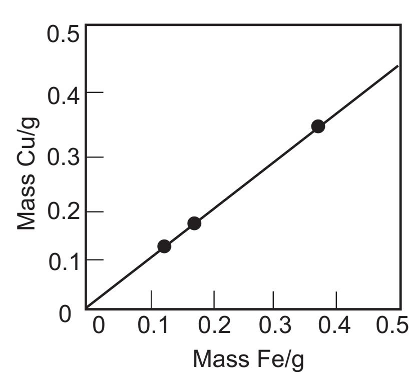
- What is the colour of copper(II) sulphate?
- What is the colour of the mixture of copper(II) sulphate and water?
- What does iron powder look like?
- What colour is the solid that you filtered? What does this tell you?
- What colour is the liquid that you filtered off? What does this mean?

Typical Results for Stoichiometric Reaction Producing a Solid

Trial	Mass $CuSO_4 \cdot 5H_20$	Mass Fe	Mass Cu	$\frac{\text{Mass Cu}}{\text{Mass Fe}}$	Comments
1	0.569 g	0.107 g	0.127 g	1.19	after drying overnight
2	0.954 g	0.131 g	0.162 g	1.24	washed with acetone; dried 15 minutes
	0.954 g	0.131 g	0.154 g	1.18	after drying overnight
3	1.490 g	0.329 g	0.417 g	1.27	washed with acetone; dried 15 minutes
	1.490 g	0.329 g	0.327 g	1.13	after dryng overnight

The graph of mass of copper versus mass of iron obtained from these results is given below. The line represents the theoretical line for a 1:1 stoichiometry—i.e., for the displacement reaction:

$$Fe_{(s)} + CuSO_{4(aq)} \rightarrow FeSO_{4(aq)} + Cu_{(s)}$$



Obviously any soluble copper compound could be used in place of copper(II) sulphate pentahydrate. Other 1:1 stoichimetric ratio combinations that work well include iron metal and any soluble tin(II) compound (e.g., tin[II] chloride) or zinc metal with any soluble lead(II) or copper(II) compound. A stoichiometric ratio of 1:2 can be demonstrated using copper metal with silver(I) nitrate.

Appendix 3.8: Indications of Chemical Reactions (Student Experiment)

Problem

Identify the characteristics that indicate the occurrence of a chemical reaction.

Equipment

- small (13 x 100 mm) test tubes
- medium (18 x 150 mm) test tubes
- medium (250 mL) beakers
- glass stirring rods
- medicine droppers
- wood splints

Materials

- ***Liquids***
  - ethanol, C_2H_5OH
  - phenolphthalein, 0.4% aqueous solution
  - sulphuric acid, H_2SO_4, concentrated
- ***Solids***
  - sugar crystals
  - zinc, Zn, powder
  - potassium permanganate, $KMnO_4$
  - diiodine (iodine), I_2, crystals,
  - calcium carbide, CaC_2
  - sodium thiosulphate, $Na_2S_2O_3$
- ***Aqueous Solutions, 1.0 mol/L***
  - lead nitrate, $Pb(NO_3)^2$
  - hydrogen chloride, $HCl_{(aq)}$
  - potassium chromate, K_2CrO_4
  - potassium iodide, KI
  - sodium sulphide, Na_2S
  - sodium hydroxide, NaOH
  - potassium dichromate, $K_2Cr_2O_7$

Appendix 3.8: Indications of Chemical Reactions (Student Experiment) *(continued)*

Safety Precautions

- Potentially toxic fumes may be liberated in the reaction of sugar with concentrated sulphuric acid. Perform the demonstration in a properly functioning fume hood.
- Sulphuric acid, H_2SO_4, is corrosive to skin and can cause serious eye damage. Wear eye protection at all times. If you get any solution on your skin or eyes, wash immediately with water for a minimum of 15 minutes.
- See the Material Safety Data Sheet before the lab demonstration/activity begins.

Procedure

Part A: Demonstration (Teacher Only)

1. Mix a small amount of potassium permanganate, $KMnO_4$, with a few drops of sulphuric acid, H_2SO_4. Use a glass stirring rod to mix the chemicals. Then touch the end of the rod used for mixing to some alcohol in an evaporating dish. Ask students to record their observations.

Part B: Student Experiments

1. Record all observations using a data sheet similar to the one that follows. Discard the resulting mixtures in the proper containers after each experiment.
2. Fill a medium test tube 1/4 full with sugar. Place the test tube in a beaker. Add 2 full droppers of sulphuric acid, H_2SO_4. **Caution:** Perform in a fume hood.
3. Fill a small test tube 1/4 full with aqueous sodium sulphide solution, Na_2S. Add 2 drops of lead nitrate solution, $Pb(NO_3)_2$.
4. Fill a small test tube 1/4 full with potassium iodide solution (KI). Add 2 drops of lead nitrate solution, $Pb(NO_3)_2$.
5. Mix 1/8 spoonful of zinc (Zn) with 1/8 spoonful of diiodine (I_2) in a small DRY test tube. Shake lightly to mix. Place the test tube in a beaker. Add 3 drops of water. **Caution:** Perform in a fume hood.
6. a) Fill a medium test tube 1/2 full with water. Place the test tube in a beaker, and add a pea-size piece of calcium carbide, CaC_2.

   b) Quickly light the gas evolved with a burning splint.
7. Fill a small test tube 1/4 full with hydrogen chloride solution, $HCl_{(aq)}$. Add 3 drops of the phenolphthalein indicator solution. With a dropper, slowly add aqueous sodium hydroxide, $NaOH_{(aq)}$.
8. Fill a medium test tube 1/3 full with water. Dissolve 1/4 spoonful of sodium thiosulphate, $Na_2S_2O_3$, in the water. Does the tube feel warmer or colder than it did before the sodium thiosulphate was added? Note the temperature change.

Appendix 3.8: Indications of Chemical Reactions (Student Experiment) *(continued)*

9. a) Add a few drops of hydrogen chloride solution, $HCl_{(aq)}$, to 1/8 test tube aqueous potassium chromate solution, K_2CrO_4.
   b) Add a few drops of hydrogen chloride solution, $HCl_{(aq)}$, to 1/8 test tube aqueous potassium dichromate solution, $K_2Cr_2O_7$.
   c) Add a few drops of aqueous NaOH to 1/8 test tube aqueous potassium chromate solution, K_2CrO_4.
   d) Add a few drops of aqueous NaOH to 1/8 test tube aqueous potassium dichromate solution, $K_2Cr_2O_7$.

Indications of Chemical Reactions—Data Sheet		
Experiment	**Observations**	**Chemical Reaction (Equation)**
1		
2		
3		
4		
5		
6		
7		
8		
9		

Question

1. List four characteristics or indicators of a chemical reaction.

Appendix 3.8: Indications of Chemical Reactions (Alternate): Chemistry in a Bag (Student Experiment)

Problem

To identify characteristics of a chemical reaction.

Materials

- a dilute phenolphthalein solution made by adding 3 drops of phenolphthalein to 30 mL water
- 10 g sodium bicarbonate, $NaHCO_3$
- 7 g calcium chloride, $CaCl_2$
- paper cups
- resealable plastic bag (e.g., plastic zippered bag)

Procedure

Caution: Goggles must be worn!

1. Carefully measure the proper amounts of sodium bicarbonate and calcium chloride in separate paper cups or weigh boats.
2. Carefully pour the samples into opposite corners of the bottom of a resealable bag.

   Note: Do not mix the chemicals at this point!
3. Lay the bag carefully on its side on a lab bench. Gently pour in the phenolphthalein solution and, without picking up the bag, seal it. Push the solution from the centre of the bag outwards to the opposite corners with your fingers, pressing from the outside of the bag. Make note of any changes in the appearance of the chemicals in each corner as the solution makes contact. Touch the bag at each corner and make note of any temperature changes.
4. Swish the bag gently to mix all the contents, and again make observations as to any changes you see, hear, or feel.

Questions

1. What changes took place when the sodium bicarbonate came into contact with the phenolphthalein solution?
2. What changes took place when the calcium chloride came into contact with the phenolphthalein solution?
3. What changes took place when all three substances were mixed together?
4. From these observations, list three characteristics of a chemical reaction.

   a) ____

   b) ____

   c) ____
5. Do you think that the mass of the bag and contents changed as a result of this chemical reaction? How could you check?

Appendix 3.8: Indications of Chemical Reactions (Discussion—Teacher Notes)

Safety Precautions

- The reaction of sugar with sulphuric acid is hazardous! Concentrated sulphuric acid is very corrosive.
- Spills must immediately be neutralized with sodium bicarbonate, and diluted with copious amounts of water.
- Spills on skin or clothing should be immediately bathed in water for a minimum of 5 minutes.
- Spills in the eyes should immediately be rinsed with water for a minimum of 15 minutes.
- Further medical attention should be sought for spills on skin or eyes.
- See Material Safety Data Sheets.

Reaction of Sugar with Sulphuric Acid

The reaction of sugar with concentrated sulphuric acid is quite slow to begin, but will eventually be vigorous, with copious amounts of steam and a rapidly growing column of black carbon. The best results for this experiment have been obtained in a large (20 x 200 mm) test tube containing about 10 mL of sugar and 7 mL of water, to which is added 10 mL of concentrated sulphuric acid. Even half these quantities will yield a sufficient reaction. The major cause of foaming is the production of steam.

Concentrated sulphuric acid is a very powerful dehydrating agent, as is shown by this demonstration.

The sulphuric acid removes water from sugar. The reaction may be represented by the equation:

$$C_{12}H_{22}O_{11(s)} + H_2SO_{4(l)} \rightarrow 12C_{(s)} + 11H_2O_{(g)}$$

However, the reaction is more complex. More than 50% of the gas evolved is carbon monoxide. Other products include sulphur dioxide, carbon dioxide, and other carbon-containing compounds.

Disposal Procedures

- Dispose of the carbon "sausage" and test tube carefully. The unreacted sulphuric acid should be neutralized with sodium bicarbonate and washed with water before disposing of the carbon "sausage" in the waste.
- Solutions or precipitates containing lead or chromium should be carefully collected and disposed of in a secure landfill site specifically designed for such hazardous wastes.
- The solutions of acids or bases should be neutralized before being flushed down the drain. Solutions of sodium thiosulphate (hypo) can also be flushed down the drain.

Appendix 3.9: Determining the Molar Mass of a Gas (Student Experiment)

Purpose

The molar mass of a compound is an important constant that, in some cases, can help identify a substance. In this lab, you will calculate the molar mass of butane using calculations involving the combined gas law and the constant 22.4 L/mole.

Materials

- goggles
- butane lighter (with flint removed)
- plastic bucket
- water
- graduated cylinder (1000 mL)
- funnel
- balance
- thermometer
- barometer

Procedure

1. Determine the initial mass of the butane lighter.
2. Pour water into the bucket until it is about three-quarters full. Then fill the graduated cylinder with water and invert it into the bucket so that the water level is within the calibrated region. Record the volume reading.
3. Place the funnel in the mouth of the graduated cylinder while it is under the water to ensure all butane gas bubbles are collected.
4. Hold the butane lighter in the water under the graduated cylinder and funnel apparatus. Release the butane gas from the lighter into the mouth of the graduated cylinder until you have displaced from half to three-quarters of the water within the cylinder.
5. Equalize the pressures inside and outside the cylinder by adjusting the position of the cylinder until the water levels inside and outside the cylinder are the same.
6. Read the measurement on the cylinder and record the volume of gas collected.
7. Record the ambient (room) temperature and pressure.
8. Thoroughly dry the butane lighter and determine its final mass.

Appendix 3.9: Determining the Molar Mass of a Gas (Student Experiment) *(continued)*

Observations

Data Chart	
Initial mass of lighter	
Final mass of lighter	
Mass of gas released	
Initial volume reading on graduated cylinder	
Final volume reading on graduated cylinder	
Volume of gas released	
Room temperature	
Room air pressure	

Calculations

1. Using the combined gas law, convert the volume of gas released in the lab to the volume it would occupy at standard temperature and pressure (STP).
2. Use the volume of gas at STP (recorded above) and the gas constant of 22.4 L/mole to determine the number of moles of gas collected at STP.
3. Use the mass of gas released (from the data chart) and divide it by the moles of gas at STP to find the molar mass of the gas.

Conclusions

1. What is the molar mass of the butane according to your lab results?
2. What is the known molar mass of butane according to the periodic table?
3. What is your experimental percent error?
4. Every experiment has some experimental error or uncertainties. State some possible flaws, limitations, experimental errors, or uncertainties that may affect the accuracy of your results. Make a list and rank them from most to least significant.
5. It is always possible to omit a source of experimental error. In the experiment, you may not have realized that butane in a lighter is not pure butane but contains a small quantity of water vapour. For a typical room temperature, this would correspond to about 2.6 kPa of the pressure you recorded. Subtract this value from the pressure you used, and use the new pressure to recalculate the molar mass and the experimental percent error. Is your answer significantly more accurate? Explain.

Appendix 3.9: Determining the Molar Mass of a Gas (Teacher Notes)

Safety Precautions

- Butane is highly flammable.
- Do not conduct this experiment near an open flame.
- Good ventilation in the laboratory is essential.
- Eye protection is required.
- Flints must be removed from the butane lighters. You can pry off the metal casing (hood) and the spark wheel of the typical disposable commercial lighter without much effort, and the flint and a long spring will just pop out.

Important Notes

- You will need one butane lighter per group.
- An ice cream pail (4 L) works fine for this experiment, but because it is small it is a little clumsy to use. The funnel could be removed so that you have more hand room, but then students have to be more careful not to lose the bubbles. A sink filled with water works well also.
- If you don't have a barometer, you can find the air pressure for your town or city on a weather website on the Internet.
- After thoroughly drying the lighter, you may also want to let it air dry for a while so that the interior parts can dry as well. A more accurate mass can be recorded after this is done.

Observations

Sample Data Chart	
Initial mass of lighter	18.17 g
Final mass of lighter	18.01 g
Mass of gas released	0.16 g
Initial volume reading on graduated cylinder	21.0 mL
Final volume reading on graduated cylinder	89.8 mL
Volume of gas released	68.8 mL
Room temperature	22°C = 295 K
Room air pressure	102.14 kPa

Appendix 3.9: Determining the Molar Mass of a Gas (Teacher Notes) *(continued)*

Calculations

1. $$\frac{P_1V_1}{T_1} = \frac{P_2V_2}{T_2}$$

$$\frac{(102.14)(68.8)}{295} = \frac{(101.3)(V_2)}{273} = 64.2\text{ mL}$$

2. Since one mole of butane occupies a volume of 22.4 L at STP, then 0.0642/22.4 = 0.00287 moles of gas is collected.
3. Since 0.00287 moles of butane gas has a mass of 0.16 grams, then 0.16/0.00287 = 55.8 g is the molar mass of butane based on our experimental data.

Conclusions

1. 55.8 g/mole
2. 58.1 g/mole
3. $$\frac{\text{absolute difference between experimental and theoretical value}}{\text{theoretical value}} \times 100$$

$$\frac{(58.1 - 55.8)}{58.1} \times 100 = 4.0\%$$

4. Answers may vary, but could include:
   - massing of butane lighter after the lab when it still had traces of water
   - losing some bubbles of butane out of the graduated cylinder
5. The answer is more accurate; therefore, this unknown source of error was significant.

Appendix 3.10: Gas Density Table (Student Resource Material)

Density of Gases at 25°C and 101.3 kPa (760 mmHg or 1.0 atm) Pressure

Name	Formula	Molar Mass (g/mol)	Density (g/L)
Ammonia	NH_3	17.03	0.696
Argon	Ar	39.944	1.633
Butane	C_4H_{10}	58.12	2.376
Carbon dioxide	CO_2	44.01	1.799
Carbon monoxide	CO	28.01	1.145
Dichlorine	Cl_2	70.91	2.898
Ethane	C_2H_6	30.07	1.229
Ethene	C_2H_4	28.05	1.147
Ethyne (acetylene)	C_2H_2	26.04	1.064
Helium	He	4.003	0.164
Dihydrogen	H_2	2.016	0.082
Hydrogen chloride	HCl	36.47	1.490
Hydrogen iodide	HI	127.93	5.228
Krypton	Kr	83.70	3.425
Methane	CH_4	16.04	0.656
Neon	Ne	20.18	0.825
Nitrogen monoxide	NO	30.01	1.226
Dinitrogen	N_2	28.02	1.145
Dinitrogen monoxide	N_2O	44.02	1.799
Nitrogen dioxide	NO_2	46.01	1.880
Dioxygen	O_2	32.00	1.308
Ozone	O_3	48.00	1.962
Propane	C_3H_8	44.09	1.802
Sulphur dioxide	SO_2	64.07	2.618
Xenon	Xe	131.30	5.367

Appendix 3.11: Creative Mole: Writing Activity

You have just been introduced to the concept of *mole*. It is a relatively abstract concept, as it is difficult to picture 6.02×10^{23} particles of anything. This creative writing assignment will let you explore this topic and become more comfortable with the new vocabulary.

The assignment will be done in the form of a RAFT:*

R → **Role of writer:** Who are you?

A → **Audience:** To whom are you writing?

F → **Format:** What form will you use? (e.g., letter, rap song, poem, advertisement)

T → **Topic and strong verb:** What important topic have you chosen? What strong verb describes your intent? (e.g., to persuade, demand, plead)

Below is a list of possible ideas for your RAFT writing activity. You may use one of these ideas, or you may choose one of your own. If you choose your own idea, please discuss it with the teacher before you start writing.

Role	Audience	Format	Topic
Atom of lead (Pb)	Other lead atoms	Want ad	Convince others to join you in a heavy metal mole band.
Pile of salt	Other formula units	News release	Announce the discovery of how many moles make up your pile.
Mole	Periodic table	Rap song	Explain how to find your molar mass.
Real estate agent	Mole of oxygen	Real estate ad	Sell the ideal home for a mole of oxygen.

Your goal is to show that you know how to convert between moles, volume, mass, and number of particles and that you understand the vocabulary.

Assessment Criteria

- At least five new vocabulary words are used correctly/explained.
- Calculations are correctly described.
- The writing is interesting and creative.
- The writing is clear. Correct spelling and grammar are used.
- The assignment is sufficient in length (from half a page to a page).

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* **Reference:** Santa, C.M. *Context Reading Including Study Systems*. Dubuque, IA: Kendall Hunt, 1988.

Appendix 3.12: The Stoichiometry of Gasoline: Internet Research Activity

Instructions

Use the Internet and your knowledge of stoichiometry to answer the following questions. Fuel consumption ratings and CO_2 emissions data can be found at the following website.

Natural Resources Canada:
<http://oee.nrcan.gc.ca/transportation/personal-vehicles-initiative.cfm>

Click on the link that says “Fuel Consumption Ratings.”

1. The energy used to propel a vehicle comes from the combustion of gasoline, C_8H_{18}. Write a balanced chemical equation for the combustion of gasoline.
2. Most gas stations offer three different grades of gasoline: regular, mid-grade, and premium. What is the difference between these three types of gasoline?
3. Answer the remaining questions, using *three* of the following vehicles as your examples:

- 2007 Lexus SC400
- 2007 Ford X150 Pickup Truck
- 2007 Honda Civic Hybrid (Gas/electric)
- 2007 VW Jetta TDI Diesel
- 2007 Honda Odyssey Van

a) Identify the size of the gas tank for each of the three selected vehicles.
b) Determine the highway fuel economy for each of the three vehicles.
c) Determine the cost of regular and premium gasoline. Use this information to calculate how much it would cost to fill the gas tank for each of the three vehicles. (Note that the Lexus requires premium gasoline.)
d) If one litre of gasoline contains 6.18 moles of gas, how many moles of gas fit in the tank of each vehicle?
e) What is the mass of a full tank for each vehicle?
f) If you were to take each vehicle on a road trip, approximately how far could you drive if you began with a full tank of gas?
g) If you were to use an entire tank of gas for each vehicle, how much carbon dioxide would be emitted?
   - How many *moles* of carbon dioxide would be emitted from each vehicle?
   - How many *grams* of carbon dioxide would be emitted from each vehicle?
   - How many *litres* of carbon dioxide would be emitted from each vehicle?

Appendix 3.13: How to Solve a Limiting Reactant Problem

Question

Roasting chalcocite (a copper sulphide) is the first step in extracting copper metal from this sulphide ore. Note that a precursor to sulphuric acid is one of the products of this "roasting of ore." Think of the implications in the atmosphere when this product connects with water vapour in the air.

$$2Cu_2S_{(s)} + 3O_{2(g)} \rightarrow 2Cu_2O_{(s)} + 2SO_{2(g)}$$

If 30.00 g of Cu_2S is reacted with 7.47 L of O_2 at STP, what is the maximum mass of copper oxide produced?

Solution

Reactant	Moles of Reactant	Moles Needed to React Completely	Type of Reagent
Cu_2S	$30.00\text{ g} \times \frac{1\text{ mol}}{159.16\text{ g}}$ = 0.1885 mol Cu_2S	0.1885 mol Cu_2S x 3 mol O_2/2 mol Cu_2S = 0.2878 mol O_2 needed	Since you need 0.2223 mol of Cu_2S, but you only have 0.1885 mol of Cu_2S, Cu_2S is the limiting reactant.
O_2	$7.47\text{ L} \times \frac{1\text{ mol }O_2}{22.4\text{ L}}$ = 0.3333 mol O_2	0.3333 L O_2 x 2 mol Cu_2S/3 mol O_2 = 0.2223 mol Cu_2S needed	Since you only need 0.2878 mol of the 0.3333 mol of O_2, O_2 is the excess reactant.

$$0.2223\text{ mol }Cu_2S \times \frac{159.16\text{ g }Cu_2S}{1\text{ mol }Cu_2S} = 35.38\text{ g }Cu_2S$$

Appendix 3.14: The Behaviour of Solid Copper Immersed in a Water Solution of the Compound Silver Nitrate

In this experiment you will weigh a sample of solid silver nitrate and prepare a water solution of it. You will also weigh a piece of copper wire, place it in the solution, and observe its behaviour. By weighing the copper wire at the close of the experiment you will be able to investigate quantitatively any changes that occur.

Before coming to the laboratory, prepare a data table in your laboratory notebook so you can record the data you will observe.

Materials

- copper wire, Cu (30 cm—No. 16 is suitable)
- vial of silver nitrate, $AgNO_3$ (provided by teacher)
- large test tube
- beaker (250 mL)
- solid glass rod
- balance
- distilled water
- wash bottle
- drying apparatus

Procedure

1. Obtain a 30 cm length of copper wire. Form a coil by wrapping the wire around a large test tube, leaving about 7 cm straight for a handle. Stretch the coil a little so there is some space between the loops (see Figure 1). Weigh the copper coil to the nearest 0.01 g.
2. Weigh a clean, thoroughly dry 250 mL beaker to the nearest 0.01 g. Weigh the vial of silver nitrate, $AgNO_3$, provided by the teacher.

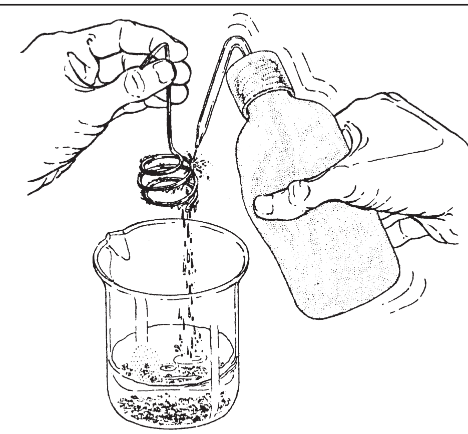

Figure 1: Rinsing with a wash bottle.

3. Fill the weighed 250 mL beaker about two-fifths full with *distilled* water. Add the solid silver nitrate, $AgNO_3$, to the water. Stir gently with a solid glass rod until all the $AgNO_3$ crystals have dissolved. Weigh the empty vial.

The Behaviour of Solid Copper Immersed in a Water Solution of the Compound Silver Nitrate—Source: Chemical Education Material Study. *Chemistry: An Experimental Science Laboratory Manual*. San Francisco, CA: W.H. Freeman and Company, 1963. 19–21. Adapted by permission of W.H. Freeman and Company/Worth Publishers.

Appendix 3.14: The Behaviour of Solid Copper Immersed in a Water Solution of the Compound Silver Nitrate *(continued)*

Caution: Silver nitrate, solid or solution, reacts with skin and will stain it black. Be careful and avoid spillage on your skin and clothing. However, don't be alarmed if you discover dark spots on your hands—they wear away in a few days. Clean hands the day following this experiment indicate good laboratory technique.

4. Bend the handle of the weighed copper wire so that it can be hung over the edge of the beaker with the coil immersed in the $AgNO_3$ solution. Place the coil into the beaker and observe any changes that take place for several minutes at least.
5. Cover the beaker with a watch glass and place it in your locker until the next laboratory period.
6. At the beginning of the next laboratory period, very carefully open your locker and lift the beaker to the desktop. Observe what has happened in the beaker. Record *all* your observations in your laboratory notebook.
7. Shake the crystals off the coil and lift the coil from the solution. Use your wash bottle to rinse into the beaker any crystals that tend to adhere to the coil. (See Figure 1.) Set the coil aside to dry. Weigh it when dry.
8. Let the crystals settle in the beaker. Carefully decant the solution. Decant means to pour off liquid, leaving solid behind, as shown in Figure 2. Add 5 mL of dilute silver nitrate solution and stir gently until any flecks of copper disappear. Carefully decant again. Wash the residue with 10 mL of water and carefully decant. Wash and decant at least three more times. You may neglect the few particles that may float over with the wash water since the amount is usually not weighable.

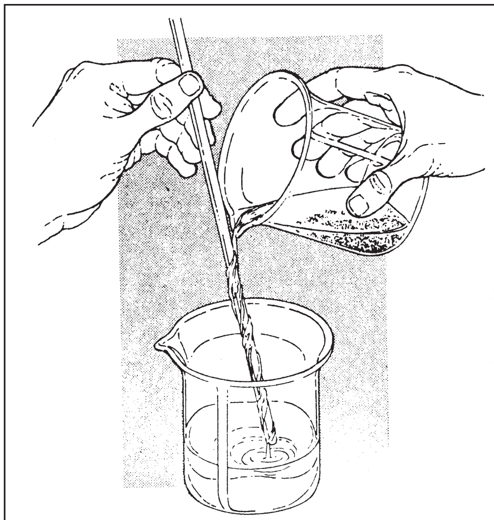

Figure 2: Decanting the supernatant liquid from the residue.

9. After the final washing, the residue must be dried. The teacher will suggest a suitable method. If the sample is dried overnight with heat lamps or in a drying oven, it should be dry when you return to the laboratory. Allow the beaker and contents to cool before weighing. Use the same balance as you used previously and record the weight together with the uncertainty.

Note: If a sand bath is used to dry the sample, you can make sure that it is dry as follows. Weigh the sample and beaker, then return the sample to the sand bath and heat it for a second time. Weigh it again. If weight was lost, you did not have a dry sample and it may not be dry this time, so heat and weigh it again. Repeat the procedure until a constant weight is obtained.

Appendix 3.14: The Behaviour of Solid Copper Immersed in a Water Solution of the Compound Silver Nitrate *(continued)*

Data Table

Your data table should include the following. Be sure to include uncertainty as part of your recorded data.

Data Table	
Weight of copper before immersion in solution	
Weight of copper at close of experiment	
Weight change of copper	
Weight of vial plus silver nitrate	
Weight of vial	
Weight of silver nitrate	
Weight of beaker plus silver	
Weight of beaker	
Weight of silver	

Calculations

1. Calculate the number of moles of copper that reacted.
2. Calculate the number of moles of silver obtained.
3. Determine the ratio of moles of silver to moles of copper involved in this reaction. Be sure to express your calculations using the correct number of significant figures.

Questions

1. What you have observed can be described by the following statement:

   One mole of copper (*solid*) + ____________ mole(s) of

   silver nitrate (*in water*) → __________mole(s) of silver (*solid*)

   + ____________ mole(s) of copper nitrate (*in water*).

   Using the results obtained in this experiment, write the proper whole number coefficients in the above statement when 1 mole of copper is used up.
2. How many atoms of solid copper were involved in your experiment?
3. How many atoms of solid silver were involved in your experiment?
4. What is the relationship between the number of atoms of silver and the number of atoms of copper calculated in questions 2 and 3?

Appendix 3.14: The Behaviour of Solid Copper Immersed in a Water Solution of the Compound Silver Nitrate *(continued)*

5. In order to evaluate the results of this experiment, the teacher will collect the data obtained by other members of your class. Make a graph, plotting the number of individuals obtaining a given silver/copper ratio along the vertical axis. Plot the Ag/Cu ratios along the horizontal axis. These should be rounded off so that each division on the graph will represent values of ±0.05. For example, values from 1.85 up to but not including 1.95 should be plotted as 1.9.
6. Considering only the middle two-thirds of the data plotted, what is the range of values obtained? How does this compare with the uncertainty you considered justifiable from your measurements?

Questions to Wonder About

1. What causes the colour in the solution after the reaction is completed?
2. What is the nature of the particles in aqueous solution?

Appendix 3.15: A Quantitative Investigation of the Reaction of a Metal with Hydrochloric Acid

In this activity, you will determine the volume of hydrogen gas that is produced when a sample of magnesium metal reacts with hydrogen chloride (also recognized as hydrochloric acid) dissolved in water. The volume of the hydrogen gas will be measured at room temperature and pressure—conditions that matter for gases.

Question

The data you obtain from this experiment will enable you to answer the following question:

How many litres of dry hydrogen gas at room temperature and 1 atmosphere (atm) can be produced per mole of magnesium metal?

Materials

- magnesium, Mg, ribbon (5 cm)
- fine copper wire
- ring stand and utility clamp
- gas measuring tube (50 mL) fitted with one- or two-hole rubber stopper
- beaker (400 mL)
- hydrochloric acid
- water
- large cylinder or battery jar

Procedure

1. Obtain a piece of magnesium, Mg, ribbon approximately 5 cm long. Measure the length of the ribbon carefully and record this to the nearest 0.05 cm. The teacher will give you the weight of 1 metre of the ribbon, and since it is uniform in thickness you can calculate the weight of the magnesium used.
2. Fold the magnesium ribbon so that it can be encased in a small spiral cage made of fine copper wire. Leave about 5 cm of copper wire to serve as a handle (see Figure 1).
3. Set up a ring stand and utility clamp in position to hold a 50 mL gas measuring tube that has been fitted with a one- or two-hole rubber stopper, as shown in Figure 1. Place a 400 mL beaker about two-thirds full of tap water near the ring stand.
4. Incline the gas measuring tube slightly from an upright position and pour in about 10 mL of moderately concentrated hydrochloric acid labelled 6 *M* HCl.

A Quantitative Investigation of the Reaction of a Metal with Hydrochloric Acid—Source: Chemical Education Material Study. *Chemistry: An Experimental Science Laboratory Manual.* San Francisco, CA: W.H. Freeman and Company, 1963. 26–30. Adapted by permission of W.H. Freeman and Company/Worth Publishers.

Appendix 3.15: A Quantitative Investigation of the Reaction of a Metal with Hydrochloric Acid *(continued)*

5. With the tube in the same position, slowly fill it with tap water from a beaker. While pouring, rinse any acid that may be on the sides of the tube so that the liquid in the top of the tube will contain very little acid. Try to avoid stirring up the acid layer in the bottom of the tube. Bubbles clinging to the sides of the tube can be dislodged by tapping the tube gently.
6. Holding the copper coil by the handle, insert the metal about 3 cm down into the tube. Hook the copper wire over the edge of the tube and clamp it there by inserting the rubber stopper. The tube should be completely filled so that the stopper displaces a little water when put in place (see the left of Figure 1).
7. Cover the hole(s) in the stopper with your finger and invert the tube in the container of water, as shown in the middle of Figure 1. Clamp it in place. The acid, being denser than water, will diffuse down through it and eventually react with the metal.
8. After the reaction stops, wait for about five minutes to allow the tube to come to room temperature. Dislodge any bubbles clinging to the sides of the tube.
9. Cover the hole(s) in the stopper with your finger and transfer the tube to a large cylinder or battery jar that is almost filled with water at room temperature (see Figure 2). Raise or lower the tube until the level of the liquid inside the tube is the same as the level outside the tube. This permits you to measure the volume of the gases in the tube (hydrogen and water vapour) at room pressure. Read the volume with your eye at the same level as the bottom of the *meniscus* (the lens shape surface taken by the water in the tube). Record the volume of the gas to the nearest 0.05 mL.
10. Remove the gas measuring tube from the water and pour the acid solution it contains down the sink. Rinse the tube with tap water.
11. Record the room temperature. The teacher will give you the room pressure or will assist you in reading the barometer to obtain a value for the pressure in the room.

The experiment may be repeated with another sample of magnesium to check your results, if time permits.

Appendix 3.15: A Quantitative Investigation of the Reaction of a Metal with Hydrochloric Acid *(continued)*

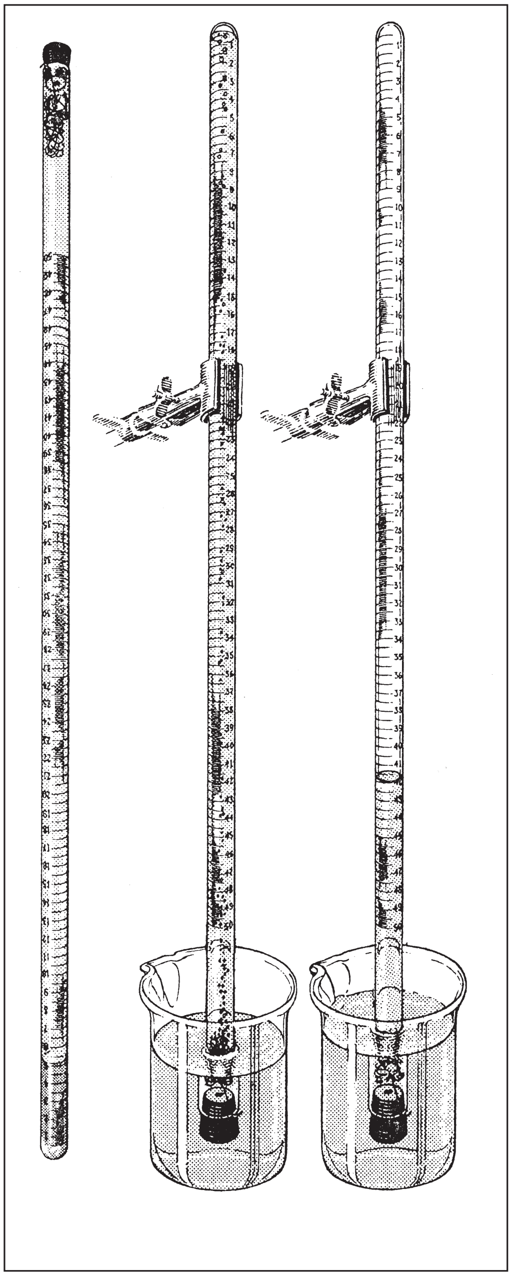

Figure 1:
Manipulating the gas measuring tube.

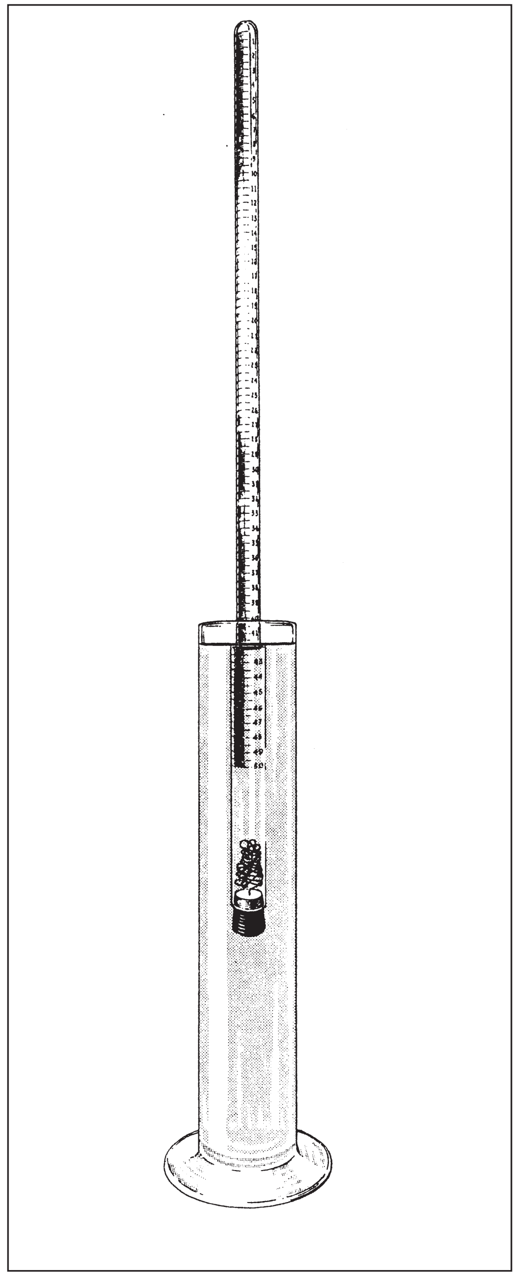

Figure 2:
Measuring the volume of gas.

Appendix 3.15: A Quantitative Investigation of the Reaction of a Metal with Hydrochloric Acid *(continued)*

Data Table

The data table should include the following.

Data Table	
Weight of magnesium ribbon in grams per metre (from teacher)	
Length of magnesium ribbon	
Volume of hydrogen (saturated with water vapour)	
Temperature of the water	
Temperature of the room	
Barometer reading (room pressure)	
Vapour pressure of water at the above temperature (see the following table)*	

* You will need this information in order to perform the necessary calculations.

Vapour Pressure of Water at Various Temperatures	
Temperature (°C)	**Pressure (mm)**
15	12.8
16	13.6
17	14.5
18	15.5
19	16.5
20	17.5
21	18.6
22	19.8
23	21.0
24	22.4
25	23.8
26	25.2
27	26.7
28	28.3
29	30.0
30	31.8

Appendix 3.15: A Quantitative Investigation of the Reaction of a Metal with Hydrochloric Acid *(continued)*

Calculations

1. Determine the weight of the magnesium you used from the grams-per-metre relationship and the length of the ribbon.
2. Determine the number of moles of magnesium used.
3. Determine the partial pressure of the hydrogen gas.

   Since the hydrogen gas was collected over water, the gas in the tube consists of a mixture of hydrogen gas and water vapour. The total pressure caused by these two gases is equal to the room pressure. Mathematically this can be expressed as follows:

   $$P_{H_2} + P_{H_2O} = P_{room}$$

   The pressure of the room may be determined by reading the barometer. The pressure of the water vapour, P_{H_2O}, can be determined from the table given above. The values in the table were obtained by measuring the pressure of water vapour above liquid water at various temperatures. The partial pressure of the hydrogen can then be calculated as follows:

   $$P_{H_2} = P_{room} - P_{H_2O}$$

4. Determine the volume of the hydrogen gas at 1 atmosphere pressure (760 mm) and 0°C.

   You have learned that for a given temperature the product of the pressure and volume of a gas is a constant. PV = k. To calculate the new volume, V_{new} at 760 mm pressure, the following mathematical relationship can be stated:

   $$V_{measured}P_{H_2} = V_{new}760$$

   or

   $$V_{new} = V_{measured} \times \frac{P_{H_2}}{760}$$

5. Calculate the volume of dry hydrogen that would be produced by 1 mole of magnesium at room temperature and 1 atmosphere pressure.

Questions

1. Given that 1 mole of Mg produces 1 mole of hydrogen, H_2, what is the volume of 1 mole of hydrogen at room temperature and 1 atmosphere pressure?
2. If 1 mole of hydrogen weighs 2.0 g, what is the weight of a litre (the density) of hydrogen at room temperature and 1 atmosphere pressure?

Appendix 3.16: Stoichiometry: Reactants, Products, and Enthalpy Changes (Student Experiment)

Through a series of chemical reactions involving the same two reactants, we will monitor the amount of product produced. By holding the total moles of reactants used constant and varying their relative amounts, maximum product formation will be determined. This method allows us to monitor the amount of product produced using the most convenient method available—in this case, the temperature. Any property that allows quantitative measurement of a product can be used.

Question

By combining various mole ratios of reactants and analyzing their products, what will be the reactant coefficients of a chemical reaction?

Prediction

If we determine the maximum amount of heat produced by a certain ratio of reactants, what will their coefficients be?

Equipment

- computer system and temperature sensor
- thermometer
- stirring rod
- foam cups
- 2 graduated cylinders
- 5% sodium hypochlorite, NaOCl—household laundry bleach (500 mL – 0.05 M)
- potassium iodide solution, KI (500 mL – 0.05 M)

Be sure the chemicals and apparatus are at room temperature before beginning.

Safety Precautions

- Wear a lab apron and safety goggles.
- Sodium hypochlorite is a severe irritant to eyes, skin, and mucous membrane.
- Potassium iodide is harmful if swallowed.
- Use caution with all solutions, as some are caustic, poisonous, or will stain clothing.
- Do not pour anything down the drain.

Appendix 3.16: Stoichiometry: Reactants, Products, and Enthalpy Changes (Student Experiment) *(continued)*

Procedure

1. Prepare a data table as shown below.

   (**Note:** The "average minimum temperature" for each row will be the same, as all the reactions will begin at the same temperature.)

Data Table

Trial #	mL Ratio	Volume of NaOCl Used (mL)	Moles of NaOCl Used (mol)	Volume of KI Used (mL)	Moles of KI Used (mol)	Average Maximum Temperature (°C)	Average Minimum Temperature (°C)	Change in Temperature (°C)
1	20:80	20		80				
2	40:60	40		60				
3	60:40	60		40				
4	80:20	80		20				

2. Connect the computer system with the temperature sensor and create a digits display and a graph display of the date.
3. Start the data recording.
4. Use a graduated cylinder to measure and pour 20 mL of NaOCl into a clean foam cup.
5. Measure the temperature of the 20 mL of NaOCl.
6. Using another graduated cylinder, measure 80 mL of KI and record its temperature.
7. Enter the average of the two temperatures into the Trial #1 row and under the Average Minimum Temperature column in the table.
8. Pour the KI into the foam cup with the NaOCl while recording the temperature. Gently stir the mixture with the temperature sensor.
9. After the temperature rises and once the temperature begins to fall, stop recording.
10. Remove the thermometer and stirring rod and wipe both clean.
11. Repeat steps 3 to 10 using volume ratios of NaOCl and KI at 40:60, 60:40, and 80:20.
12. Discard the solutions as directed by the teacher. Do not pour anything down the drain.

Appendix 3.16: Stoichiometry: Reactants, Products, and Enthalpy Changes (Student Experiment) *(continued)*

Questions

Analysis

1. For each of the volumes listed for both NaOCl and KI, calculate the number of moles of each used in the trials (remember that each chemical's concentration is 0.50 M and number of moles = M x volume). Enter these values into the table.
2. For each trial, find the maximum temperature attained. Record this under the Average Maximum Temperature column.
3. Calculate and record the Change in Temperature by subtracting the average minimum temperature from the average maximum temperature in each row.
4. With which trial number did the temperature rise the most?
5. What was this maximum temperature change?
6. What was the ratio of volume of NaOCl to KI during the maximum change in temperature?
7. What was the ratio of moles of NaOCl to KI during the maximum change in temperature?

Conclusions

8. What are the simplest whole numbers, making use of the maximum temperature mole ratio, that would satisfy the equation xNaOCl + yKI → products + heat? You are not expected to know the coefficients of the products.
9. Which chemical was the limiting agent?

Applications

10. Other reactants you could mix with NaOCl might include sodium thiosulphate, $Na_2S_2O_2$, and sodium sulphite, Na_2SO_3.

Appendix 3.16: Stoichiometry: Reactants, Products, and Enthalpy Changes (Teacher Notes)

Sample Data: Anwer Key

Observations

Data Table								
Trial #	mL Ratio	Volume of NaOCl Used (mL)	Moles of NaOCl Used (mol)	Volume of KI Used (mL)	Moles of KI Used (mol)	Average Maximum Temperature (°C)	Average Minimum Temperature (°C)	Change in Temperature (°C)
1	20:80	20	0.01	80	0.04	22.4	19.3	3.1
2	40:60	40	0.02	60	0.03	25.9	19.3	6.6
3	60:40	60	0.03	40	0.02	29.6	19.3	10.3
4	80:20	80	0.04	20	0.01	27.7	19.3	8.4

Analysis

1.

Trial #	Volume of NaOCl Used (mL)	Moles of NaOCl Used (mol)	Volume of KI Used (mL)	Moles of KI Used (mol)
1	20	0.01	80	0.04
2	40	0.02	60	0.03
3	60	0.03	40	0.02
4	80	0.04	20	0.01

2.

Trial #	Average Maximum Temperature (°C)	Average Minimum Temperature (°C)
1	22.4	19.3
2	25.9	19.3
3	29.6	19.3
4	27.7	19.3

Appendix 3.16: Stoichiometry: Reactants, Products, and Enthalpy Changes (Teacher Notes) *(continued)*

3.

Trial #	Change in Temperature (°C)
1	3.1
2	6.6
3	10.3
4	8.4

4. Trial #3
5. Approximately 10.3°C
6. 60:40 or 3:2
7. 0.03:0.02 or 3:2

Conclusions

8. $3NaOCl + 2KI \rightarrow$ products + heat
9. NaOCl was the limiting reagent.

Applications

10. Other reactants you could mix with NaOCl might include sodium thiosulphate, $Na_2S_2O_2$, and sodium sulphite, Na_2SO_3.