

APPENDICES

Appendices

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Appendix A. Physical Properties of Gases at 1 atm

	MM (g/mol)	mp °C	bp °C	density: $d_{\text{gas}}/d_{\text{air}}$	density ^a g/L	Solubility in H ₂ O at 0°C		
						g/100 mL	moles/L	mL/dL ^b
C₂H₂	26.0379	-80.8	-84.0	0.90	1.0643	0.106	0.041	99.600
C₂H₄	28.0539	-169	-103.7	0.97	1.1467	0.013 ^a	0.00449 ^a	10.99 ^a
CH₄	16.0429	-182	-164	0.55	0.6557	0.00235 ^c	0.00146 ^c	3.49 ^c
Cl₂	70.906	-100.98	-34.6	2.45	2.8982	1.460	0.206	462
CO	28.0104	-205.0	-191.5	0.97	1.1449	0.00437	0.00156	3.50
CO₂	44.0098	-	-78.5 ^d	1.52	1.7988	0.348	0.079	177
H₂	2.01594	-259.14	-252.5	0.07	0.0824	0.00019	0.00096	2.141
H₂S	34.0799	-85.5	-60.7	1.18	1.393	0.665	0.195	437
HCl	36.461	-114.8	-84.9	1.26	1.4903	82.300	22.572	50591
N₂	28.0134	-209.86	-195.8	0.97	1.145	0.003	0.001	2.328
N₂O	44.0128	-90.81	-88.46	1.52	1.799	0.255	.0580	130
NH₃	17.0306	-77.7	-33.3	0.59	0.6961	89.900	52.787	118314
NO	30.0061	-163.6	-151.8	1.04	1.2265	0.00983	0.00328	7.343
NO₂	46.0055	-11.2	21.2	1.59	1.8804	e	e	e
O₂	31.9988	-214.8	-183.0	1.10	1.3079	0.00700	0.00219	4.903
SiH₄	32.1179	-185	-111.8	1.11	1.3128	e	e	e
SO₂	64.0628	-72.7	-10.0	2.21	2.6185	22.800	3.559	7977

^a at 25 °C

^bmL gas/100 mL H₂O

^c at 17 °C

^d sublimes at 1 atm

^e reacts with water

Appendix B. Summary Table of Gas Preparation

Gas:	Method/ Chapter	Reagent:	Second Reagent	Disposal:	Wash Gas?	Chem13 Ref	Notes:
C_2H_2	In-Syringe Chap. 14	0.20 g CaC_2	5 mL water	Discharge out of doors	Yes	Apr 1997	4
C_2H_4	Thermal Chap. 21	3 mL conc H_2SO_4	3 mL anhydrous C_2H_5OH	Burn / discharge out of doors	Yes	Sep 1999	4
CH_4	Thermal Chapter 22	2 g $NaC_2H_3O_2$	2 g NaOH	OK to discharge indoors	Opt- ional	Apr 2000	4, 9
Cl_2	In-Syringe Chapter 16	0.22 g $Ca(OCl)_2$	5 mL 2 M HCl	Bubble through 1M NaOH(aq)	No	Oct 1997	1, 5, 10
CO	Thermal Chapter 20	8 drops H_2SO_4	8 drops HCOOH	Discard outdoors	No	Mar 1999	9
CO_2	In-Syringe Chapter 1	0.22 g $NaHCO_3$	5 mL vinegar or 1 M HCl	OK to discharge indoors	Yes	Nov 1996	
H_2	In-Syringe Chapter 3	0.05 g Mg turnings	5 mL 2 M HCl	OK to discharge indoors	Yes	Dec 1996	
H_2S	In-Syringe Chapter 26	0.22 g ZnS	3 - 5 mL 6 M HCl	Bubble through 1M NaOH(aq)	Yes	May 1997	1, 6
HCl	Thermal Chapter 19	1.4 g $NaHSO_4$	0.7 g NaCl(s)	Bubble through water	No	Apr 1998	9
N_2	In-Syringe Chapter 17	0.20 g HSO_3NH_2 (sulfamic acid)	5 mL of 0.5 M $NaNO_2$	OK to discharge indoors	Yes	Jan 1997	7
N_2O	Both Chapter 23	5 mL 6 M HNO_3	1 g NH_4NO_3 and 0.05 g NaCl	OK to discharge indoors	No	Jan 2002	
NH_3	In-Syringe Chapter 13	none	3 mL strong NH_3 solution	Bubble through water	No	Mar 1997	9
NO	In-Syringe Chapter 12	0.25 g $NaNO_2^2$	3 - 5 mL acidic $FeSO_4$ solution ³	Bubble through 1M NaOH(aq)	Yes	Feb 1997	
NO_2	In-Syringe Chapter 12	$NO(g)$	$O_2(g)$	Bubble through 1M NaOH(aq)	No	Feb 1997	
O_2	In-Syringe Chapter 4	0.10 g KI crystals	5 mL 6% H_2O_2	OK to discharge indoors	Yes	Jan 1997	9
SO_2	In-Syringe Chapter 15	1.7 g $NaHSO_3^2$	5 mL of 6 M HCl	Bubble through 1M NaOH(aq)	No	Sep 1997	8, 9
SiH_4	In-Syringe Chapter 25	0.2 g Mg_2Si	10 ml 1 M HCl	Burn	No	Jan 2000	1

Notes:

¹ Appropriate for classroom demo by teacher; not recommended for use by first year high school students.

² This quantity exceeds the calculated value but gives good results (generates 50 - 60 mL gas)

³ To prepare enough acidic ferrous sulfate solution to perform the syringe reaction ten times, add 13.5 g $FeSO_4 \cdot 7 H_2O$ to 28 mL distilled water. Add 12 mL 6 M H_2SO_4 and stir for a few minutes until all of the solid is dissolved.

⁴ Flammable gas

⁵ Use large vial cap; poisonous gas; heavier than air; quantities listed generate only 30 mL gas

⁶ Reacts slowly, stench, poisonous

⁷ Reaction is slow; trace NO_2 is initially formed but soon disappears

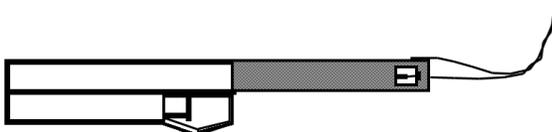
⁸ Use large vial cap

⁹ Also can be prepared by the microwave method; see Chapter 24.

¹⁰ Alternatively, use 1.0 mL 6 M HCl(aq) as the first reagent and 3 mL household bleach as the second reagent

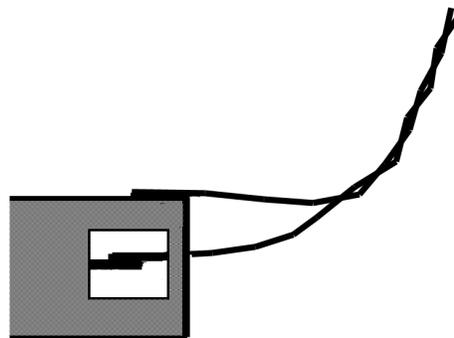
Appendix C. Construction of a Sparker from an Empty Piezoelectric Lighter and Launching Pipet Rockets

An extremely useful sparker used for initiating reactions of gas mixtures can be constructed from an empty (out of fuel) piezoelectric igniter such as an Aim-&-Flame. The assembled device is shown here.



How to build a sparker

Make sure the piezoelectric lighter is completely out of fuel. Use the pinch-nosed pliers to remove the end cap from the lighter. Locate the terminal in the center of the barrel of the piezo lighter. Click the lighter to observe a spark arcing between the terminal and the chassis of the lighter. Remove 5 mm of insulation from one end of each of the two 8 cm lengths of 24-gauge wire. Solder one of the bare wire leads to the chassis and the other to the piezo electrode as shown at right. The cavity around the piezo electrode is then filled with hot glue which adds stability and provides insulation from arcing. Carefully twist the two wires together as shown. Simultaneously cut both wires so they end in close proximity. Do NOT remove insulation from the arcing leads; the spark will easily jump and the insulation prevents the leads from making contact. Place a band of electrician's tape around the chassis/wire lead solder joint to complete the project. Click the lighter to check for spark.

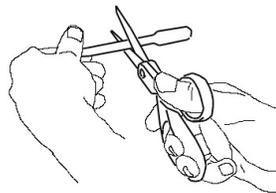


A step-by-step color photo sequence on filling and launching rockets is available at our website and summarized here:

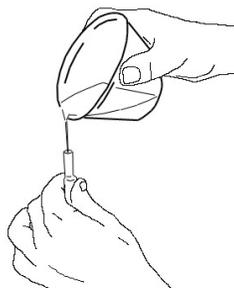
Instructions for filling and launching rockets

Prepare a mixture of the two gases via the method of syringe-to-syringe transfer as described in Chapter 1, page14. The procedure for filling rockets with gas is:

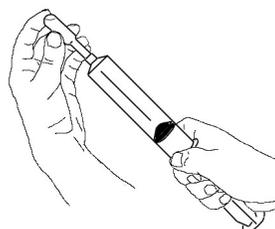
1. Cut the end off of a pipet leaving at about 3 cm of stem attached to the bulb.



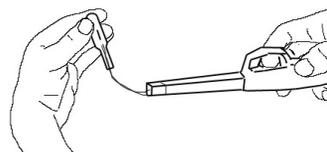
2. Completely fill the “rocket” with water.



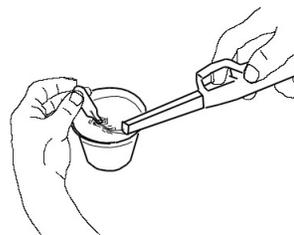
3. Slip the water-filled “pipet rocket” into the syringe fitting. Slowly displace the water with the hydrogen-oxygen mixture. The water will dribble out onto the bench top or floor. Holding the syringe and bulb at a 45° angle works best.



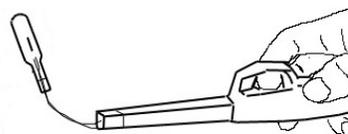
4. Slip the pipet rocket over the wire leads of the piezoelectric sparker. Remember that hydrogen is lighter-than-air! Never tip the gas-filled rocket open-end up — the gas will escape.



5. Draw some water into the stem — without this, the rocket will not fly far when “launched”. The ends of the wire leads of the piezoelectric lighter must be above the water in the gas-filled region of the rocket.



6. With some water in the stem, launch the rocket by triggering a spark. **DO NOT** aim the rocket at anyone! If the water leaks out of the stem while positioning the rocket over the wires, immediately fill the stem again by holding the wires and rocket in a cup of water and drawing a very small amount of water into the stem. The rocket should fly up to 10 m! Water must remain in the stem.



Appendix D. Preparation of Stock Solutions

1. Acids

A. Hydrochloric Acid, HCl(aq)

- ❖ 6 M HCl(aq). Concentrated HCl is 12 M. To prepare 500 mL 6 M HCl, slowly add 250 mL concentrated HCl to 250 mL distilled water. Store in a polyethylene bottle for safety. Mix thoroughly. Label the solution "6 M HCl"
- ❖ 3 M HCl(aq). To prepare 500 mL 3 M HCl, add 250 mL 6 M HCl to 250 mL distilled water. Store in a polyethylene bottle for safety. Mix thoroughly. Label the solution "3 M HCl"
- ❖ 2 M HCl(aq). To prepare 500 mL 2 M HCl, add 167 mL 6 M HCl to 333 mL distilled water. Store in a polyethylene bottle for safety. Mix thoroughly. Label the solution "2 M HCl"
- ❖ 1 M HCl(aq). To prepare 500 mL 1 M HCl, add 83 mL 6 M HCl to 417 mL distilled water. Store in a polyethylene bottle for safety. Mix thoroughly. Label the solution "1 M HCl"

B. Sulfuric Acid, H₂SO₄(aq)

- ❖ 6 M H₂SO₄(aq). Concentrated H₂SO₄ is 18.0 M. To prepare 500 mL 6 M H₂SO₄, slowly add 167 mL concentrated H₂SO₄ to distilled water to make a final volume of 500.0 mL. CAUTION! Solution will become very hot! Allow to cool and then store in a polyethylene bottle for safety. Mix thoroughly. Label the solution "6 M H₂SO₄"
- ❖ 5 M H₂SO₄(aq). To prepare 100 mL 5 M H₂SO₄, add 83 mL 6 M H₂SO₄(aq) to 17 mL distilled water to make a final volume of 100.0 mL. Store in a polyethylene bottle for safety. Mix thoroughly. Label the solution "5 M H₂SO₄"

C. Nitric Acid, HNO₃(aq)

- ❖ 6 M HNO₃(aq). Concentrated HNO₃ is 15.8 M. To prepare 100 mL 6 M HNO₃(aq), slowly add 38 mL concentrated HNO₃ to distilled water to make a final volume of 100.0 mL. CAUTION! Solution will become hot. Store in a polyethylene bottle for safety. Mix thoroughly. Label the solution "6 M HNO₃(aq)"

2. Bases

- ❖ 6 M NaOH(aq). To prepare 500 mL 6 M HCl, slowly dissolve 120 g NaOH(s) in water to make 500 mL solution. The solution will become exceedingly hot as the NaOH dissolves. Mix thoroughly. Label the solution "6 M NaOH"

- ❖ 3 M NaOH(aq). To prepare 500 mL 3 M NaOH, add 250 mL 6 M NaOH to 250 mL distilled water. Store in a polyethylene bottle for safety. Mix thoroughly. Label the solution “3 M NaOH”
- ❖ 1 M NaOH(aq). To prepare 500 mL 1 M NaOH, add 83 mL 6 M NaOH to 417 mL distilled water. Store in a polyethylene bottle for safety. Mix thoroughly. Label the solution “1 M NaOH”

3. 3% Dish soap solution

Add 30 g Dawn blue dish soap to a 1-L polyethylene bottle and fill with distilled water. Mix thoroughly. Label the solution “3% Dish Soap”

4. Dilute Potassium Permanganate Solution

Add a few small crystals of potassium permanganate (0.004 g) to a clean glass bottle containing 100 mL distilled water. Mix thoroughly. Store tightly covered. Label the solution “Dilute KMnO_4 ”. After a prolonged period of time, the solution will turn brown and should be discarded. Treat with one drop of 1 M sodium bisulfite solution and discard down the drain.

5. 1 M Sodium Bisulfite Solution

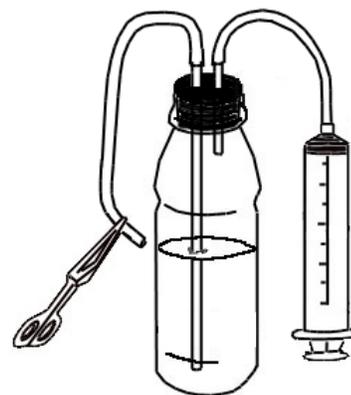
Add 10.4 g sodium bisulfite (NaHSO_3) to a clean glass bottle and add distilled water to make 100 mL solution. Mix thoroughly. Store tightly covered. Label the solution “1 M NaHSO_3 ”

6. 6% Hydrogen Peroxide Solution

To prepare 100 mL 6% H_2O_2 , add 20 mL 30% H_2O_2 (aq) to 80 mL distilled water to make a final volume of 100.0 mL. Store in a special polyethylene bottle with venting cap (use an old H_2O_2 bottle). Mix thoroughly. Label the solution “6 % H_2O_2 ”

7. Limewater Solution

Limewater is a clear colorless saturated Ca(OH)_2 (aq) prepared by mixing 1.5 g Ca(OH)_2 (s) per liter of water. Stir or shake vigorously and allow the solid to settle overnight. When using limewater, decant carefully to avoid transferring any solid or suspended Ca(OH)_2 (aq). In the drawing, limewater is stored in a 1-L plastic water or soft drink bottle. A syringe is used to apply pressure in controlled increments, say 5-mL at a time. The limewater



used in this device is carefully decanted from a 2.5-L acid bottle where it is prepared. The contents of this simple dispenser remain clear for long periods of time. Replace the rubber stopper with the original bottle cap for long-term storage.

8. Bromine water

Flinn Scientific has developed a convenient method for the preparation of bromine water from sodium bromide, household bleach and 1 M hydrochloric acid. Their method does not require the use of liquid bromine. In order to make 50-mL bromine water, you will need:

- Glass bottle with screw cap, label the bottle "Br₂(aq)"
- 1.1 g sodium bromide, NaBr
- 10.7 mL 1 M HCl
- 7.6 mL sodium hypochlorite (household laundry bleach)
- 32 mL water (preferably distilled water)

Dissolve the NaBr in the HCl and transfer to the labeled bottle. Add the sodium hypochlorite to the bottle and swirl to mix the reagents. Keep the bottle capped when not in use.

Disposal of bromine water

Excess bromine water can be "neutralized" to harmless bromide by reacting it with a few drops of 1 M sodium bisulfite solution or 1 M sodium thiosulfate solution. Add either of these solutions with stirring until the red color of the bromine water becomes colorless.

9. Universal indicator solution at pH 8

Numerous experiments require a slightly basic universal indicator solution. Prepare a solution by mixing 200 mL distilled water plus 20 mL universal indicator solution. Raise the pH to 8 by bubbling through the solution a pipetful of gaseous ammonia taken from the vapors above a solution of concentrated ammonium hydroxide solution.

10. Red cabbage juice indicator solution

Red cabbage juice solutions provide a practical alternative to universal indicator solution. Preparation of a red cabbage juice solution may be accomplished in a variety of ways. Perhaps the easiest is to buy a can of red cabbage and drain the juice through a kitchen sieve. Fresh red cabbage can be used. Cook a small wedge with 500 mL water until the water becomes red in color. Allow to cool and drain through a kitchen sieve. Excess solution can be frozen for later if not used within a few days. Store tightly capped in a plastic bottle.

Appendix E Ordering Information

Chem13 News

Chem13 News is an informal magazine published for teachers of introductory chemistry courses, particularly high school teachers. Issues, each with 16 pages or more of teaching ideas, articles, experiments, puzzles, appear monthly, September through May. For information on current subscription rates, contact **Chem13 News**, Department of Chemistry, University of Waterloo, Waterloo, ON, N2L 3G1, CANADA, Telephone: 519-888-4567, ext. 2505 Fax: 519-888-9168 e-Mail: kjackson@waterloo.ca

Educational Innovations, website: <http://www.teachersource.com/>

Syringes, etc.

syringes, 60-mL plastic syringe with a LuerLOK fitting (Educational Innovations GAS-140)

syringe cap fittings, latex LuerLOK (Educational Innovations GAS-160)

plastic vial caps (Educational Innovations GAS-180)

silicone oil for \$5.95. Part #GAS-150

metal locking hemostat (Educational Innovations GAS-200)

Latex tubing, 1/8-inch (3.175 mm) ID, 5 ft (1.75 m) lengths (Educational Innovations GAS-220)

Flinn Scientific, website: <http://www.flinnsci.com/>

Chemistry of Gases, A Microscale Approach, Mattson, Anderson, Schwennsen, (Flinn AP4849)

Flinn has also published our 350-page book designed for use by high school students and their teachers. Each experiment contains a significant pedagogical component. The book covers the gases CO₂, H₂, O₂, N₂, NO_x, NH₃, C₂H₂, SO₂, H₂S, and Cl₂.

Syringes and Gas Kits

Gas Kit containing two syringes, latex caps, latex tubing and more (Flinn AP 4850)

Latex tubing, 1/8-inch (3.175 mm) ID, 15 cm lengths (Flinn AP2076)

syringe cap fittings, latex LuerLOK (Flinn AP8958)

syringes, 60-mL plastic syringe with a LuerLOK fitting (Flinn AP8754)

Fisher Scientific

Most of the items needed to perform the experiments in this book can be found in Fisher Scientific's catalog. Their prices are excellent, however, usually it is necessary to purchase large quantities of materials.

Aldrich

magnesium silicide, Mg_2Si (Aldrich 343196) (for Preparation of Silanes)

platinum sponge (Aldrich 26715-5)

Alltech, telephone: 847-948-8600

Gas chromatography column: Porapak N 80/100, 6-ft (180 cm), inside diameter = 0.085 inches (2.2 mm); helium flow: 30-mL/minute; room temperature; Alltech Part Number 2716

Appendix F. Gas Equipment

Microscale Gas Chemistry Kits

Each pair of students will need certain equipment in order to prepare gases and perform experiments with the gases. We recommend organizing this equipment in 8-cup plastic food storage containers. Each kit should contain:

- ❖ 8-cup plastic food storage container such as Gladware
- ❖ two 60-mL plastic syringes with a LuerLOK fitting
- ❖ two Latex LuerLOK syringe caps
- ❖ two plastic vial caps
- ❖ one 15 cm length of Latex tubing
- ❖ one 3 cm length of Latex tubing
- ❖ one small bottle of silicone oil
- ❖ one plastic pipet
- ❖ one clear plastic beverage cup (250 mL/9 oz)
- ❖ one small plastic weighing dish
- ❖ one small test tube (12 x 100 mm)
- ❖ one medium test tube (18 x 150 mm)
- ❖ one birthday candle

All of this will fit into the food storage container. In addition, each pair of students will need a wide-mouth beverage bottle for draining and supporting their syringes. Ordering information for kit materials is given at the end of this chapter and in Appendix E.

Equipment for generating gases by the Thermal Method (Chapter 19)

Each pair of students will need certain equipment in order to prepare gases and perform experiments with the gases. We recommend organizing this equipment in 8-cup plastic food storage containers. Each kit should contain:

- ❖ 18 x 150 mm test tube (in addition to the one in the kit)
- ❖ two-hole #1 stopper fitted with two short lengths (2 cm) of glass tubing
- ❖ two pieces of tubing, 1/8-inch (3.175 mm) ID, 5 cm length
- ❖ pinch clamp or hemostat
- ❖ ring stand and a suitable clamp to hold test tube
- ❖ wooden dowel or aluminum rod, 1 cm diameter x 15 cm length with suitable clamp to hold it to the ring stand
- ❖ three heavy duty rubber bands
- ❖ small Bunsen burner
- ❖ matches or a lighter
- ❖ marker pen

Appendix F. Answers to the Questions

Chapter 1. Getting Started: Generating Carbon Dioxide

- (a) NaHCO_3 ; (b) $\text{HC}_2\text{H}_3\text{O}_2$; and (c) CO_2
- (a) $\text{NaHCO}_3(\text{s})$; (b) $\text{HC}_2\text{H}_3\text{O}_2(\text{aq})$; (c) $\text{CO}_2(\text{g})$
- It is actually the syringe containing the gas that must be washed in order to remove all traces of chemicals that were used to make the gas. After that has happened, the syringe contains our desired gas (CO_2), a little bit of air and a few droplets of water.
- Use only the specified amounts of reagents to avoid making more gas than the syringe can hold.
- The syringe should be upright when removing the cap so that the liquid will not spray out onto your hand, clothing, laboratory book, etc.
- Vigorously shaking the syringe mixes the chemicals. This becomes even more important if one of the reagents is a gas.
- If some of the solid reagent spills out of the vial cap as it is being lowered, that amount of reagent will be removed from the syringe with the water used to lower the vial cap.
- The molar mass of carbon dioxide is 44.01 g/mol.
- $\text{NaHCO}_3(\text{s}) + \text{HC}_2\text{H}_3\text{O}_2(\text{aq}) \longrightarrow \text{CO}_2(\text{g}) + \text{H}_2\text{O}(\text{l}) + \text{NaC}_2\text{H}_3\text{O}_2(\text{aq})$
- 0.22 g of NaHCO_3 corresponds to 0.0026 mol NaHCO_3 and 5.0 mL of 1.0 M HCl corresponds to 0.0050 mol HCl. Because they react 1:1, NaHCO_3 is the limiting reactant.
- 64 mL at 25 °C and 1 atm pressure
- The density of carbon dioxide at 25 °C and 1 atm pressure is 1.80 g/L.
- Using the “average molar mass of air”, 28.964 g/mol, the density of air is 1.18 g/L. The ratio of densities, $d_{\text{carbon dioxide}}/d_{\text{air}}$ is 1.52:1.00. Note that this is true under all conditions of temperature and pressure and can be determined as a simple ratio of their molar masses.

Chapter 2. Experiments with Carbon Dioxide

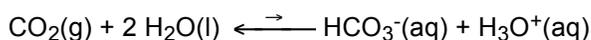
Chapter 2, Experiment 1. Traditional limewater test for carbon dioxide

- A precipitate is an insoluble solid substance that is formed from an aqueous solution. Usually, precipitates are noticed as a cloudiness in the solution or as suspended particles. Eventually they settle to the bottom.
- Limewater is a saturated solution of calcium hydroxide, $\text{Ca}(\text{OH})_2(\text{aq})$
- Carbon dioxide is $\text{CO}_2(\text{g})$
- The symbol “(aq)” stands for aqueous, “(g)” is gas “(s)” is solid (including both solids out of a bottle and precipitates from a reaction) and “(l)” stand for liquid.
- Yes, most precipitates settle if allowed to stand.
- Calcium carbonate is a white solid.
- The carbonate ion has the formula CO_3^{2-} .
- A good test for carbon dioxide is the limewater test. Shake a sample of gas, suspected of being carbon dioxide with limewater. The formation of a white precipitate of calcium carbonate confirms carbon dioxide.

9. The traditional test for the carbonate ion is also the limewater test.
10. Calcium carbonate is an ionic substance.
11. Aqueous calcium hydroxide and carbon dioxide gas produced solid calcium carbonate and water.

Chapter 2, Experiment 2. Acidity of carbon dioxide

1. Carbonated beverages contain carbon dioxide which makes them slightly acidic.
2. Acids such as vinegar cause the pH to drop to less than 7.0 so that universal indicator will be yellow, orange or red.
3. Your friend used cabbage indicator if an acidic solution turned the indicator purple. Refer to the indicator color chart.
4. Ammonia is much more soluble than carbon dioxide. You may suspected that when all that was necessary to dissolve ammonia is to discharge some of the gas near the surface of the water.
5. The indicator solution eventually turns the color associated with acid because there is more carbon dioxide than ammonia.
6. The pH of the distilled water in a laboratory tends to run slightly less than 7.0 because all water tends to dissolve carbon dioxide naturally present in air. While this doesn't change the pH of tap water much, distilled water is free of all other substances so that the carbon dioxide causes a noticeable change in pH.
7. Carbon dioxide forms $\text{CO}_2(\text{aq})$ in water. Some of it reacts with water forming the bicarbonate ion, $\text{HCO}_3^-(\text{aq})$ according to the equilibrium:



Chapter 2, Experiment 3. Carbon dioxide extinguishes fires

1. A candle cannot burn in carbon dioxide. Carbon dioxide does not burn.
2. A candle can burn in oxygen. In fact, it requires oxygen (air contains 21% oxygen).
3. The candle was extinguished. Carbon dioxide is used in certain types of fire extinguishers.
4. We released the carbon dioxide in the bottom of the cup because the syringe only holds 60 mL of carbon dioxide and the cup is capable of holding 3 or 4 times that volume.
5. We used a short candle so that the flame would be towards the bottom of the cup where the concentration of carbon dioxide is greatest.
6. Carbon dioxide has a greater density than air. See Questions 12 and 13 from Chapter 1, page 16. Air is mostly nitrogen (78%) and oxygen (21%). Nitrogen, N_2 , has a molar mass of 28 g/mol and oxygen, O_2 , is 32 g/mol. Both are less than that of CO_2 , 44 g/mol.
7. The heat from the candle creates air currents whereby the hot air rises above the candle. Cool air flows into the cup along the inside edge of the cup. This current flow of gases would "pump" the carbon dioxide out of the cup as part of the general air flowing before it could extinguish the fire if the carbon dioxide were discharged too slowly.
8. Many possibilities.

Chapter 2, Experiment 4. Carbon dioxide and aqueous sodium hydroxide react

1. NaOH and Na_2CO_3

2. Carbon dioxide reacted with the sodium hydroxide solution.
3. Shaking the syringe mixes the carbon dioxide gas and the aqueous sodium hydroxide to mix together better and this speeds up the reaction.
4. The carbon dioxide came into contact with the aqueous solution. Some carbon dioxide dissolves in the water and reacts with the sodium hydroxide. Then more carbon dioxide dissolves in the water and the process continues. As the amount of carbon dioxide is depleted, the plunger is pushed inward by the fact that the pressure outside the syringe is greater than that inside the syringe.
5. Solutions of sodium hydroxide react with $\text{CO}_2(\text{g})$ from the air. The reaction is slow because the amount of carbon dioxide in air is small, but over time, uncapped bottles of most bases will react to a noticeable extent. For example, if a bottle of $\text{NaOH}(\text{aq})$ had been determined to be 0.1122 M a year ago, most chemists would not trust that concentration any longer. Most likely, the actual concentration is now lower than that.
6. Sodium carbonate is an ionic compound.
7. Aqueous sodium hydroxide and carbon dioxide gas produce aqueous sodium carbonate and water.

Chapter 2, Experiment 5. Carbon dioxide/carbonic acid equilibrium

1. We used ice because carbon dioxide is more soluble in cold water.
2. Carbon dioxide is more soluble under higher pressures. Think of a bottle of a carbonated beverage; the contents are under pressure. When the pressure is released by opening the cap, some of the carbon dioxide escapes.
3. Carbonated beverages are the familiar product of carbon dioxide. It was discovered in 1772 by Joseph Priestley.
4. If an opened bottle or bottle of soda pop if it is allowed to warm up, the carbon dioxide escapes and the beverage becomes "flat".
5. If a can of soda pop is tapped, carbon dioxide bubbles will be formed.
6. Use a series of syringes, each filled with carbon dioxide and each prepared in exactly the same way so that each initially contained the same amount of carbon dioxide. Remove the caps of all the syringes. Test one syringe for carbon dioxide content by reaction with 6 M NaOH as per Experiment 4. Test subsequent syringes after 10 minutes, 20 minutes and 30 minutes.

Chapter 3. Experiments with Hydrogen

Preparation of Hydrogen

1. The formula for hydrogen gas is $\text{H}_2(\text{g})$.
2. You should hold the syringe with the opening down when the cap is off to prevent loss of hydrogen due to effusion. Hydrogen is much lighter than air and will quickly escape if the open syringe is directed upward. Nevertheless, you should remove the syringe cap with the cap directed upward to avoid discharging liquid chemicals on hands, etc. Then quickly rotate the syringe so that the open end is directed downward.
3. The volume of hydrogen you obtained will vary, but is usually 50+ mL. Try to read the syringe to the nearest 0.5 mL.
4. The reaction exothermic — it gives off heat.
5. You can read the syringe to within +/- 0.2 mL

- The funny shape of the plunger deal is accounted for by the fact that the volume that it occupies is the same as the volume at the
- Divide mass by 24.31 g/mol. If you used 0.050 g Mg as directed, your answer should be 0.0021 mol magnesium.
- Using 5 mL of 2.0 M HCl corresponds to 0.010 mol HCl.
- $\text{Mg(s)} + 2 \text{HCl(aq)} \longrightarrow \text{H}_2(\text{g}) + \text{MgCl}_2(\text{aq})$
- Mg(s) is the limiting reagent
- Using the ideal gas law and 0.050 g Mg as the limiting reagent, the volume of gas is expected at 25 °C and standard pressure is 50.3 mL.

Chapter 3, Experiment 1. Traditional test for hydrogen

- Helium is also much lighter than air.
- The purpose of collecting the gas under water by water displacement is to collect pure hydrogen free of undue contamination by air.
- The atomic symbol for oxygen is O and the formula of the molecular substance it forms is O₂.
- The test tube containing hydrogen gas was stored upside down in the water to prevent diffusion of the gas. A rubber stopper would work as well if it made a tight seal.
- In order to distinguish between a test tube of hydrogen and carbon dioxide, the limewater test would be easiest. The traditional test for hydrogen would work too.
- The hydrogen-air mixture is louder than that of pure hydrogen because the oxygen from air is already mixed in with the hydrogen. With pure hydrogen, the closest oxygen is in the air just outside the mouth of the test tube.
- Your sketch should show the flow of hydrogen gas upward and the flow of air into the test tube.
- H₂O(g) is formed when H₂(g) is ignited in air.
- $2 \text{H}_2(\text{g}) + \text{O}_2(\text{g}) \longrightarrow 2 \text{H}_2\text{O}(\text{g})$ Note: H₂O is probably a gas at the high temperature associated with the explosion.

Chapter 3, Experiment 2. Hydrogen forms explosive mixtures with air

- Part 2 is much louder because of the oxygen present from the air that has been added.
- The purpose of the soap solution is to trap the gas until the we are ready to trigger the explosion.
- Extreme care must be exercised when working with hydrogen gas because it forms explosive mixtures with air as the name of this experiment implies.
- The two major gases found in the air are nitrogen and oxygen. It is the oxygen that reacts with hydrogen in this experiment.
- If one bubble of hydrogen and one bubble of air were side-by-side and ignited, the bang would be not as loud than if the same amount of gas were contained in a single bubble because the two gases are not mixed together as well as they would be in a single bubble.
- The explosion would be louder if oxygen were used instead of air.
- $2 \text{H}_2(\text{g}) + \text{O}_2(\text{g}) \longrightarrow 2 \text{H}_2\text{O}(\text{g})$
- Create various mixtures of hydrogen/air in a syringe. Transfer each mixture to a bubble, trying to maintain similar sizes to the bubbles. Ignite the bubbles and rate the results.

Chapter 3, Experiment 3. Reversible conversion of copper metal and copper(II) oxide

1. (a) CuO(s); (b) Cu(s); (c) H₂(g); (d) H₂O(l)
2. The most obvious indication should have been the color change. In the second part, you may have noticed the formation of water droplets.
3. This reaction could be repeated over and over without consuming the copper.
4. You would know for sure that the conversion of copper to copper(II) oxide and then copper(II) oxide to copper is reversible if you repeated it.
5. Many reactions need heat to get them started. This is such a reaction. The molecules do not have enough kinetic energy to react at room temperature.
6. (a) hydrogen is a molecular compound; (b) copper is a metal; (c) copper(II) oxide is an ionic compound ; (d) water is a molecular compound; and (e) silicon dioxide network covalent compound
7. Of the four fundamental types of substances, only molecular compounds can be either solid, liquid or gas under standard conditions. The rest are almost always solids.
8. In the oxidation of copper with air ($2 \text{ Cu(s)} + \text{O}_2(\text{g}) \longrightarrow 2 \text{ CuO(s)}$), the oxidizing agent is oxygen and the reducing agent is copper. In the second reaction ($2 \text{ CuO(s)} + \text{H}_2(\text{g}) \longrightarrow 2 \text{ Cu(s)} + \text{H}_2\text{O(l)}$), the oxidizing agent is CuO and the reducing agent is H₂.

Chapter 3, Experiment 4. Reduction of iron(III) oxide with hydrogen

1. The formula contains two iron atoms for every three oxygen atoms.
2. You may have observe a color change as the iron darkens. Also, you used a magnet to test the solid in the pipet before and after the reaction and noticed quite different results. Also, you may have noticed the formation of water droplets.
3. The magnet tell us that a reaction has taken place because the product is attracted to the magnet, while the reactant, Fe₂O₃, was not. This test only works in rather rare circumstances — either the reactant or product must be ferromagnetic, the property of being attracted to a magnet.
4. Fe₂O₃ is not ferromagnetic; Fe is ferromagnetic.
5. Iron(III) oxide and gaseous hydrogen react at high temperature to produce elemental iron and water vapor.
6. Your rate of hydrogen flow depends on how you did the experiment. Suppose that you passed 45 mL of hydrogen through the pipet in 55 seconds. Convert 55 seconds to minutes, $55 / 60 = 0.92 \text{ min}$. Rate = $45 \text{ mL} / 0.92 \text{ min} = 49 \text{ mL/min}$.
7. Many reactions need heat to get them started. This is such a reaction. The molecules do not have enough kinetic energy to react at room temperature.
8. 0.25 g Fe₂O₃ corresponds to 0.0016 mol Fe₂O₃. 40 mL H₂ at room temperature and standard pressure corresponds to 0.0016 mol H₂. Referring to the reaction ($\text{Fe}_2\text{O}_3(\text{s}) + 3 \text{ H}_2(\text{g}) \longrightarrow 2 \text{ Fe(s)} + 3 \text{ H}_2\text{O(g)}$), we need three moles of H₂ for every mole of Fe₂O₃, thus H₂(g) is the limiting reagent.
9. The bright white glow indicates an exothermic reaction.

Chapter 3, Experiment 5. Effusion of hydrogen is faster than air

1. Hydrogen is effusing through the balloon faster than air can effuse dur to its smaller molecular size. As a result, the quantity of gas inside the syringe is constantly decreasing.

Because pressure is related to moles of gas at constant temperature and volume, the pressure inside the syringe would decrease with time when compared to the external pressure. The rubber balloon, however, stretches inward to minimize the pressure difference.

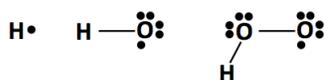
2. Effusion refers to random molecular motion involving an orifice. Diffusion refers to random molecular motion in all directions.
3. In theory, carbon dioxide molecules should move through the rubber balloon more slowly than air molecules. This should cause the balloon to bulge outward rather than inward. In the actual experiment, the results are noticed fairly quickly because hydrogen passes relatively quickly through the balloon. Both air and carbon dioxide would be much slower.

Chapter 3, Experiment 6. Hydrogen burns with a gentle flame

1. Hydrogen and air are not allowed to mix so there is no explosion
2. $2 \text{H}_2(\text{g}) + \text{O}_2(\text{g}) \longrightarrow 2 \text{H}_2\text{O}(\text{g})$
3. In this case, if air were to get inside the syringe, the explosion would take place in a closed container and cause a potentially dangerous explosion. With the soap bubbles, the container (the bubbles) harmlessly bursts as the gases explode.
4. A flame is associated with a reaction that is slow enough for our eyes to “catch” the action. When hydrogen/air mixtures explode, the reaction is over in an instant — often faster than the eye can see.
5. If you were to hold a piece of glass above the flame, water vapor would condense.

Chapter 3, Experiment 7. Incomplete combustion of hydrogen

1. Normally, hydrocarbons burn to produce carbon dioxide and water. When pure hydrogen burns, the normal product is only water.
2. The yellow color is from sodium in the glass.
- 3.



Chapter 3, Experiment 8. Disappearing/reappearing candle flame

1. The explanation is that inside the syringe there is no air, only hydrogen and pure hydrogen does not burn. The wick continues to glow for a second or two and is enough to re-ignite the hydrogen-air mixture as the syringe is lifted.
3. Not all flames are visible. Hydrogen, as well as alcohols burn without a visible flame.

Chapter 3, Experiment 9. Calcium and calcium hydride produce hydrogen in reactions with water

1. Calcium metal looks very much like most common metals. It is shiny and silvery.
2. Calcium hydride looks like gray chunks or small rocks — like marble chips.
3. The formula of calcium hydride is $\text{CaH}_2(\text{s})$. Calcium hydride is ionic.
4. Both reactions are vigorous and evolve hydrogen gas. They are over within a few seconds.
5. Both are oxidation-reduction reactions.
6. Do the traditional test for hydrogen, Experiment 1, on the gases collected.

Chapter 3, Experiment 10. Deuterium isotope effects

1. The word isotope refers to the fact that some elements exist in more than one form, differing only by the number of neutrons. Hydrogen exists in three different isotopic forms; the first has one proton and zero neutrons; the second one has one proton and one neutron, and the third one has one proton and two neutrons. It is impossible to predict the existence of isotopes.
2. H is ${}^1_1\text{H}$, and D is ${}^2_1\text{H}$.
3. Tritium would be the slowest of the three isotopes because of its largest molar mass. The ranked order of rate of reaction from slowest to fastest. ${}^3_1\text{H} < {}^2_1\text{H} < {}^1_1\text{H}$.

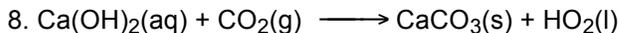
Chapter 4. Experiments with Oxygen

Preparation of Oxygen

1. The formula for oxygen gas is $\text{O}_2(\text{g})$.
2. (a) $\text{KI}(\text{s})$; (b) $\text{H}_2\text{O}_2(\text{aq})$; and (c) $\text{H}_2\text{O}(\text{l})$
3. The reaction should have sped up when you tapped the syringe while holding the plunger slightly outward. This works because the agitation causes the molecules to collide with the surface with more kinetic energy and often that is enough to initiate the formation of a bubble. The same thing happens when you tap a bottle of carbonated beverage.
4. If the potassium iodide were to spill out of the cap before you were able to draw up the hydrogen peroxide, it would be impossible to draw up the $\text{H}_2\text{O}_2(\text{aq})$. As the H_2O_2 was being drawn into the syringe, it would react with the KI outside of the cap, form $\text{O}_2(\text{g})$ and that would push the $\text{H}_2\text{O}_2(\text{aq})$ back out of the syringe.
5. $2 \text{H}_2\text{O}_2(\text{aq}) \longrightarrow 2 \text{H}_2\text{O}(\text{l}) + \text{O}_2(\text{g})$
6. A catalyst is usually a chemical that speeds up a chemical reaction by providing an alternative reaction pathway that is faster.
7. 0.3 g of H_2O_2 is present in 5 mL of 6% $\text{H}_2\text{O}_2(\text{aq})$. This converts to 0.0088 mol H_2O_2 .
8. From the reaction stoichiometry, we expect to make 0.0044 mol $\text{O}_2(\text{g})$. At 25 °C and standard pressure, this gas would have a volume of 108 mL.

Chapter 4, Experiment 1. Traditional test for oxygen

1. The glowing splint should have re-ignited when it came in contact with oxygen. Carbon dioxide would extinguish the glowing splint.
2. Flames are associated with rapid reactions. When wood is glowing and not openly burning, it is because there is not enough available fuel surface area or oxygen to support an open fire. Blowing on glowing wood causes it to burst into flames by providing more oxygen. Stirring the coals also works by creating more surface area or exposing previously unburned wood.
3. The glowing splint test would positively identify oxygen from the other two gases.
4. Carbon dioxide could be determined by the limewater test and hydrogen could be determined by the traditional test for hydrogen (popping sound as test tube of hydrogen is ignited)
5. An oxidation-reduction reaction occurred in the experiment.
6. Carbon dioxide is being produced when the wood splint burns?
7. The limewater test should have been positive.



9. Your sketch should show oxygen pushing the air out. Unlike with some liquids, there is not a sharp boundary between oxygen and air. The gas at the bottom of the test tube becomes enriched with oxygen, but some air undoubtedly remains unless a large amount of oxygen is used.

Chapter 4, Experiment 2. Oxygen supports combustion

1. The candle should burn more brightly in oxygen.
2. The candle should burn longer in oxygen.
3. The candle should burn hotter in oxygen.
4. The candle goes out when the level of oxygen falls to a level too low to support combustion.
5. The candle wax may have dripped a lot, indicating that the fire was unusually hot in the presence of oxygen.
6. The two products of combustion of all carbon-hydrogen-containing substances are carbon dioxide and water.
7. The limewater test confirmed the presence of carbon dioxide as one of the products.
8. The three components needed to sustain a fire are fuel, oxygen and heat.
9. Assuming that the test tube were completely filled with oxygen (100%), the ratio of oxygen content in the two test tubes is 100:21 or 4.8:1.

Chapter 4, Experiment 3. Dynamite Soap

2. The soap solution provides a film to encapsulate the gas mixture and prevent it from escaping into the air.
3. Extreme care must be exercised when working with hydrogen-oxygen mixtures because they will explode by spark or flame.
4. Mini-explosions are relatively safe because of the small amount of chemicals being used. Scaling reactions up in size risks creating a situation in which you are not able to control or escape what is happening. By analogy, compare a simple firecracker with a stick of dynamite.
6. Using air would result in a softer BANG.
7. It is common for explosive mixtures to need a flame or spark to get the reaction to go. The energy of activation is so high that the gases can be stored without having them react. Once the reaction is started, it is self-sustaining.
8. According to the law of combining volumes, 5 mL oxygen would be needed to react with 10 mL hydrogen.
9. Energy is released by this reaction.

Chapter 4, Experiment 4. Hydrogen-oxygen rockets

2. We filled the rocket with water to eliminate air present so that the rocket could be filled with the prescribed amount of hydrogen and oxygen.
3. A rocket filled with a mixture of hydrogen and oxygen would fly the furthest.
4. The explosion creates a force that pushes against the water plug in the stem. If the water is not present, most of the force is dissipated as the gases leave the rocket.
5. $2 \text{H}_2(\text{g}) + \text{O}_2(\text{g}) \longrightarrow 2 \text{H}_2\text{O}(\text{l})$

- A rocket that is mostly filled with hydrogen and some oxygen would fly farther.
- Rockets filled with hydrogen and air would fly , but not as far.
- The optimal ratio of hydrogen to oxygen is 2 H₂(g) to 1 O₂(g).
- The gases must be stored separately until they are to be reacted.

Chapter 4, Experiment 5. Steel wool burns in oxygen

- If you used coarse steel wool instead of the fine steel wool, a similar result would be obtained. The fine steel wool probably burns faster due to its greater surface area.
- $4 \text{ Fe(s)} + 3 \text{ O}_2\text{(g)} \longrightarrow 2 \text{ Fe}_2\text{O}_3\text{(s)}$

Chapter 4, Experiment 6. The Blue Bottle experiment

- Shaking the syringe increases the concentration of dissolved oxygen, O₂(g).
- When the solution in the syringe becomes colorless, the O₂(aq) has been depleted.

Chapter 4, Experiment 7. Oxygen makes the flame hotter

- The colors of the various parts of a flame varies with fuel used. The outer part of the flame, regardless of its color is the hottest region.
- Oxy-acetylene torches. Used for welding, is an example of a real-world application that takes advantage of oxygen's ability to increase the temperature of a flame.
- $\text{CH}_4\text{(g)} + 2 \text{ O}_2\text{(g)} \longrightarrow 2 \text{ H}_2\text{O(l)} + \text{CO}_2\text{(g)}$
- The two most common fragments that are found in the inner cone are CH₃(g) and H(g).

Chapter 4, Experiment 8. Mini-sponge shooter

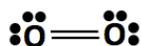
- After the alcohol has evaporated, it is mixed with air in the gas phase so the two reactants are in the form of a homogeneous mixture and are highly reactive. A liquid and a gas can only react at the surface of the liquid.
- Gases react more explosively than do liquids with air. The reaction between gasoline and air in a cars cylinder is an example of this. The carburetor or fuel injector converts liquid gasoline into the gas phase.
- Reactions are always faster at higher temperatures.

Chapter 4, Experiment 9. Chemiluminescence

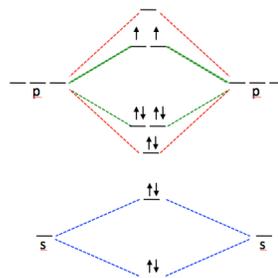
- The three ways in which a chemical reaction can give off energy are heat, light, and electricity.
- Other examples of chemiluminescence include fireflies, glow sticks, etc.
- Shaking the syringe brings more oxygen into solution and because oxygen is one of the reagents this makes the light brighter.
- The reaction eventually stops when one of the reagents is depleted.

Chapter 4, Experiment 10. Paramagnetism of liquid oxygen

- It is faintly blue.
-



- paramagnetic



Chapter 6 Mystery Gas

1. Several possible answers to this question. Perhaps the easiest would be to use the limewater test for carbon dioxide, the glowing splint test for oxygen and the “popping” test tube (traditional test) for hydrogen.
2. Performed the test for one particular gas on a different gas will lead to either no results (they do not react) or they may react in a way not expected. An example of the latter is what would happen if you performed the glowing splint test on hydrogen: you would hear the “pop” associated with hydrogen, but the splint would not re-ignite.
3. Oxygen.
4. “Dynamite Soap” is not as useful as other experiments for the identification of gases because it uses a mixture of two gases. Assuming you do not know the identity of any of the gas samples, this test would require three experiment to try every combination and then all that you would know is which pair is hydrogen or oxygen.
5. See the answer to Question 1.

Chapter 7 Percent Composition

6. Analyses of Tums tablets by these two methods should give similar results. The syringe method gives 36 – 39% CaCO_3 in Tums. The traditional method, gives 41 – 43% CaCO_3 .
7. Two factors contribute to the discrepancy: (1) CO_2 is soluble in water so that the syringe method gives slightly low results and (2) The Tums tablets splatter water as they react with considerable fizzing; some water is undoubtedly lost this way, giving results that are slightly high. Nevertheless, the results are similar.

Chapter 8 Percent Carbonated Beverages — Priestley’s Soda-water

Chapter 8, Experiment 1. Estimating the amount of carbon dioxide in a carbonated beverage

1. When you withdrew the plunger of the syringe containing the carbonated beverage, you should have seen bubbles forming.
5. If you used 20 mL as instructed, that would be 20 g water or 1.1 mol water. Suppose that the volume of carbon dioxide collected was 35 mL. Using the ideal gas law, this converts to 0.0014 mol carbon dioxide.

Chapter 8, Experiment 2. Carbonating water by Priestley’s method

1. Holding the syringe in a horizontal position maximizes the surface area of the water so that carbon dioxide can interact with a large surface.
2. Carbon dioxide dissolves slowly. Your sketch of the universal indicator color changes should account for the fact that the pH drops most near the surface where the carbon dioxide is actively dissolving. With time, the aqueous molecules migrate and the solution should become homogeneous in color.
3. $\text{CO}_2(\text{aq}) + \text{H}_2\text{O}(\text{l}) \rightleftharpoons \text{H}_2\text{CO}_3(\text{aq})$
4. The long arrow in the equilibrium expression points towards the primary component. In the equilibrium above, there is relatively little $\text{H}_2\text{CO}_3(\text{aq})$

Chapter 8, Experiment 3. Freezing carbonated beverages produces “snowy” ice

1. Water should look clear and glass-like. Carbonated water appears white and highly fractured — like frozen, compacted snow.

2. The plunger moves outward in the syringe with the sparkling water because as the solution freezes, carbon dioxide is released from the solution.
3. The sparkling water would taste flat after it were allowed to melt.
4. The carbonation could be returned to the liquid as was done in Experiment 2.
5. When ice forms, the intermolecular forces that form are hydrogen bonds. The intermolecular forces that exist between carbon dioxide and water are dipole forces.
6. Your graph should show that the solubility of a gas decreases as the temperature increases for the liquid region. If your graph shows frozen water, the solubility of carbon dioxide is considerably less in solid water as we learned in this experiment.

Chapter 8, Experiment 4. Spectacular crystallization of super-cooled carbonated beverages.

2. Salt lowers the freezing point of water, a colligative property.
4. The carbonated beverage would eventually freeze if the temperature drops low enough. People living in cold climates have perhaps accidentally experienced this if they left a carbonated beverage in a cold car in sub-freezing temperatures.
5. The frozen beverage has less dissolved carbon dioxide than the liquid form.
6. When ice forms, hydrogen bonds are formed, some of the strongest intermolecular forces. London dispersion intermolecular forces exist between carbon dioxide and water (ignoring the small amount that reacts with water to form carbonic acid.)
7. A nucleation site was created when the cap was removed and the ice formed quickly. Most likely a bubble started the process. The solution remains liquid while the cap was in place because no nucleation sites existed.

Chapter 8, Experiment 5. On the carbon dioxide/carbonic acid equilibrium

1. Using 20 mL sparkling water as instructed, there were 0.07 g of CO_2 dissolved, corresponding to 0.0016 moles.
2. Using 20 mL diluted vinegar as instructed, there were 0.0016 moles.
3. Using 5 drops as instructed, the volume of 3 M $\text{NaOH}(\text{aq})$ used was 0.25 mL. Given the concentration is 3 M $\text{NaOH}(\text{aq})$, the 5 drops contained 0.00075 mol NaOH , thus making it the limiting reagent.
4. The reaction between $\text{H}_2\text{CO}_3(\text{aq})$ and $\text{NaOH}(\text{aq})$ is fast. The slow step is the formation of $\text{H}_2\text{CO}_3(\text{aq})$ from $\text{CO}_2(\text{aq})$, Step 1. As Step 1 slowly shifts to the right, the $\text{H}_2\text{CO}_3(\text{aq})$ produced is immediately consumed by reaction with $\text{NaOH}(\text{aq})$, so for awhile, $\text{NaOH}(\text{aq})$ is effectively the reagent in excess. Within a few seconds, the slow step has caught up and, because $\text{NaOH}(\text{aq})$ is actually the limiting reagent, the final color of the solution is clear (acidic).

Chapter 9. Molar Mass of a Gas

1. Calculation II usually gives results that are closer to the actual values. Both give quite good results, however.
2. The largest source of error is that results depend on very small differences in mass. Drops of water, picking up the wrong syringe cap, forgetting to use the syringe cap, etc., all cause errors. A hidden drop of water is the most common error. Gases prepared in the laboratory always have a small quantity of air (5 mL/60 mL), and this alters the molar mass obtained by each method. Method I required use of the temperature and barometric pressure, also potential sources of error.

3. Generally gases with larger molar masses give better results, however, affinity for water is an issue. Argon is a heavy gas that gives great results; HCl(g) does not.
4. Gases that are water-soluble tend to give higher-than-expected results because they dissolve in droplets of water present. This removes them from the gas phase and more gas moves into the syringe. The mass of the compound in the syringe is too high because it includes the gas dissolved.

Chapter 10, Limiting Reagent

2. The amount of HCl used, in mmol, is given by the formula $n_{\text{HCl}} = \text{mass}_{\text{HCl}} \times \text{concentration}_{\text{HCl}} / \text{density}_{\text{HCl}}$. The concentration, approximately 2 mol/L can be written as 2 mmol/mL. Your answer will be approximately 2.8 mmol HCl.
3. The relationship between length of Mg and mass will be provided. For example, you may be given 20 cm Mg has a mass of 0.20 g Mg. This can be written as 0.20 g Mg/20 cm Mg or 0.010 g Mg/cm Mg. The mass of any length of Mg can be estimated by multiplying this constant by the length.
4. The reaction stoichiometry requires 2 mol HCl for every 1 mol Mg. Given the example in Answer 1 (2.8 mmol HCl), 1.4 mmoles of Mg are needed to react with the HCl.
5. Convert the amount of Mg (mmol) into mass of Mg (g). Then convert mass into length using the conversion provided by your instructor.
6. There should be a “kink” in the plot around the length determined in Question 5. In the figure on page 125, this kink occurs at 4.0 cm Mg.

Chapter 11. Determining the Barometric Pressure without a Barometer

2. This equation has the general form of $y = mx + b$. where y relates to V and x relates to n . The slope of the line m is RT/P .
3. The computer would give a better value of m , rather than estimating Δx and Δy to determine m .

Chapter 12 Experiments with Nitrogen Oxides

Preparation of Nitric Oxide

1. If you used 0.25 g NaNO₂(s) as recommended, the amount of sodium nitrite used 0.0036 mol NaNO₂(s).
2. The balanced equation (page 147) requires that 2 mol NaNO₂ produces 2 mol NO(g). Therefore, we can expect 0.0036 mol NO(g).
3. The expected volume is 89 mL NO(g).
4. The calculated volume of nitric oxide is considerably larger than the experimentally obtained volume.
5. The reactants: NaNO₂: Na is +1, N is +3, and O is -2; FeSO₄: Fe is +2, S is +6 and O is -2; H₂SO₄: H is +1, S is +6 and O is -2; The products: NO: N is +2 and O is -2; Fe₂(SO₄)₃: Fe is +3, S is +6, and O is -2; NaHSO₄: Na is +1, H is +1, S is +6, and O is -2; H₂O: H is +1 and O is -2. The oxidizing agent is the nitrogen in NaNO₂ and the reducing agent is the Fe⁺² in FeSO₄.

Chapter 12, Experiment 1. Conversion of nitric oxide to nitrogen dioxide

1. $2 \text{NO(g)} + \text{O}_2\text{(g)} \longrightarrow 2 \text{NO}_2\text{(g)}$
2. The reaction was exothermic and kinetically fast.

3. According to the reaction stoichiometry, we would expect 40 mL $\text{NO}_2(\text{g})$.
4. If 10 mL of $\text{NO}(\text{g})$ and 30 mL of $\text{O}_2(\text{g})$ were placed together in a container, only 5 mL $\text{O}_2(\text{g})$ would react before the $\text{NO}(\text{g})$ present was depleted. Together they would make 10 mL $\text{NO}_2(\text{g})$. There would be 25 mL unreacted $\text{O}_2(\text{g})$ for a total volume of gases equal to 10 mL + 25 mL = 35 mL.
5. The law of combining volumes is useful in chemical industry because it is crucial to know what change in volume is expected when a reaction takes place. If additional moles of gas are being produced, there must be some accommodation allowed for the new volume or pressure. In this experiment, the final volume is less than the combined original volumes. When scaled up, this could be equally disastrous because the reaction vessel could collapse (implode) due to the reduced pressure inside. Again, carefully designing the equipment can prevent this problem.

Chapter 12, Experiment 2. From nitrogen dioxide to nitric acid

2. Dissolving does not entail a chemical change. The substance is now in a solution.
3. Not all of the gas in the syringe reacted because oxygen is required to complete the reaction.
4. $4 \text{NO}_2(\text{g}) + 2 \text{H}_2\text{O}(\text{l}) + \text{O}_2(\text{g}) \longrightarrow 4 \text{HNO}_3(\text{aq})$

Chapter 12, Experiment 3. LeChatelier principle and the $\text{NO}_2/\text{N}_2\text{O}_4$ equilibrium

1. The color of the gas darkened because as the volume decreased, the concentration increased. Within a second, however, the color lightened somewhat as the system returned to equilibrium.
2. The dissociation of any bond is endothermic. Bond energies are all positive numbers; it takes that amount of energy to break the bond.
3. The nitrogen-nitrogen single bond dissociation energy in your chemistry book is probably in the range 150 - 250 kJ/mol. Most bond dissociation enthalpies are known with more certainty and represent a reasonable estimation of the strength of the bond in question, independent of the molecule. With nitrogen-nitrogen single bonds, a good estimation is difficult to make because there are relatively few compounds with N-N single bonds. If N_2O_4 were the only compound with a N-N single bond, the N-N single bond dissociation energy would be +57 kJ/mol, exactly opposite of $\text{DH}_{\text{dissociation}} = -57 \text{ kJ}$.
4. See page 143 for the Lewis structure of NO_2 . To draw a Lewis structure for N_2O_4 , draw two NO_2 molecules with the lone pair connected to form a single bond. Resonance exists and the nitrogen-oxygen bond order is 1.5.
5. The quadratic equation is required here. The results are $[\text{NO}_2] = 0.0204 \text{ mol/L}$ and $[\text{N}_2\text{O}_4] = 0.0898 \text{ mol/L}$.

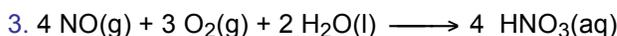
Chapter 12, Experiment 4. High temperature favors the endothermic substance

1. The red-brown color intensified when the gas was heated because adding heat favors the endothermic direction and that means more red-brown $\text{NO}_2(\text{g})$ is formed.
2. Is the reaction as written is exothermic.
3. See the answer to Question 1.
4. $K_c^{25} = 0.00465$ and $\text{DH} +57 \text{ kJ/mol}$.
5. In your sketch, the activation energy should be relatively small.

Chapter 12, Experiment 5. Acid rain microchemistry

1. The initial pH of the buffered lakes should have been around 10 or so due to the sodium bicarbonate ion. The initial pH of the unbuffered lakes would depend on the water source — but should be between 6 and 8. After the addition of $\text{NO}_2(\text{g})$, the pH of the unbuffered lakes rises within a few minutes — well before the pH changes for the buffered lakes.
2. The buffer moderates the effect of added acid or base in order to stabilize the pH.
3. A rock-bottomed lake would be more affected by acid rain.

Chapter 12, Experiment 6. Acidic nature of nitrogen oxides



4. Gases have the most mobility; they mix and migrate rapidly so that molecules that are capable of reacting will sooner or later come in contact and the reaction may take place, although there are a variety of other factors to consider. After gas-gas mixtures, reagents in a homogeneous solution would have good mobility. Solids and liquids have difficulties because the solid cannot migrate and the only place that the reaction can take place is at the surface of the solid.

Chapter 12, Experiment 7. Well plate reactions involving nitric oxide

1. Test tube #1 ($\text{KMnO}_4(\text{aq})$): the purple color turned either colorless or possibly produced brownish solid; Test tube #2 ($\text{Br}_2(\text{aq})$): the red color turned colorless; Test tube #3 ($\text{Fe}^{+2}(\text{aq})$): the solution turned black; Test tube #4 ($\text{I}^-(\text{aq})$): the solution turned brown.
2. I_2 (probably in the form of I_3^-) causes the solution to be yellow.
3. The oxidation number of the nitrogen atom in $\text{NO}(\text{g})$ is +2 and in NO_3^- is +5.
4. $\text{NO}(\text{g})$ acts as a reducing agent when it reacts with the $\text{Br}_2(\text{aq})$.
5. The brown ring test could be used to qualitatively test for Fe^{+2} .

Chapter 12, Experiment 8. Dinitrogen trioxide is a blue liquid

1. The Lewis dot structure of NO_2 and NO are given on page 143. The Lewis dot structure of N_2O_3 is obtained by joining the two lone pair electrons on NO_2 and NO to form a N-N bond.
2. N_2O_3 is a liquid and is diamagnetic, while NO_2 and NO are gases and paramagnetic.
3. If the nitrogen-nitrogen bond is broken, energy is absorbed. The energy to accomplish this must come from the surroundings.

Chapter 13 Experiments with Ammonia

Preparation of Ammonia

1. The equilibrium $\text{NH}_3(\text{aq}) \xrightleftharpoons{\hspace{1cm}} \text{NH}_3(\text{g})$ ($\Delta H = +34.2 \text{ kJ/mol}$) indicates that $\text{NH}_3(\text{aq})$ is favored. $\text{NH}_3(\text{g})$ is the endothermic substance and to produce more of it, we must add heat.
2. The volume of $\text{NH}_3(\text{g})$ decreases upon standing because ammonia is exceedingly soluble in water and will dissolve in tiny drops of water present. Also, the ideal gas law accounts for some decrease in the gas volume with a decrease in temperature.

Chapter 13, Experiment 1. Ammonia is a base

1. Universal indicator and would have been red, indicating the pH was 4 or lower. Red cabbage solution would have been red indicating the pH was 4 or lower.
2. Because ammonia is a base, the pH of the solution should have risen to 9 as the $\text{NH}_3(\text{g})$ was discharged over the surface.
3. $\text{NH}_3(\text{g})$ is very soluble in the water. Some ammonia odor may have been detectable because the equilibrium ($\text{NH}_3(\text{aq}) \rightleftharpoons \text{NH}_3(\text{g})$) always provides for some ammonia in the gas phase.
4. $\text{NH}_3(\text{g})$ is a basic compound; it causes the pH of solutions to rise.
5. The color of the solution may have been different at the top than at the bottom because the ammonia dissolves at the surface and is thus regionally concentrated there. With time, the solution will mix and become pH homogeneous.

Chapter 13, Experiment 2. Ammonia Fountain

1. The small fountain results from ammonia's incredible affinity for and solubility in water.
2. Ammonia is very soluble in water, and it dissolves rapidly.
3. The Lewis structure of ammonia should have featured three single bonds to three hydrogen atoms and there should have been a lone pair of electrons on nitrogen. The Lewis structure of water should have featured two single bonds to two hydrogen atoms and there should have been two lone pairs of electrons on oxygen. Hydrogen bonding forces exist between ammonia molecules and water molecules.
4. It is not possible to collect $\text{NH}_3(\text{g})$ over water due to its solubility in water. Joseph Priestley, in the 1770s, collected it over mercury.

Chapter 13, Experiment 3. Acid-base reactions with fruit juices

1. Unsweetened fruit juices have a sour taste and are therefore acidic.
2. $\text{NH}_3(\text{g})$ is a base.
3. The pH of the fruit juice increased after adding the $\text{NH}_3(\text{g})$.
4. As the acid concentration increases, the pH decreases. As the acid concentration decreases, the pH increases.
5. An acid and base reaction occurred when we added $\text{NH}_3(\text{g})$ to fruit juice samples.

Chapter 13, Experiment 4. Ammonia Is More Soluble at Low Temperature

1. Ammonia re-dissolved in the water present as the syringe cooled down.
2. $\text{NH}_3(\text{g})$ is so water-soluble because of hydrogen bonding between the O-H hydrogen on water and the lone pair on ammonia and between the N-H hydrogen on ammonia and the lone pair on water. See your answer to Experiment 2, Question 3 for the Lewis dot structures.
3. Re-heating the $\text{NH}_3(\text{aq})$ would produce $\text{NH}_3(\text{g})$.
4. Your sketch should show ammonia's solubility (y-axis) decreasing as the temperature (x-axis) increases.
5. The reaction profile should show $\text{NH}_3(\text{aq})$ lower on energy than $\text{NH}_3(\text{g})$. The energy of activation should be very small.
6. $\Delta H_{\text{solution}} = -34.2 \text{ kJ/mol}$
7. $\Delta G = +10 \text{ kJ/mol}$ at 298 K.

Chapter 13, Experiment 5. Gaseous Ammonia Reacts with Hydrogen Chloride

1. The salt produced is ammonium chloride with the formula NH_4Cl .
2. $\text{NH}_3(\text{g}) + \text{HCl}(\text{g}) \longrightarrow \text{NH}_4\text{Cl}(\text{s})$
3. Yes, ammonium chloride is water-soluble. All ammonium salts are soluble.
4. The plunger(s) moved inward because the two gaseous reagents were being converted into a solid. See the balanced equation in Answer 2.
5. The plunger(s) stop moving inward after all of the $\text{HCl}(\text{g})$ had been consumed. Recall that we obtained $\text{HCl}(\text{g})$ from the headspace above a solution of concentrated $\text{HCl}(\text{aq})$. This headspace gas is mostly air. If we used equal volumes of pure (100%) $\text{HCl}(\text{g})$ and $\text{NH}_3(\text{g})$, there would have been no volume of gas left.

Chapter 13, Experiment 6. Ammonia Forms Nitric Oxide

1. $\text{NO}(\text{g})$ rapidly reacts with oxygen from the air (Chapter 12, Experiment 2) as per $2 \text{NO}(\text{g}) + \text{O}_2(\text{g}) \longrightarrow 2 \text{NO}_2(\text{g})$
2. $\text{NO}(\text{g})$ is converted into $\text{NO}_2(\text{g})$ as in the equation in the previous answer. The $\text{NO}_2(\text{g})$ is then reacted with water to produce nitric acid: $3 \text{NO}_2(\text{g}) + \text{H}_2\text{O}(\text{l}) \longrightarrow 2 \text{HNO}_3(\text{aq}) + \text{NO}(\text{g})$
3. A catalyst is usually a chemical that speeds up a chemical reaction by providing an alternative reaction pathway that is faster.
4. The copper wire stays hot as long as the ammonia/oxygen mixture is being discharged over its surface because the reaction is exothermic and the heat keeps the copper hot.

Chapter 13, Experiment 7. Ammonia Forms Complex Ions

1. $[\text{Co}(\text{NH}_3)_6]^{+2}(\text{aq})$, $[\text{Cu}(\text{NH}_3)_4]^{+2}(\text{aq})$, $[\text{Ni}(\text{NH}_3)_6]^{+2}(\text{aq})$, $[\text{Ag}(\text{NH}_3)_2]^+(\text{aq})$
2. Determine the oxidation state for the metal in each complex ion. The first one is: Co^{+2} , Cu^{+2} , Ni^{+2} , Ag^+
3. What is the coordination number of the complex ion in each compound? For Co^{+2} , the answer is six, Cu^{+2} is four, Ni^{+2} is six and Ag^+ is two.

Chapter 14 Experiments with Ethyne

Preparation of ethyne

1. $\text{CaC}_2(\text{s}) + 2 \text{H}_2\text{O}(\text{l}) \longrightarrow \text{C}_2\text{H}_2(\text{g}) + \text{Ca}(\text{OH})_2(\text{s}, \text{aq})$
2. 0.20-g calcium carbide, CaC_2 (MM = 64) corresponds to 0.0031 mol calcium carbide
3. 5 mL water (density = 1 g/mL, MM = 18 g/mol), corresponds to 0.28 mol H_2O
4. Calcium carbide is the limiting reactant.
5. The theoretical yield of $\text{C}_2\text{H}_2(\text{g})$ is 0.0031 mol $\text{C}_2\text{H}_2(\text{g})$.
6. The theoretically expected volume (mL) of $\text{C}_2\text{H}_2(\text{g})$ is 76 mL.

Chapter 14, Experiment 1. Ethyne reacts with permanganate

1. The Lewis structure of ethyne features a triple bond between the two carbons and from each carbon, a single bond to a hydrogen.
2. Permanganate ion was reduced as it was converted into MnO_2 .
3. Ethyne was oxidized.

- The triple bond can be thought of as increasing the overall bond strength of the double bond by only 225 kJ/mol ($614 + 225 = 839$). The reactivity of the C-C triple bond is based on the fact that C-C single bonds, or in our case, carbon-oxygen double bonds, are much stronger than the second or third bond in a C-C triple bond.

Chapter 14, Experiment 2. Sooty combustible of ethyne

- $2 \text{C}_2\text{H}_2(\text{g}) + \text{O}_2(\text{g}) \longrightarrow 4 \text{C}(\text{g}) + 2 \text{H}_2\text{O}(\text{g})$
- The combustion reaction was incomplete if soot were formed. Soot is elemental carbon, C(s). It can be further combusted to form $\text{CO}_2(\text{g})$.
- Carbon dioxide and water are expected as the products of complete combustion. Would you not expect to see soot with the complete combustion of ethyne.

Chapter 14, Experiment 3. Banging bubbles!

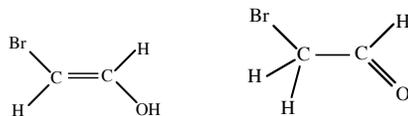
- The loudest bang probably came from a bubble with an excess of oxygen present.
- These are explosions are not particularly dangerous because (a) they are small in scale and (b) the gases are not tightly bound in a rigid container; the soap film easily bursts and the expanding gases can dissipate into the air without causing damage.
- Explosive mixtures are sometimes stable if the reaction they undergo has a high energy of activation. A spark or open flame, however, provides enough energy to trigger the reaction. Ethyne/oxygen mixtures are very dangerous because the energy of activation is not especially high — even static electricity will trigger the explosion.

Chapter 14, Experiment 4. Ethyne/oxygen rockets

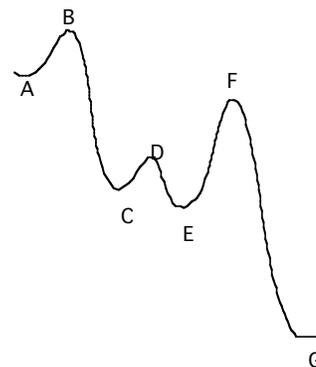
- $2 \text{C}_2\text{H}_2(\text{g}) + 5 \text{O}_2(\text{g}) \longrightarrow 4 \text{CO}_2(\text{g}) + 2 \text{H}_2\text{O}(\text{g})$
- The explosion creates a force that pushes against the water plug in the stem. If the water is not present, most of the force is dissipated as the gases leave the rocket.
- The combustion of ethyne is more exothermic than that of methane or propane as a fuel. Carbon-carbon triple bonds are “high energy” bonds — easily forming stronger bonds. (See Experiment 5, Question 1). In the combustion of ethyne vs. methane or propane, in all three cases, carbon dioxide is formed. With ethyne/oxygen, the reactants are just starting much higher in energy than methane/oxygen.
- The contents of the rocket should have passed the limewater test for carbon dioxide.

Chapter 14, Experiment 5. Ethyne reacts with aqueous bromine

- 1 and 2. The structure of CHBrCHOH features a double bond (left) and CH_2BrCHO has an aldehyde function (right):



- Ethyne and bromine are indicated as A. They quickly react (small energy of activation, B) to form the compound CHBrCHOH , C, which exists in equilibrium (very small energy of activation, D) with CH_2BrCHO , E. This equilibrium mixture eventually (larger energy of activation, F) forms the final stable products, G.



- Test both with bromine-water.

Chapter 15 Experiments with Sulfur Dioxide

1. $\text{NaHSO}_3(\text{s}) + \text{HCl}(\text{aq}) \longrightarrow \text{SO}_2(\text{g}) + \text{NaCl}(\text{aq}) + \text{H}_2\text{O}(\text{l})$
2. Sulfur dioxide is a polar molecule and is expected to be soluble in water.
3. From the ideal gas law, we know that volume is proportional to moles, so we need not convert 10 mL into moles. We can simply calculate $\text{Conc}(\text{ppm}) = 106 \times V_{\text{sulfur dioxide}}/V_{\text{air}}$. The two volumes must be in the same units, mL, for example. The concentration of sulfur dioxide is 10 ppm.

Chapter 15, Experiment 1. Sulfur dioxide reacts with water

1. $\text{SO}_2(\text{g})$ is an acid.
2. Sulfur dioxide is heavier than air and is extremely soluble in water.
3. (a) $\text{Na}_2\text{O}(\text{s}) + \text{H}_2\text{O}(\text{l}) \longrightarrow 2 \text{NaOH}(\text{aq})$; (b) $\text{P}_2\text{O}_5(\text{s}) + 3 \text{H}_2\text{O}(\text{l}) \longrightarrow 2 \text{H}_3\text{PO}_4$

Chapter 15, Experiment 2. Sulfur dioxide reacts quickly with sodium hydroxide

1. Sulfur dioxide is the acid and aqueous sodium hydroxide is the base.
2. Sulfur dioxide dissolved in water and reacted with the sodium hydroxide.
3. The water molecule should have two lone pairs of electrons and the sulfur dioxide molecule should have one lone pair.

Chapter 15, Experiment 3. Sulfur dioxide and potassium permanganate react

1. The color change is the most obvious indication that a reaction has taken place.
2. $3 \text{SO}_2(\text{g}) + 2 \text{MnO}_4^-(\text{aq}) + 2 \text{H}_2\text{O}(\text{l}) \longrightarrow 2 \text{MnO}_2(\text{s}) + 3 \text{SO}_4^{2-}(\text{aq}) + 4 \text{H}^+(\text{aq})$
3. $\text{Mn}^{2+}(\text{aq})$ is a colorless solution and $\text{MnO}_2(\text{s})$ is a brown precipitate.

Chapter 15, Experiment 4. Sulfur dioxide discolors many natural colors

1. Expose the flowers to a base such as ammonia gas.
2. The first one is an oxidation-reduction reaction and the second one is an acid-base reaction.

Chapter 15, Experiment 5. Acid-rain microchemistry

1. The buffered lakes do not become acidic as fast as the unbuffered lakes.
2. The initial pH of the buffered lakes should have been around 10 or so due to the sodium bicarbonate ion. The initial pH of the unbuffered lakes would depend on the water source — but should be between 6 and 8. After the addition of $\text{SO}_2(\text{g})$, the pH of the unbuffered lakes rises within a few minutes — well before the pH changes for the buffered lakes.
3. The buffer moderates the effect of added acid or base in order to stabilize the pH.
4. Acid rain caused by sulfur dioxide originates from burning coal, not petroleum.

Chapter 15, Experiment 6. Sulfur dioxide reacts with aqueous bromine

1. $\text{Br}_2(\text{aq})$ is the oxidizing agent and $\text{SO}_2(\text{g})$ is the reducing agent.
2. oxidation: $\text{SO}_2(\text{g}) + 2 \text{H}_2\text{O}(\text{l}) \longrightarrow \text{SO}_4^{2-}(\text{aq}) + 4 \text{H}^+(\text{aq}) + 2 \text{e}^-$; (b) reduction: $\text{Br}_2(\text{aq}) + 2 \text{e}^- \longrightarrow 2 \text{Br}^-(\text{aq})$
3. You could confirm the presence of bromide ion by precipitating it with $\text{Ag}^+(\text{aq})$
4. The oxidation number of sulfur in SO_2 and SO_3^{2-} is +4 and in SO_3 and SO_4^{2-} is +6.
5. One pair is SO_2 and SO_3^{2-} and the other pair is SO_3 and SO_4^{2-} .

Chapter 16 Experiments with Chlorine

1. See page 223 for the equation for both methods.
2. Method A: $0.22 \text{ g Ca(OCl)}_2(\text{s}) = 0.00154 \text{ mol Ca(OCl)}_2$; $5 \text{ mL } 2 \text{ M HCl}(\text{aq}) = 0.010 \text{ mol HCl}$. Limiting reagent is Ca(OCl)_2 . Theoretical yield is $0.0308 \text{ mol Cl}_2(\text{g})$ from the reaction stoichiometry. At $25 \text{ }^\circ\text{C}$ and standard pressure, this volume would occupy 75 mL . Method B: $1.0 \text{ mL } 6 \text{ M HCl}(\text{aq}) = 0.006 \text{ mol HCl}$; $3.0 \text{ mL household bleach} = 0.18 \text{ mol NaOCl}$. Limiting reagent is HCl . Theoretical yield is $0.003 \text{ mol Cl}_2(\text{g})$ from the reaction stoichiometry. At $25 \text{ }^\circ\text{C}$ and standard pressure, the volume is 73 mL .
3. Method A: Oxidation numbers for the chlorine in Ca(OCl)_2 , HCl , and Cl_2 are $+1$, -1 and 0 , respectively. Method B: Oxidation numbers for the chlorine in HCl , NaOCl , Cl_2 , and NaCl are -1 , $+1$, 0 and -1 , respectively.

Chapter 16, Experiment 1. Chlorine and sodium hydroxide form bleach

1. (a) $\text{Cl}_2(\text{g}) + \text{NaOH}(\text{aq}) \longrightarrow \text{NaOCl}(\text{aq}) + \text{NaCl}(\text{aq})$; (b) The evidence that a chemical change took place is not compelling. The chlorine gas goes into solution, but we have seen numerous examples of gases that simply dissolve in water without reacting.
2. Oxidation numbers for the chlorine in Cl_2 , NaOCl , and NaCl are 0 , $+1$, and -1 , respectively. Chlorine, Cl_2 , was being both oxidized and reduced. This is called disproportionation.
3. $2 \text{ HCl}(\text{aq}) + \text{NaOCl}(\text{aq}) \longrightarrow \text{Cl}_2(\text{g}) + \text{NaCl}(\text{aq}) + \text{H}_2\text{O}(\text{l})$
4. $\text{Cl}_2 + \text{SO}_3^{2-} + \text{H}_2\text{O} \longrightarrow 2 \text{ Cl}^- + \text{SO}_4^{2-} + 2 \text{ H}^+$

Chapter 16, Expt 2. Chlorine disproportionates in water to form acidic species

1. The starting pH was around 7 and the final pH should have been around 4 . In terms of concentration, at pH 7 , $[\text{H}_3\text{O}^+] = 1 \times 10^{-7}$ and at pH 4 , $[\text{H}_3\text{O}^+] = 1 \times 10^{-4}$. The ratio $[\text{H}_3\text{O}^+]_{\text{solution}}/[\text{H}_3\text{O}^+]_{\text{water}} = 1000:1$
2. The main difference is that $\text{HCl}(\text{aq})$ is produced rather than $\text{NaCl}(\text{aq})$. This accounts for the differences in resulting pH.
3. Disproportionation refers to reactions in which one species is both oxidized and reduced. In this case it is Cl_2 .
4. HCl is a strong acid and HOCl is a weak acid.

Chapter 16, Experiment 3. Chlorine discolors the natural colors of fruit juices

1. The color changes involving ammonia were explained in terms of acid-base chemistry.
2. The color changes involving sulfur dioxide were explained in terms of either acid-base chemistry or reduction chemistry.
3. Chlorine is an oxidizing agent.

Chapter 16, Experiment 4. Testing colorfast fabrics

2. Green food coloring is made from blue and yellow food coloring. Treating green food coloring with bleach will cause it to turn yellow within a few seconds and then finally colorless a short time later.

Chapter 16, Experiment 5. Chlorine reacts with aqueous sodium sulfite

1. One may eliminate residual $\text{Cl}_2(\text{g})$ from a fish tank by treating it with sodium bisulfite solution or sodium sulfite solution.
2. Sulfite was oxidized and chlorine was reduced.
3. $\text{Cl}_2(\text{g}) + \text{HSO}_3^-(\text{aq}) + \text{H}_2\text{O}(\text{l}) \longrightarrow 2 \text{ Cl}^-(\text{aq}) + \text{HSO}_4^-(\text{aq}) + 2 \text{ H}^+(\text{aq})$

Chapter 16, Experiment 6. The halogen activity series

1. The bromide ion is colorless; bromine (Br_2) is red. Bromide was oxidized in the reaction.
2. The iodide ion is colorless; iodine (I_2) is brown-yellow in aqueous solution. Iodide was oxidized in the reaction.
3. Chloride and aqueous bromine would not react. Bromide and aqueous chlorine do react to produce chloride and aqueous bromine.
4. Chlorine is the easiest to reduce. Iodide is easiest to oxidize.

Chapter 16, Experiment 7. Chlorine and sodium form sodium chloride

1. Oxidation-reduction.
2. Sodium is a neutral metal, chlorine is a covalent molecules and sodium chloride is an ionic compound.
3. Metallic substances have high melting and boiling points, no solubility in water, high ductility, high malleability, and high ability to conduct electricity; ionic substances have high melting and boiling points, variable solubility in water (some are highly soluble and others have low solubility), no ductility, no malleability, and low ability to conduct electricity except when molten or in aqueous solution; covalent molecules have variable melting and boiling points (they can be solids, liquids or gases), variable solubility in water, generally low ductility and malleability, and generally low or no ability to conduct electricity.

Chapter 16, Experiment 8. Hydrogen/chlorine rockets

2. The pH test results should have shown an acid present in the rocket.
3. The rocket reactions require a spark to initiate because the energy of activation is quite high. The reaction is also highly exothermic, so once started, it is self-sustaining.

Chapter 16, Experiment 9. Chemiluminescence and singlet oxygen.

1. 630 nm is red.
2. Singlet oxygen is higher in energy than triplet oxygen.
3. $\Delta H = -92 \text{ kJ}$.
4. We breathe triplet oxygen.

Chapter 16, Experiment 10. Spectacular underwater fireworks!

1. Light energy was released in the reaction. The other types of energy associated with chemical reactions are heat and electrical.
2. Soot is carbon, $\text{C}(\text{s})$.
3. A similar reaction would take place if chlorine and ethyne were directed towards each other in air.
4. The reaction profile should show an exothermic reaction with a very small activation energy.

Chapter 16, Experiment 11. Liquid and solid chlorine

1. $\text{Cl}_2(\text{s}) \rightleftharpoons \text{Cl}_2(\text{l}) \rightleftharpoons \text{Cl}_2(\text{g})$
2. Liquid and solid chlorine are green-yellow.
3. The plunger moved inward as the amount of gas decreased due to the phase change.
4. London dispersion forces must be overcome when liquid chlorine becomes gaseous chlorine?

Chapter 19 Experiments with Hydrogen Chloride

1. 1.1 g anhydrous NaHSO₄ corresponds to 0.0092 mol NaHSO₄ and 0.7 g sodium chloride, NaCl corresponds to 0.012 mol NaCl. Because they react 1:1, NaHSO₄ is the limiting reagent and there should be 0.0092 mol HCl(g) produced.
2. 224 mL HCl(g)
3. They melt together and effervesce.
4. We made HCl by this method because the reaction is too slow at room temperature.
5. We are most familiar with HCl in aqueous form as hydrochloric acid.
6. 294 L
7. $\text{NaHSO}_4 \cdot \text{H}_2\text{O} (\text{s}) + \text{NaCl}(\text{s}) \xrightarrow{\Delta} \text{HCl}(\text{g}) + \text{Na}_2\text{SO}_4(\text{s}) + \text{H}_2\text{O}(\text{g})$

Chapter 19, Experiment 1. Formation of an aerosol

1. The aerosol eventually condensed into larger droplets and returned to the surface of the water.
2. Refer to pH color chart.
3. The red solution, containing HCl(aq) is slightly more dense than water
4. The solution will eventually become all red, representing a homogeneous pH

Chapter 19, Experiment 2. White clouds

1. The white cloud above the ammonium hydroxide solution is solid ammonium chloride, NH₄Cl(s).
2. The HCl reacted with the fumes of ammonia above the liquid.
3. The white cloud dropped as it spilled over the top of the vial indicating that it is heavier than air.
4. If one connected a syringe of HCl(g) to a syringe of NH₃(g), solid, NH₄Cl would form and the volume of gas in the syringes would decrease.

Chapter 19, Experiment 3. Acid Snow?

1. LeChatelier's theory predicts the formation of more solid NaCl if more chloride is added to the solution. This occurs when HCl(g) dissolves: $\text{HCl}(\text{g}) \longrightarrow \text{H}^+(\text{aq}) + \text{Cl}^-(\text{aq})$ and then the presence of more chloride shifts the equilibrium left.
2. The glittering is caused from light being reflected off of NaCl(s) crystals. The crystals have flat surfaces and reflect light well.
3. Saturated NaCl(aq) is prepared by heating NaCl(s) and water together. It is necessary to use more NaCl than can dissolve. Allow the solution to cool to room temperature. Note: NaCl is more soluble at higher temperatures, but not all salts are. Some salts, like calcium acetate, are more soluble at low temperatures.
4. The one solid NaCl produced can be redissolved by heating the solution.

Chapter 19, Experiment 4. Fizzzzzz!

1. The reaction is audible because thousands of bubbles are breaking and each makes a small amount of noise. (Put your ear up to an open carbonated beverage container.) The pitch would be lower if you were to use a larger vial. Variations in pitch would also occur if three vials of various depth were used.
2. $\text{HCl}(\text{g}) + \text{HCO}_3^-(\text{aq}) \longrightarrow \text{CO}_2(\text{g}) + \text{H}_2\text{O} + \text{Cl}^-(\text{aq})$

3. The spectator ion $\text{Na}^+(\text{aq})$ is not listed in the above chemical reactions. If the reaction were carried out to completion, the product $\text{NaCl}(\text{aq})$ would be obtained.

Chapter 19, Experiment 5. Curdles away

1. Skim milk may not work as well.
2. Salad dressing, mayonnaise and yogurt are other common emulsions.
3. Vinegar is used in the kitchen to curdle milk?

Chapter 19, Experiment 6. Levitating paper

1. calcium carbonate in office paper causes it to be white.
2. The mass of office paper decreases with the reaction with acid because the solid calcium carbonate produces gas and water soluble calcium: $\text{CaCO}_3(\text{s}) + 2 \text{H}^+(\text{aq}) \longrightarrow \text{Ca}^{2+}(\text{aq}) + \text{CO}_2(\text{g}) + \text{H}_2\text{O}(\text{l})$
3. Yes, bubbles would form on a sheet of paper if vinegar were spilled on it.
4. Because we used $\text{HCl}(\text{aq})$, the final form of calcium is $\text{CaCl}_2(\text{aq})$.

Chapter 19, Experiment 7. Hydrogen chloride fountain

1. This is a physical change.
2. The resulting pH is low.
3. The fountain occurs because the water is being inwards due to the difference in pressure between the external pressure and the pressure inside the syringe.
4. Some air remains in the syringe after the reaction because it was there from the beginning — by the design of the preparation, air is present.

Chapter 20 Experiments with Carbon Monoxide

1. A mass of 0.23 g formic acid correspond to 0.005 mol HCOOH .
2. 122 mL $\text{CO}(\text{g})$
3. Some bubbles form as soon as the two acids are mixed. If we did not heat the reaction, the reaction would proceed, but very slowly.
4. Sulfuric acid is not a reagent, per se. It functions as a desiccant, removing water (H_2O) from formic acid, HCOOH to give CO .
5. As sulfuric acid ties up more and more water molecules, its ability to function as a desiccant decreases.
6. The generation of $\text{HCl}(\text{g})$ would have a the larger energy of activation.

Chapter 20, Experiment 1. Blue jets!

1. $\text{CO}(\text{g}) + \frac{1}{2} \text{O}_2(\text{g}) \longrightarrow \text{CO}_2(\text{g})$
2. A blue color is produced when CO burns.
3. The blue color is evidence that $\text{CO}(\text{g})$ is present near the base of the flame.
4. Because carbon monoxide burns, but in carbon dioxide, the carbon is fully oxidized and will not burn.

Chapter 20, Experiment 2. Wimpy soap bubble explosions and wimpy rockets.

1. In addition to the carbon monoxide and oxygen rockets, we have experimented with hydrogen/oxygen, methane/oxygen, hydrogen/chlorine, ethyne/oxygen, ethene/oxygen,

and nitrous oxide/hydrogen. The carbon monoxide/oxygen system makes one of the weakest explosions and wimpiest of rockets.

- Equation B: $\Delta H = -393.5$; Equation C: $\Delta H = -110.5$. To get ΔH for Equation A, one must flip Equation C, making it $+110.5$ kJ. Adding Equation B + Flipped Equation C gives $\Delta H = -283.0$ kJ

Chapter 20, Experiment 3. Carbon monoxide detectors

- The detector waits until a sustained high level of CO is detected so that "bursts" of CO, such as from cigarette smoke does not trigger the alarm.
- 5000 ppm CO
- 0.017 ppm or 17 ppb

Chapter 20, Experiment 4. Reduction of copper(II) oxide with carbon monoxide

- $2 \text{Cu(s)} + \text{O}_2\text{(g)} \longrightarrow 2 \text{CuO(s)}$
- The reaction has a high energy of activation.
- $\text{CuO(s)} + \text{CO(g)} \longrightarrow 2 \text{Cu(s)} + \text{CO}_2\text{(g)}$
- Oxidation numbers: for CuO: copper +2, oxygen -2; CO: carbon +2, oxygen -2; Cu 0; CO_2 : carbon +4, oxygen -2. The oxidizing agent is CuO and the reducing agent is CO.
- $\Delta H = -126$ kJ

Chapter 20, Experiment 5. Reduction of palladium and silver ions with carbon monoxide

- The ions Pd^{+2} and Ag^+ are easily reduced to their elements.
- Metals are not soluble in water; metal ions are.
- Limewater test
- In the reduction of Pd^{+2} , the pH decreases due to the formation of $\text{H}^+\text{(aq)}$.

Chapter 20, Experiment 6. Carbon monoxide and potassium permanganate react

- $2 \text{MnO}_4^-\text{(aq)} + 5 \text{CO(g)} + 6 \text{H}^+\text{(aq)} \longrightarrow 2 \text{Mn}^{+2}\text{(aq)} + 5 \text{CO}_2\text{(g)} + 3 \text{H}_2\text{O(l)}$
 $2 \text{MnO}_4^-\text{(aq)} + 3 \text{CO(g)} + 2 \text{H}^+\text{(aq)} \longrightarrow 2 \text{MnO}_2\text{(aq)} + 3 \text{CO}_2\text{(g)} + \text{H}_2\text{O(l)}$
- In both cases $\text{MnO}_4^-\text{(aq)}$ is the oxidizing agent and CO is the reducing agent.

Chapter 20, Experiment 7. Carbon monoxide poisoning

- The CO-hemoglobin complex is red in color.
- CO binds reversibly to hemoglobin.
- The person would exhibit some of the symptoms of CO-poisoning.

Chapter 20, Experiment 8. Formation of a simple organometallic compound

- There is a triple bond between C and O with a lone pair on each C and O. The formal charge is -1 for C and +1 for O.
- The electron pair on C is most easily donated to a Lewis acid.
- The oxidation number on copper is +1, making it a Lewis acid.
- Carbon has a formal charge of -1 so it is the electron pair on carbon that is donated to the Fe^{+2} of hemoglobin. The interaction is Lewis acid-base with the metal ion being the acid.

In a similar way, $O_2(g)$ has two lone pairs on each oxygen atoms and formal charges of 0. The lone pairs interact with Fe^{+2} of hemoglobin in an acid-base interaction, however, the interaction is not as strong as it was with CO/ Fe^{+2} because of the formal charge differences.

Chapter 20, Experiment 9. The Water-gas Shift reaction

1. Water-gas is the mixture of $CO(g)$ and $H_2(g)$. The water-gas shift reaction is $CO(g) + H_2O(g) \longrightarrow CO_2(g) + H_2(g)$
2. Both ΔH and ΔS are negative. The reaction is enthalpy favored, but entropy unfavored.
3. $\Delta G = 0$ at 969 K. At temperatures below 969 K, the reaction is spontaneous, with $\Delta G < 0$.
4. The water-gas shift reaction is useful because it converts $CO(g)$ and water into more hydrogen. The $CO(g)$ is converted into harmless carbon dioxide.

Chapter 21 Experiments with Ethene

1. Anhydrous ethanol must be used because sulfuric acid dehydrates the ethanol — removes a water molecule from the formula of ethanol. If the ethanol contained water, such as 95% ethanol, the sulfuric acid would not react with ethanol.
2. Diethyl ether is slightly polar; there are two electron pair groups on the oxygen. These can form hydrogen-bonding interactions with water — and thus dissolve in the water. Ethene is hydrophobic; it is non-polar.
3. pungent odor
4. Ethanol minus a water molecule is ethene.

Chapter 21, Experiment 1. Thermal cracking of polyethylene — an alternative preparation of ethene

1. Carbon-carbon double bond; each carbon bonded to two hydrogen atoms.

Chapter 21, Experiment 2. Going bananas!

1. Green fruit is easier to ship because it is less prone to bruising.
2. Yes, the fruit in the bag without the ethene would ripen with adequate time.

Chapter 21, Experiment 3. Reaction with bromine-water

1. ethene: See Experiment 1, Question 1; 2-bromoethanol: one carbon bonded to two hydrogen atoms and bromine and to the other carbon with a single bond. The second carbon is bonded to two hydrogen atoms and an OH group.
2. 1,2-dibromoethene looks the same as bromoethanol, but with a second bromine instead of the OH.
3. Bromine water could be used to estimate the purity of a gas mixture containing ethene several ways. Reacting the sample with excess bromine, and measuring the absorbance would allow one to calculate the molarity of bromine and the number of moles. From this, the number of moles of ethene originally present could be determined.

Chapter 21, Experiment 4. Ethene reacts with potassium permanganate

1. The color change was purple to colorless (or brown).
2. Ethene gas forms water soluble ethandiol (ethylene glycol).
3. Ethylene glycol is water soluble because of hydrogen-bonding.
4. Hydroxyl functional groups were added.

Chapter 21, Experiment 5. Flammability of ethene

1. $C_2H_4(g) + O_2(g) \rightarrow 2 CO_2(g) + 2 H_2O(g)$
2. As ethene burns, it produces hot gases that are lighter than air and leave the test tube. Cool air from the room takes its place and brings more oxygen into the test tube.
3. You probably would not see a ring of fire. Perhaps an explosion would occur instead.

Chapter 21, Experiment 6. Ethene rockets

2. The limewater test should have produced white calcium carbonate.
3. The fuel mixture for these rockets should have been among the best in this book.
4. Large energy of activation.

Chapter 21, Experiment 7. Solubility of ethene in alcohol and Henry's law

1. Dissolving takes place at the interface between the liquid and the gas. Agitating the mixture speeds this up.
2. The amount of gas that dissolves in a solvent increases with an increase in pressure. Reducing the pressure causes dissolved gases to come out of solution. When pressurized spray cans are discharged, gases (the propellants) come out of the solution (paints, deodorants, etc.).
3. The plot of solubility of a gas (y-axis) vs. pressure (x-axis) will have a positive slope and a y-intercept of 0.

Chapter 21, Experiment 8. Ethene reactions with chlorine

2. The first reaction was slow and followed a low-energy pathway. The second reaction had a much higher energy of activation, but once it was initiated, it was self-sustaining because it was exothermic.
3. See the answer to Question 2.

Chapter 22, Experiments with Methane.

1. 2 g sodium acetate (MM 82) corresponds to 0.024 mol $NaC_2H_3O_2(s)$ and 2 g sodium hydroxide (MM 40) corresponds to 0.05 mol $NaOH$. The reaction stoichiometry is 1:1 sodium acetate is the limiting reagent. The other product obtained is $Na_2CO_3(s)$. The theoretical yield is 0.024 moles of $CH_4(g)$.
2. The theoretical yield is 600 mL $CH_4(g)$ at 298 K and standard pressure.
4. The density of air is 1.19 g/L. The density of methane is 0.65 g/L. The ratio of densities, $density_{methane}/density_{air}$ is 0.55:1.

Chapter 22, Experiment 1. Products of combustion

2. If the flow is too fast, the flame is blown out.
3. Not all combustion reactions form carbon dioxide. Carbon monoxide can be formed if carbon-based substances are combusted. If non-carbon-based substances are being combusted, such as iron, oxides are formed.
4. The two products of combustion of hydrocarbon materials are carbon dioxide and water.
5. It would burn hotter and brighter.

Chapter 22, Experiment 2. How a Bunsen burner works

1. If the flow of methane is increased, the flame becomes larger.

2. If the flow of air is increased, the flame becomes shorter and brighter (hotter).
3. Combustion required fuel (methane) and oxygen (air).
4. Bunsen's burner design efficiently mixes fuel and air in proportions that can be adjusted.
5. The hot flame vaporizes a small amount of glass. Sodium in this glass colors the flame.

Chapter 22, Experiment 3. Flame chemistry

1. The inner cone (pyrolysis zone) of the flame contains the fuel.
2. The outer, combustion zone is the hotter of the two.
3. The experiment should still work if you drew gases from the outer cone. In the outer cone, fuel/air mixtures are reacting; siphoned-off samples should be combustible.

Chapter 22, Experiment 4. Burned rings in paper

2. The pyrolysis zone is the inner cone and the combustion zone is the outer cone.
3. The burned circle on the card would be darker and more focused if the card were held closer to the burner. Near the tip of the flame, the diameter of the circle would be smallest.

Chapter 22, Experiment 5. Window screen provides thermal insulation

2. Yes, you could ignite the gas on both sides of the screen at the same time.
3. Not all of the fuel is consumed if the screen is in the way. Gases that pass through the screen are cooled to below their combustion temperature, but are, nevertheless, combustible.

Chapter 22, Experiment 6. Density of methane: Lighter-than-air methane bubbles.

2. The smaller bubbles drop because the buoyancy of the methane in the smaller bubble is not adequate to lift the weight of the soap film.
3. For example, a bubble with a diameter of 3 cm has a volume of 4.2 mL. At 298 K and standard pressure, this corresponds to 1.7×10^{-4} mol.
4. 140 J

Chapter 22, Experiment 7. Burning methane. Controlling gas flow by density

1. Experiment 6
2. The ring of fire moved down the tube as air moves into the tube to replace the hot gases that are leaving to their lighter density.

Chapter 22, Experiment 8. Explosive mixture of methane/air

1. Again, the density of methane plays a role. Lighter-than-air methane leaves through the hole in the foil where it burns.
2. The flow of CH_4 should be out through the hole. Oxygen, in the form of air, comes into the flame above the foil **and** into the apparatus through the bottom opening. CO_2 and H_2O are formed in the flame.
3. The methane suddenly exploded when the air content in the apparatus reached a level where it formed an explosive mixture with the methane present.

Chapter 22, Experiment 9. Bubble domes

1. The bubble film is strong and as long as the tubing is moist, it can be inserted through the soap film.
2. Yes, you could poke your finger, moistened with dish soap, through the soap film.
3. The surface tension is reduced.

Chapter 22, Experiment 10. Big water thrasher

1. The explosion shot water out of the 2-L container because the energy released has to go somewhere. Trapped under water, the energy is imparted to the water.
2. The volume of the film container is about 40 mL. Suppose that 20 mL methane and 30 mL oxygen were used to form the mixture, and 40 mL of the mixture is used. Oxygen is the limiting reagent ($\text{CH}_4 + 2 \text{O}_2$). The film canister contains 24 mL oxygen or 9.8×10^{-4} mol at 298 K and standard pressure. This should produce 394 J of energy.
3. Explosions require an explosive mixture of fuel and air and a spark or open flame.

Chapter 22, Experiment 11. Reaction with chlorine: Formation of soot

1. This reaction is an oxidation/reduction. CH_4 is being oxidized by Cl_2 , the oxidizing agent. Chlorine is being reduced by methane.
2. Soot is evident. Testing the gas produced shows that it is acidic, whereas air is not.
3. Under milder conditions, chloromethane, CH_3Cl might be formed

Chapter 23. Experiments with nitrous oxide

1. Only $\text{NH}_4\text{NO}_3(\text{s})$ matters in terms of calculating the theoretical yield (equation: $\text{NH}_4\text{NO}_3(\text{aq}) \longrightarrow \text{N}_2\text{O}(\text{g}) + 2 \text{H}_2\text{O}(\text{l})$). 1.0 g $\text{NH}_4\text{NO}_3(\text{s})$ (MM 80) corresponds to 0.0125 mol NH_4NO_3 .
2. 306 mL
3. The rate of $\text{N}_2\text{O}(\text{g})$ formation is extremely slow.

Chapter 23, Experiment 1. Wooden splint test for nitrous oxide

1. Oxygen.
2. Nitrous oxide's ability to support combustion used is used in race-car engines.
3. $4 \text{N}_2\text{O}(\text{g}) + \text{CH}_4(\text{g}) \longrightarrow \text{CO}_2(\text{g}) + 2 \text{H}_2\text{O}(\text{g}) + 4 \text{N}_2(\text{g})$
4. Nitrous oxide is not as reactive as oxygen. You may have had more difficulty getting the splint to ignite.

Chapter 23, Experiment 2. Nitrous oxide forms explosive mixtures with hydrogen

1. $\text{N}_2\text{O}(\text{g}) + \text{H}_2(\text{g}) \longrightarrow \text{H}_2\text{O}(\text{g}) + \text{N}_2(\text{g})$
2. $\text{DH}_f = -324 \text{ kJ}$
3. $\text{H}_2(\text{g})$ is the limiting reagent.
4. 106 J

Chapter 23, Experiment 3. Nitrous oxide rockets

2. These ratios reflect the reaction stoichiometries.

- hydrogen/oxygen, methane/oxygen, hydrogen/chlorine, ethyne/oxygen, ethene/oxygen, nitrous oxide/hydrogen, carbon monoxide and oxygen
- $\text{N}_2\text{O} + \text{O}_2$ would not make a reasonable fuel mixture for a rocket because both are oxidants.

Chapter 23, Experiment 4. Solubility in water and oil

- In both cases, a homogeneous solution was formed.
- Nitrous oxide dissolves to a small extent in water because it is not completely non-polar; it has some polar characteristics. It dissolves in vegetable oil because it is basically a non-polar molecule.

Chapter 23, Experiment 5. Magnesium burns in nitrous oxide

- $\text{N}_2\text{O}(\text{g}) + \text{Mg}(\text{s}) \longrightarrow \text{MgO}(\text{s}) + \text{N}_2(\text{g})$
- oxidation reduction
- Magnesium reacts better with air because of the oxygen present.
- oxygen

Chapter 23, Experiment 6. Enlarged candle flame of nitrous oxide

- The flame is the brightest in oxygen, then nitrous oxide, then air.
- The candle burned the longest in oxygen, followed by nitrous oxide.
- Nitrous oxide is an oxidant, not a fuel.
- (a) Air is to candle wax as nitrous oxide is to methane.
(b) Nitrous oxide is to oxygen as hydrogen is to methane.

Chapter 25. Ozone

- Ozone is produced by electrical discharge – electrical motors, lightening, and so on.
- See page 401.
- Ozone has an electron pair on the central oxygen atom making it polar, like water and thus more soluble. Molecular O_2 is non-polar and is less soluble in water.
- $\text{O}_3(\text{g}) + 2 \text{Ag}(\text{s}) \rightarrow \text{O}_2(\text{g}) + \text{O}^{2-}(\text{g}) + 2 \text{Ag}^+(\text{g})$
- $\text{O}_3(\text{g})$ and $\text{O}_2(\text{g})$ are allotropes.
- $2 \text{H}^+ + 2 \text{e}^- \rightarrow \text{H}_2$ $E_{\text{red}}^\circ = 0.00 \text{ V}$
- $2 \text{H}_2\text{O} \rightarrow \text{O}_2 + 4 \text{H}^+ + 4 \text{e}^-$ $E_{\text{ox}}^\circ = -1.23 \text{ V}$ (major product)
- $2 \text{H}_2\text{O} \rightarrow \text{O}_2 + 2 \text{H}_2$ $E_{\text{tot}} = -1.23 \text{ V}$

Chapter 25, Experiment 1. Traditional test for ozone

- It turned blue
- oxidized
- It would turn blue where the ozone comes out. Diffusion may turn more of the paper blue, however ozone is more dense than air (MM = 48 g/mol) vs. ~ 29 g/mol for air, so density would have the ozone travel downward after leaving the generator tube.

Chapter 25, Experiment 2. Quantitative determination of ozone

1. flow = 0.0023 g/hr
2. Concentration (ppm) = $0.40 \times 10^4 = 400$ ppm
3. 18.1 days

Chapter 25, Experiment 3. Reaction between food coloring and ozone

2. The reaction is zero order in terms of blue food coloring. The rate law is rate = k. In order to determine k (the slope), we would need the relationship between absorbance and [Blue] from $A = a \times b \times c$; we do not know a, the molar absorptivity.
3. The blue color "gone" by 300 s
4. Blue
5. The orange solution would lose its red component first and would appear yellow: Orange → Yellow → Colorless

Chapter 25, Experiment 4. Ozone reacts with natural and artificial colors

1. The quantity of colored substance is too great for the ozone to react with in a short period of time. Basically, the moles of chromophore \gg moles ozone.
3. The chromophore is being oxidized to non-colored products
4. If the artificial colors are listed on the package label, one can perform experiments such as Experiment 3 in order to determine which artificial colors are the most stable with respect to ozone degradation.

Chapter 25, Experiment 5. Ozone reacts with rubber

1.
$$\text{CH}_3-\text{C}=\text{C}-\text{CH}_3 \quad \text{CH}_3-\overset{\overset{\text{O}}{\text{O}}}{\text{C}}-\text{C}-\text{CH}_3$$
2. When the rubber band is stretched, it gives off heat, suggesting that new bonds are formed. Try this with a thick rubber band: touch it to your forehead, note the temperature and then stretch it and immediately feel the new temperature with your forehead. It should feel noticeable warmer. The new bonds cannot be covalent bonds, so they are intermolecular forces, which apparently can more easily form when the polymers are stretched and not as tangled. The web offers a variety of alternative explanations, and I am not sure which one is correct.
3. Everything made of rubber is susceptible to ozone degradation. Tires are the most common example.

Chapter 25, Experiment 6. Ozone kills mold on cheese.

3. Ozone kills bacteria much like hydrogen peroxide does. Ozone does not persist in any form, so there is no danger of ingesting ozone, however, the degradation products of the treated substance are still present. For example, mold treated with ozone is dead, but is not gone. It may affect the flavor of the cheese.
4. Ozone stops bacterial growth, but cannot undo the biological degradation reactions that it has caused.

Chapter 25, Experiment 7. Ozone discolors colored office paper

1. A color change indicates a chemical reaction has taken place. In the case of blue food coloring, Blue → Colorless, however Colorless is still one or more chemical substances. In the case of green food coloring, Green is a mixture of two food colorings, Blue and Yellow, both pure chemical substances. In this case, as we saw in Experiment 3, Yellow is more stable than Blue in the presence of ozone.
2. If a substance is slow to react, its activation energy for the reaction is higher and the rate of the reaction will be slower.
3. Probably the same one that reacts the fastest with ozone

Chapter 25, Experiment 8. Ozone reacts with wood fibers.

1. Lignin
3. Ozone degradation causes one chemical to become one or more chemicals, nominally $A \rightarrow B$. With food coloring, A is colored and B is not. In the case of wood fibers, A has a pale color (the color of wood), while B has a dark color.

Chapter 25, Experiment 9. Ozone kills bacteria in water.

1. 24000 nmol or 24 μ mol
2. Bacteria are ubiquitous in nature
4. Chlorine. Chlorine is easier to transport, ozone has to be generated on the spot. Ozone does not damage the atmosphere, whereas chlorine degrades the ozone layer.
5. Ozone the limiting reagent at this low rate of production

Chapter 26. Experiments with silane.

1. 0.20 g magnesium silicide corresponds to 0.0026 mol $Mg_2Si(s)$.
2. 0.0026 mol silane is expected
3. 64 mL
5. The water in the upper syringe excludes air and allows an air-free transfer.
6. What is the purpose of the water in the “silane dispenser” provides an air-trap to prevent silane from coming into contact with air.

Chapter 26, Experiment 1. Silanes react with air

1. The bursts of fire first appear inside the pipet, because there was a little air present. After that was consumed. The fire appeared at the tip of the pipet instead.
2. $SiH_4(g) + 2 O_2(g) \longrightarrow SiO_2(s) + 2 H_2O(l)$
3. Heat and light were the forms of energy release.
4. Silane and air could not be stored together in the same syringe.
5. Silicon dioxide is a network covalent molecule most familiar to us in the form of quartz. Carbon dioxide is molecular.

Chapter 26, Experiment 2. Silane reacts with oxygen

1. These maneuvers are carried out under water to prevent exposure to air.
2. Use the same general approach, but use some liquid other than water, such as oil.
3. The white smoke is actually silicon dioxide, $SiO_2(s)$.
4. The test tube not explode because pressure can be released by pushing the water downward.
5. The silicon-oxygen single bond energy, 452 kJ/mol, is relatively large.

6. CO_2 is sp^2 hybridized and SiO_2 is sp^3 hybridized. The carbon-oxygen bond order is 2 and the silicon-oxygen bond order is 1.

Chapter 26, Experiment 3. Silane reacts with chlorine

2. Cleaner, purer products in higher yields are usually obtained when reactions are slower and less exothermic.
3. Chlorine is very reactive.
4. $\text{SiCl}_4(\text{l}) + 4 \text{H}_2\text{O}(\text{l}) \longrightarrow \text{Si}(\text{OH})_4(\text{l}) + 4 \text{HCl}(\text{aq})$
6. SiCl_4 is sp^3 hybridized with a silicon-chlorine bond order of 1.

Chapter 26, Experiment 4. Thermal decomposition of silane

2. $\Delta H = -32 \text{ kJ}$, slightly exothermic.
3. It is necessary to provide heat because this reaction is kinetically slow.
4. Hydrogen is not soluble in water.
5. Hydrogen does not ignite because there is no oxidant such as oxygen present.

Chapter 26, Experiment 5. Reaction with aqueous potassium hydroxide

2. change in volume
3. The volume of gas in the syringe increased.
4. Oxidation numbers for SiH_4 : Si = -4 and H = +1; H_2O : H = +1 and O = -2; SiO_2 : Si = +4 and O = -2; and H_2 : H = 0. The oxidizing agent is H_2O and the reducing agent is SiH_4 .

Chapter 27. Experiments with Hydrogen Sulfide

1. 0.22 g $\text{ZnS}(\text{s})$ corresponds to 0.0023 mol $\text{ZnS}(\text{s})$
2. 0.0023 mol $\text{H}_2\text{S}(\text{g})$
3. 55 mL $\text{H}_2\text{S}(\text{g})$
4. The volume produced is usually less than the theoretical yield.
5. The densities of air and hydrogen sulfide are 1.19 g/L and 1.39 g/L, respectively. The ratio of densities, $\text{density}_{\text{hydrogen sulfide}}/\text{density}_{\text{air}}$, is 1.17:1. Hydrogen sulfide is 17 % heavier than air.
6. 0.19 moles $\text{H}_2\text{S}/\text{kg}$

Chapter 27, Experiment 1. Hydrogen sulfide is slowly oxidized

1. There are several possibilities for why this reaction is slow. The rate constant is likely small. Hydrogen sulfide does not react rapidly with oxygen in the air. In water, oxygen is not present in a large amount. As it is depleted, more oxygen from the air dissolves. This process takes time and is hastened by agitating the surface of the water. The water acts as a catalyst for the reaction, however, so that is one thing that would increase the rate of the reaction. The reaction would have been faster if the test tube contained oxygen instead of air.
2. The milky-colored precipitate formed is sulfur. It is formed in extremely small particles that scatter so much light that the solid appears white instead of yellow.
3. After hydrogen sulfide enters the atmosphere or natural waters, it slowly oxidizes to elemental sulfur.

4. Either filter the solid or allow it to settle and then melt it and allow it to solidify. Grind the solid and it will look like the familiar yellow sulfur.

Chapter 27, Experiment 2. Hydrogen sulfide is a weak acid

1. The pH change is caused by a physical property. Acid dissociation in water is a physical property of substances we call “acids”.
2. (a) When comparing two weak acids, the one with the *larger* K_a is the stronger of the two and will be *more* dissociated than the weaker acid. (b) When comparing two weak acids, the one with the *smaller* pK_a is the stronger of the two. (c) If solutions of identical concentration were made of these two weak acids, the one with the larger K_a would exhibit the *lower* pH. (d) The weak acid with the larger K_a will have the *smaller* pK_a .
3. $HS^-(aq)$ is the conjugate base of hydrogen sulfide.
4. $pH = 3.85$. Using the quadratic equation was not necessary.
5. Hydrogen selenide should be acidic.

Chapter 27, Experiment 3. Reaction between hydrogen sulfide and aqueous sodium hydroxide

1. The volume of gas decreased.
2. $Na_2S(aq); NaHS(aq) + NaOH(aq) \longrightarrow Na_2S(aq) + H_2O(l)$
3. The conjugate base produced, $HS^-(aq)$ is more stable than water's conjugate base, $OH^-(aq)$ due to its larger size. This allows it to accommodate the charge. The same rationale is used to explain why HCl is a stronger acid than HF.
4. Aqueous hydrogen sulfide and aqueous sodium hydroxide react to form aqueous sodium hydrogen sulfide and water.

Chapter 27, Experiment 4. Hydrogen sulfide burns in oxygen with a howling blue flame

1. The combustion reaction in oxygen is brighter.
2. Sulfur is easily oxidized. In sulfur dioxide, one of the most familiar compounds of sulfur, the sulfur has an oxidation number of +4. Oxygen has no positive oxidation states.
3. $S(s) + O_2(g) \longrightarrow SO_2(g)$
4. $2 SO_2(g) + O_2(g) \longrightarrow 2 SO_3(g); SO_3(g) + H_2O \longrightarrow H_2SO_4(g)$
5. Alkali metal oxides and alkaline earth oxides are base anhydrides. Nitric oxide and carbon dioxide are non-metal oxides that are also acid anhydrides.

Chapter 27, Experiment 5. Reaction between hydrogen sulfide and sulfur dioxide yields elemental sulfur

2. The droplets of water dissolved both reactants and brought them into proximity so they could react.
3. Oxidation numbers for $H_2S(g)$: H = +1 and S = -2; SO_2 : S = +4 and O = -2; S_8 : S = 0; and H_2O : H = +1 and O = -2.
4. The experiment was designed with hydrogen sulfide as the limiting reagent to minimize odor.
5. The gases are both more dense than air and stay in the test tube.
6. Burn one third of the hydrogen sulfide in air to produce sulfur dioxide, then combine the rest of the hydrogen sulfide with sulfur dioxide. See the reaction stoichiometry, page 419.

Chapter 27, Experiment 6. Metal sulfide precipitation reactions

1. All were precipitations.
2. The reactions take place at the surface because that is where the $\text{H}_2\text{S}(\text{g})$ first encounters the aqueous metal ion solutions.
3. $\text{Cu}^{+2}(\text{aq}) + \text{H}_2\text{S}(\text{aq}) + 2 \text{H}_2\text{O}(\text{l}) \longrightarrow \text{CuS}(\text{s}) + 2 \text{H}_3\text{O}^{+}(\text{aq})$; $\text{Pb}^{+2}(\text{aq}) + \text{H}_2\text{S}(\text{aq}) + 2 \text{H}_2\text{O}(\text{l}) \longrightarrow \text{PbS}(\text{s}) + 2 \text{H}_3\text{O}^{+}(\text{aq})$; $2 \text{Bi}^{+3}(\text{aq}) + 3 \text{H}_2\text{S}(\text{aq}) + 6 \text{H}_2\text{O}(\text{l}) \longrightarrow 2 \text{Bi}_2\text{S}_3(\text{s}) + 6 \text{H}_3\text{O}^{+}(\text{aq})$
4. The solution becomes slightly acidic so the pH decreases. ; molar solubility = 1.0×10^{-28} mol/L
5. $\text{CdS}(\text{s})$: $K_{\text{sp}} = 1.0 \times 10^{-28}$; molar solubility = 1.0×10^{-14} mol/L; $\text{CuS}(\text{s})$: $K_{\text{sp}} = 8.5 \times 10^{-45}$; molar solubility = 9.2×10^{-23} mol/L; $\text{PbS}(\text{s})$: $K_{\text{sp}} = 7 \times 10^{-29}$; molar solubility = 8.4×10^{-15} mol/L

Chapter 27, Experiment 7. Oxidation of metal sulfides

2. (a) $\text{CuS}(\text{s}) + 4 \text{H}_2\text{O}_2(\text{aq}) \longrightarrow \text{Cu}^{+2}(\text{aq}) + \text{SO}_4^{2-}(\text{aq}) + 4 \text{H}_2\text{O}(\text{l})$; (b) $\text{PbS}(\text{s}) + 4 \text{H}_2\text{O}_2(\text{aq}) \longrightarrow \text{PbSO}_4(\text{s}) + 4 \text{H}_2\text{O}(\text{l})$
3. The sulfide ion was oxidized.
4. Excess hydrogen sulfide oxidizes with oxygen in air. See Experiment 1: $2 \text{H}_2\text{S}(\text{aq}) + \text{O}_2(\text{aq}) \longrightarrow 2 \text{S}(\text{s}) + 2 \text{H}_2\text{O}(\text{l})$
5. The most common reason is to produce an aqueous solution of the metal ion.

Appendix H. Reference to Original Articles that Appeared in *Chem13 News**, 1996 - Present.

For ordering information for Chem13 News, see Appendix E.

1996:

- "Microscale Gas Chemistry: Generating Gases in Large Syringes" Mattson, B. M., *Chem13 News*, **251**, October, 1996.
- "Microscale Gas Chemistry, Part 2. Experiments with Carbon Dioxide" Mattson, B. M., *Chem13 News*, **252**, November, 1996.
- "Microscale Gas Chemistry, Part 3. Experiments with Hydrogen" Mattson, B. M., *Chem13 News*, **253**, December, 1996.

1997:

- "Microscale Gas Chemistry, Part 4. Experiments with Oxygen and Nitrogen" Mattson, B. M.; Lannan, J. *Chem13 News*, **254**, January, 1997.
- "Microscale Gas Chemistry, Part 5. Experiments with Nitrogen Oxides" Mattson, B. M.; Lannan, J. *Chem13 News*, **255**, February, 1997.
- "Microscale Gas Chemistry, Part 6. Experiments with Ammonia" Mattson, B. M., *Chem13 News*, **256**, March, 1997.
- "Microscale Gas Chemistry, Part 7. Experiments with Hydrocarbons" Mattson, B. M.; Anderson, M.; Nguyen, J; Harrison, B., *Chem13 News*, **257**, April, 1997. See * below.
- "Microscale Gas Chemistry, Part 8. Experiments with Hydrogen Sulfide" Mattson, B. M.; Anderson, M.; Nguyen J; Lannan, J., *Chem13 News*, **258**, May, 1997.
- "Microscale Gas Chemistry, Part 9. Experiments with Sulfur Dioxide" Mattson, B.; Anderson, M.; Nguyen J; Lannan, J., *Chem13 News*, **259**, September, 1997.
- "Microscale Gas Chemistry, Part 10. Experiments with Chlorine" Mattson, B. M.; Harrison, B.; Lannan, J., *Chem13 News*, **260**, October, 1997.

1998:

- "Microscale Gas Chemistry, Part 11. Experiments with Hydrogen Chloride" Mattson, B. M.; Catahan, R.; Vaitkus, R., *Chem13 News*, **266**, April, 1998.

1999:

- "Microscale Gas Chemistry, Part 12. Experiments with Carbon Monoxide" Mattson, B. M., Anderson, M. P., Catahan, R., Bansal, M., Khandhar, P., Mattson, A., Rajani, A., Obendrauf, V., Vaitkus, R., *Chem13 News*, **274**, March, 1999.
- "Microscale Gas Chemistry, Part 13. Experiments with Ethene" Mattson, B. M., Hulce, M., Fujita, J., Anderson, M. P., Catahan, R., Bansal, M., Khandhar, P., Mattson, A., Rajani, A., Worth, L., Obendrauf, V., *Chem13 News*, **277**, September, 1999.

2000:

- "Microscale Gas Chemistry, Part 14. Experiments with Silane" Mattson, B. M., Anderson, M. P., Nguyen, J, Bansal, M., *Chem13 News*, **281**, January, 2000.
- "Microscale Gas Chemistry, Part 15. Experiments with Methane" Mattson, B. M., Catahan, R., Nguyen, J., Patel, A., Khandhar, P., Mattson, A. and Anand Rajani, A., *Chem13 News*, **284**, April, 2000.

2001:

"Microscale Gas Chemistry, Part 16. Generating Gases in a Microwave Oven" Mattson, B. M., Pound, K., **Chem13 News**, **292**, March, 2001.

"Microscale Gas Chemistry, Part 17. Molar Mass of Gas Determination. A chemistry laboratory experiment;" Mattson, B. M., Greimann, J., Dedhia, R., Saunders, E., **Chem13 News**, **295**, September, 2001.

2002:

"Microscale Gas Chemistry, Part 18. Experiments with Nitrous Oxide," Mattson, B., Sullivan, P., Fujita, J., Pound, K., Cheng, W., Eskestrand, S., Obendrauf, V., **Chem13 News**, **299**, January, 2002.

"Microscale Gas Chemistry, Part 19. Formation of Elemental Iron. A laboratory experiment in a pipette," Mattson, B., Saunders, E., **Chem13 News**, **303**, May, 2002.

"Microscale Gas Chemistry, Part 20. Reversible Oxidation of Metallic Copper. A laboratory experiment in a pipette." Mattson, B., Saunders, E., DiSapio, C., Hamilton, R., **Chem13 News**, **304**, September, 2002.

"Microscale Gas Chemistry, Part 21. The Limiting Reagent. A chemistry laboratory experiment," Mattson, B., Eskestrand, S., Meyer, A., **Chem13 News**, **305**, October, 2002.

2003:

"Demonstrating Heterogeneous Gas Phase Catalysis with the Gas Reaction Catalyst Tube," Mattson, B., Fujita, J., Catahan, R., Cheng, C., Greimann, J., Hoette, T., Khandhar, P., Mattson, A., Rajani, A., Sullivan, P., Perkins, R" accepted for publication **Journal of Chemical Education**, 2003, **00**, 0000.

"Microscale Gas Chemistry, Part 22. Ziploc[®] Bags for Temporary Gas Storage and Transfer"; Bruce Mattson, B.; and Meyer, A., **Chem13 News**, **311**, April, 2003.

"Microscale Gas Chemistry, Part 23. Improved Chlorine Preparation", Mattson, B.; Eskestrand, S.; and Mordechai Livneh, M., **Chem13 News**, **312**, May, 2003.

"Microscale Gas Chemistry, Part 24. Percent Composition of Calcium Carbonate in Tums[®]. A chemistry laboratory experiment.", Mattson, B.; and Saunders, E., tentatively: **Chem13 News**, **313**, September, 2003.

"Microscale Gas Chemistry, Part 25. Mystery Gas", Mattson, B.; and Pottenbaum, D.; tentatively: **Chem13 News**, **314**, October, 2003.

"Microscale Gas Chemistry, Part 27. **Carbonated Beverages — Priestley's Soda-water**", Mattson, B.; tentatively: **Chem13 News**, **316**, December, 2003.

2004:

"Microscale Gas Chemistry, Part 26. **Carbonated Beverages — Priestley's Soda-water**", Mattson, B.; Saunders, E.; Sconzo, P.; **Chem13 News**, **324**, 6-11, November, 2004.

"Microscale Gas Chemistry, Part 25. Mystery Gas", Mattson, B.; and Pottebaum, D.; tentatively: **Chem13 News**, **318**, February, 2004, 10 - 11.

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"Microscale Gas Chemistry, Part 28. **Barometric Pressure without a Barometer**", Mattson, B.; and Snipp, R.; Eskestrand, S.: **Chem13 News**, **326**, 12 - 14, Jan, 2005.

*Part 7 was titled "Experiments with Hydrocarbons" but largely describes the chemistry of ethyne. Ethene and methane are described in subsequent articles.

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