STUDYWALKS · HOMESCHOOL PACK · FREE TO PRINT

Chemistry Lab Guide

Twelve home labs keyed to the units of Chemistry, from density and stoichiometry to calorimetry, gas laws, titration and a lemon battery, using kitchen and hardware-store materials.

Twelve hands-on labs keyed to the units of the StudyWalks Chemistry course. They use kitchen and hardware-store materials, so several classic school labs (strong acids and bases, flame tests, organic solvents) are replaced here with safe household versions. A kitchen scale that reads to 0.1 g, or better a 0.01 g pocket scale, makes the quantitative labs much more precise.

Safety first

Read this page with your student before the first lab, and reread each lab's own safety notes before starting it.

- **An adult supervises every lab** and reads the whole lab, safety notes included, before it starts. An adult handles knives, the electric kettle, hot water and anything hot. No lab in this guide uses a stove, a burner, a candle or any other open flame.
- **Wear eye protection** (safety goggles marked ANSI Z87.1, sold inexpensively at hardware stores) in every lab with liquids, heat or anything that can splash or snap.
- **Wear gloves** (disposable nitrile, or rubber kitchen gloves) when a lab's safety box says so: iodine, hydrogen peroxide, cabbage juice, steel wool, cold salted ice, the sealed mold bags and outdoor work.
- **Never taste, eat or drink** anything used in a lab, even when it began as food. No food or drinks in the work area. Label every container, and wash borrowed kitchen tools well (or keep a set just for labs).
- **Never mix unknown chemicals** or household products, and use only the materials a lab lists. Never use bleach, ammonia, drain, oven or toilet-bowl cleaners in these labs: some combinations release toxic gases.
- **Ventilation:** work in a well-aired room, with a window open if you can.
- **Dress for it:** closed-toe shoes, long hair tied back, loose sleeves rolled up, old clothes or an apron.
- **Allergies:** before each lab, check for allergies to anything it uses, such as latex (balloons), iodine, foods, mold, plants, pollen or insect stings. Substitute or skip.
- **Cleanup and disposal:** the liquids in these labs can go down the sink with running water; food scraps go in the trash or compost; follow each lab's own disposal note. Wipe the surfaces and wash hands with soap after every lab.
- **If something goes wrong:** for a splash in the eye, rinse with running water for 15 minutes and get medical help; cool a burn under cool running water. In the U.S., Poison Help is 1-800-222-1222. Stop if anything unexpected happens.
- **Young children and pets** stay out of the work area. Balloons, marbles and small parts are choking hazards.
- **No cleaning products.** Never use bleach, ammonia, drain cleaner, oven cleaner, toilet-bowl cleaner or any other cleaning product in these labs, and never mix household products to "see what happens". Some combinations release toxic gases.
- **Closed containers.** Never put a cap or cork on a container in which a gas is being made. Pressure can build until it bursts. The one closed-system lab (Lab 2) uses a balloon, which stretches.

- **No flames, no stove.** No lab here uses a candle, a match, a lighter, a burner or the stove. Hot water comes from an electric kettle and is handled by an adult (Labs 4 and 10).
- **No mercury.** The spectrum lab (Lab 5) uses LED, incandescent and phone light only: never fluorescent tubes or compact fluorescent (CFL) bulbs, which contain mercury, and never a broken bulb of any kind.

What this guide leaves out on purpose: open flames of every kind (no candles, matches, lighters or burners) and stove-top heating; mercury (no fluorescent tubes or CFL bulbs, and never a broken bulb of any kind); hydrogen peroxide stronger than the 3% sold for first aid; strong acids and bases; flammable solvents (which is why there is no DNA-extraction lab: it needs rubbing alcohol); dissection; opening anything that has grown mold; and anything that makes a toxic gas.

How to use this guide

1. Do each lab after the student has finished the course unit it names. The unit is listed with each lab.
2. Read the whole lab together first: purpose, safety, materials and procedure. Gather everything before starting.
3. The student records everything in a lab notebook as the lab happens (see the notebook guide), then writes a lab report using the template.
4. Grade the report with the parent rubric and log the hours.

Each lab ends with **Go further**, a second investigation that changes one variable. For a lab every week of the school year, add these and investigations of your student's own design.

Lab hours

There is no single national rule for how much lab work makes a "lab science", and rules differ by state. Many homeschool families plan a lab most weeks, about 1 to 2 hours of hands-on work plus time to write it up. The twelve labs here come to roughly 25 to 40 hours with write-ups; the Go further investigations add about as much again. Use the lab-hours log at the end. Lab time can also be logged as offline study hours for the course in StudyWalks, which counts them separately from lesson time. You decide what counts and what credit to grant.

The labs

Lab	Chemistry unit	Time
1. Density and Measurement Uncertainty	Unit 1: Essential Ideas: Matter, Measurement, and Uncertainty	1 hour
2. Conservation of Mass: Open and Closed Systems	Unit 2: Atoms, Molecules, and Ions	45 minutes
3. Limiting Reactant: Baking Soda and Vinegar	Unit 3: Composition, Solutions, and Stoichiometry	1.5 hours
4. Specific Heat of a Metal	Unit 4: Thermochemistry	1.5 hours
5. Seeing Spectra with a CD Spectroscope	Unit 5: Electronic Structure and Periodic Properties	1 hour
6. Charles's Law: Gas Volume and Temperature	Unit 7: Gases	1.5 hours

Lab	Chemistry unit	Time
7. Paper Chromatography of Marker Inks	Unit 8: Liquids, Solids, Solutions, and Colloids	1 hour
8. Freezing-Point Depression: Ice and Salt	Unit 8: Liquids, Solids, Solutions, and Colloids	45 minutes
9. Reaction Rate: Temperature and Surface Area	Unit 9: Kinetics	1 hour
10. Red Cabbage pH Indicator	Unit 10: Equilibrium: Fundamentals, Acids and Bases, and Solubility	1 hour
11. Titration: How Much Acid Is in Vinegar?	Unit 3: Composition, Solutions, and Stoichiometry; Unit 10: Equilibrium: Fundamentals, Acids and Bases, and Solubility	1.5 hours
12. A Lemon Battery	Unit 11: Thermodynamics and Electrochemistry	1 hour

Keeping a lab notebook

- Use a bound notebook (a composition book works) and write in pen. Number the pages and keep a table of contents on the first page.
- Start each lab on a new page with the date, the lab title and anyone who helped.
- Record raw data the moment you take it, with units. Never copy it from scrap paper later.
- Do not erase or tear out pages. Cross out a mistake with a single line so it can still be read, and note why.
- Sketch what you see and label the sketch. Write down anything unexpected and any change you made to the procedure.
- End each entry with a one-line summary of the result. The parent initials and dates the entry.

Lab 1. Density and Measurement Uncertainty

CHEMISTRY Unit 1: Essential Ideas: Matter, Measurement, and Uncertainty · About 1 hour

Purpose

Measure the densities of water, vegetable oil and corn syrup, report them with sensible significant figures, and compare them with accepted values.

Materials

- Kitchen scale (0.1 g if possible)
- A measuring cup marked in mL, or an inexpensive 100 mL plastic graduated cylinder
- Water, vegetable oil and light corn syrup
- Food coloring and a tall clear glass
- Graph paper and a calculator

Safety

Adult present for the whole lab · **Wear:** nothing special · **Disposal:** Water and syrup down the sink with running water; wipe oil up with paper towels for the trash, never down the drain.

- Never taste the liquids. Oil spills are slippery; wipe them up right away.

Procedure

1. Record the scale's smallest reading and the cylinder's smallest marking. These set the uncertainty of each measurement.
2. Put the empty cylinder on the scale and zero it. Add water to about 25 mL, reading the bottom of the curved surface (the meniscus) at eye level. Record the volume and the mass.
3. Add water to about 50 mL and then about 75 mL, recording volume and mass each time.
4. Repeat for oil and for corn syrup, cleaning and drying the cylinder between liquids.
5. Calculate each density (mass ÷ volume). Then graph mass against volume for each liquid; the slope of each line is its density.
6. Predict the order, then slowly pour corn syrup, colored water and oil, in that order, into the tall glass to make a density column.

Data

Liquid	Volume (mL)	Mass (g)	Density (g/mL)

Analysis questions

1. Which digit in each of your measurements is estimated? How many significant figures should each density have?
2. Using your three trials per liquid, describe the precision of your results. Using the accepted values, describe their accuracy.
3. Why is the slope of a graph often better than a single measurement?
4. Did the density column match your predictions? Predict whether a grape would float on the water layer, then test it.
5. Density is an intensive property. Explain what that means, using your data.

What to expect. Water comes out close to 1.00 g/mL; vegetable oil about 0.91 to 0.93 g/mL; light corn syrup about 1.3 to 1.4 g/mL (brands differ). Results within a few percent are good with kitchen tools.

Typical sources of error. Misreading the meniscus; syrup and oil clinging to the cylinder; the scale's resolution; air bubbles in the syrup.

Go further. Measure the density of salt water at several concentrations and graph density against grams of salt.

Lab 2. Conservation of Mass: Open and Closed Systems

CHEMISTRY Unit 2: Atoms, Molecules, and Ions · About 45 minutes

Purpose

Test the law of conservation of mass by weighing a reaction in an open container and in a closed one.

Background

Baking soda (sodium hydrogen carbonate) reacts with the acetic acid in vinegar: $NaHCO_3 + CH_3COOH \rightarrow NaCH_3COO + H_2O + CO_2$. The carbon dioxide is a gas.

Materials

- Kitchen scale (0.1 g if possible)
- A 500 mL plastic bottle and a balloon (non-latex if needed)
- A funnel, a cup and a small square of paper
- Baking soda, vinegar, measuring spoons and a measuring cup

Safety

Adult present for the whole lab · **Wear:** safety goggles · **Disposal:** Liquid down the sink with running water; balloon in the trash.

- Wear eye protection; vinegar stings eyes.
- Never put the cap on the bottle during the reaction. Use only the balloon.

Procedure

1. Open system: pour 100 mL of vinegar into the cup. Put 1 teaspoon of baking soda on the paper. Weigh the cup and the paper together.
2. Tip the baking soda into the vinegar. When the fizzing stops, weigh the cup and the paper together again.
3. Closed system: pour 100 mL of vinegar into the bottle. Use the funnel to put 1 teaspoon of baking soda into the balloon.
4. Stretch the balloon over the bottle's mouth, letting it hang to the side so no powder falls in. Weigh the whole setup.
5. Lift the balloon so the powder drops into the vinegar. When the fizzing stops, weigh the setup again with the balloon still on.
6. Do each setup three times.

Data

System	Trial	Mass before (g)	Mass after (g)	Change (g)
Open				
Open				
Open				
Closed				
Closed				
Closed				

Analysis questions

1. Which system lost mass, and where did that mass go?
2. Did the closed system's mass change? An inflated balloon pushes aside air, and the air pushes up on it a little (about 1.2 g for every liter of balloon). How could that affect your reading?
3. Write the balanced equation and check that each kind of atom is conserved.
4. Why must a law about mass be tested in a closed system?

What to expect. The open cup loses mass as carbon dioxide escapes, often about 1 to 2 g with these amounts. The closed setup changes by about the scale's resolution, or reads slightly lower because of the air pushing up on the inflated balloon.

Typical sources of error. Powder that drops in early; a leaking balloon seal; scale resolution; the air's push on the balloon.

Go further. Repeat the open system with twice as much baking soda and predict the mass lost before you measure it.

Lab 3. Limiting Reactant: Baking Soda and Vinegar

CHEMISTRY Unit 3: Composition, Solutions, and Stoichiometry · About 1.5 hours

Purpose

Find the amount of baking soda that just uses up a fixed amount of vinegar, by measuring the mass of carbon dioxide lost, and compare it with the amount predicted from the balanced equation.

Background

Vinegar labeled "5% acidity" holds about 5 g of acetic acid (molar mass 60.05 g/mol) per 100 mL, so 50 mL holds about 2.5 g, or 0.042 mol. The reaction is 1:1, so it needs about 0.042 mol of baking soda (84.01 g/mol), about 3.5 g, and can release at most about 1.8 g of carbon dioxide (44.01 g/mol).

Materials

- Kitchen scale (a 0.01 g pocket scale is best; 0.1 g works)
- Vinegar labeled 5% acidity and baking soda
- 6 cups and 6 small squares of paper
- A measuring syringe (no needle), measuring cup or graduated cylinder
- Graph paper

Safety

Adult present for the whole lab · **Wear:** safety goggles · **Disposal:** Liquids down the sink with running water.

- Wear eye protection; vinegar stings eyes.
- Add the powder slowly so the foam does not overflow.

Procedure

1. Measure 50 mL of vinegar into each of six cups.
2. Weigh out baking soda portions of 0.5, 1.0, 2.0, 3.0, 4.0 and 5.0 g, each on its own paper.
3. For each cup: weigh the cup and its paper of baking soda together (starting mass). Add the powder slowly, swirl, wait 5 minutes until the fizzing stops, swirl again, and weigh the cup and paper together (final mass).
4. Mass lost = starting − final. This is the carbon dioxide that escaped.
5. Note whether undissolved powder is left at the bottom of each cup.
6. Graph carbon dioxide lost (y-axis) against baking soda added (x-axis).

Data

Baking soda (g)	Starting mass (g)	Final mass (g)	CO2 lost (g)	Powder left over?
0.5				
1.0				
2.0				
3.0				
4.0				
5.0				

Analysis questions

1. Where does your graph level off? Which reactant is limiting before that point, and which after?
2. Using the molar masses, calculate the baking soda and carbon dioxide expected for 50 mL of your vinegar. Compare them with your graph.
3. Calculate the percent yield of carbon dioxide at the plateau. Why is it below 100 percent?
4. Before the plateau, the graph's slope should be about 0.52 g of CO_2 per gram of baking soda. Where does 0.52 come from?

What to expect. Carbon dioxide lost rises in proportion to the baking soda, then levels off near 3.5 g of baking soda. The measured carbon dioxide is usually somewhat below the theoretical 1.8 g, because some carbon dioxide stays dissolved in the liquid.

Typical sources of error. Dissolved carbon dioxide; droplets carried out with the foam; scale resolution; vinegar that is not exactly 5 percent.

Go further. Repeat with 100 mL of vinegar and predict where the new plateau will be before you measure.

Lab 4. Specific Heat of a Metal

CHEMISTRY Unit 4: Thermochemistry · About 1.5 hours

Purpose

Check heat balance by mixing hot and cold water, then use a foam-cup calorimeter to measure the specific heat of steel.

Background

Heat lost by the hot metal equals heat gained by the water (if none escapes): $m_{metal}\, c_{metal}\, (T_{metal} - T_{final}) = m_{water} \times 4.18 \times (T_{final} - T_{water})$, with c in J/g·°C. With 150 g of steel at about 95 °C dropped into 100 g of water at 20 °C, expect a rise of about 10 °C.

Materials

- 2 nested foam cups with a cardboard lid
- A digital kitchen thermometer (0.1 °C if possible)
- Kitchen scale
- About 150 g of steel washers or nuts on a string loop
- An electric kettle, a vacuum flask or insulated jug (a thick ceramic mug works), and tongs (adult)
- A measuring cup

Safety

Adult present for the whole lab · **Wear:** safety goggles · **Disposal:** Let the water cool, then pour it down the sink; dry the washers.

- Just-boiled water and hot metal cause burns. An adult handles the kettle, the hot water and the hot washers, and uses the tongs. There is no stove in this lab.
- Wear eye protection. Work on a tray away from the table edge, and keep young children and pets away while the water is hot.

Procedure

1. Part A: weigh 100 g of cold water into the foam cup and record its temperature. Weigh 100 g of warm tap water and record its temperature. Predict the final temperature, then pour the warm water into the cup, stir, and record the final temperature.
2. Part B: weigh the washers. An adult puts them in the flask or jug and covers them with just-boiled water from the electric kettle. After 5 minutes, measure the water's temperature beside the washers just before the transfer; that is the washers' starting temperature (usually about 85 to 95 °C).
3. Weigh 100 g of room-temperature water into a clean foam cup and record its temperature.
4. The adult lifts the washers out with the tongs, lets drops fall off for 2 seconds, and lowers them into the cup. Put on the lid.
5. Stir gently with the thermometer and record the highest steady temperature.
6. Calculate the specific heat of steel. Repeat for a second trial.

Data

Quantity	Trial 1	Trial 2
Mass of water (g)		
Mass of metal (g)		
Water starting temperature (°C)		
Metal starting temperature (°C)		
Final temperature (°C)		
Heat gained by water (J)		
Specific heat of metal (J/g·°C)		

Analysis questions

1. In Part A, how close was the final temperature to your prediction? What does a difference tell you?
2. Compare your specific heat for steel with the accepted value for iron, 0.449 J/g·°C. Calculate the percent error.
3. Hot water carried into the cup on the washers would make your answer too high or too low? What about heat lost while moving the washers?
4. Why does water make a good calorimeter liquid?

What to expect. Part A lands close to the average of the two temperatures, slightly lower because the cup and air absorb some heat. Steel gives about 0.45 to 0.50 J/g·°C; results within 10 to 20 percent are typical with foam cups.

Typical sources of error. Heat lost to the cup and air; water carried over on the washers (raises the result); the washers cooling during transfer, or the hot water cooling between the reading and the transfer (lowers it); thermometer response time.

Go further. Repeat with aluminum or brass hardware and compare with accepted values (aluminum 0.897, brass about 0.38 J/g·°C).

Lab 5. Seeing Spectra with a CD Spectroscope

CHEMISTRY Unit 5: Electronic Structure and Periodic Properties · About 1 hour

Purpose

Use a CD, DVD or diffraction-grating glasses to compare the spectra of safe light sources, see that a hot filament gives a continuous spectrum while LEDs and screens give separate bands, and connect the bright lines of gases to electrons moving between energy levels using published reference spectra.

Materials

- Diffraction grating glasses ("rainbow glasses", inexpensive) or a whole CD or DVD (do not cut or break it)
- Two index cards and tape (to make a narrow slit)
- Light sources: an LED bulb in a lamp, a phone flashlight, a phone screen showing white and then pure red, green and blue, an incandescent bulb in a lamp if you have one, and colored LED holiday lights (optional)
- Reference line spectra of hydrogen, helium and neon: the atomic-spectra figures in the course's free textbook (OpenStax *Chemistry 2e*) or any printed chart
- Colored pencils

Safety

Adult present for the whole lab · **Wear:** nothing special · **Disposal:** Nothing to throw away. A broken bulb is cleaned up by an adult, never by the student.

- Never look at the Sun or a laser pointer, directly or through a grating or CD. It can permanently damage your eyes.
- No fluorescent tubes or compact fluorescent (CFL) bulbs: they contain mercury. Never use a broken or cracked bulb of any kind; if a bulb breaks, leave it for an adult to clean up.
- View the phone flashlight from at least 1 m away and only for a few seconds at a time. An incandescent bulb gets hot: leave it in its lamp and do not touch it.
- Do not cut or break the CD or DVD; broken discs have sharp edges.

Procedure

1. Darken the room except for the light source. Tape two index cards side by side with a gap of about 1 mm and hold the slit in front of the source, or view a small, bright source from across the room.
2. Look at the source through the grating glasses, or hold the CD near your eye, label side down, and tilt it until you see a spectrum beside the source.
3. For each source, record whether the spectrum is a smooth rainbow (continuous) or has separate bands, and sketch the colors in order with colored pencils.
4. Compare at least four sources. On the phone screen, switch between white, red, green and blue and note which bands remain.
5. Compare your sketches with the reference line spectra of hydrogen, helium and neon. Note how a line spectrum differs from what you saw.

Data

Light source	Continuous or bands?	Colors seen, in order	Sketch or notes

Analysis questions

1. Which sources gave continuous spectra and which gave separate bands?
2. Look at the reference spectrum of hydrogen. Explain its bright lines using electrons moving between energy levels. Why does each element have its own pattern?
3. A white LED usually shows a strong blue band plus a broad yellow-green-red band. What does that suggest about how it makes white light?
4. How do astronomers use line spectra to learn what stars are made of?
5. Why does a hot solid, such as a bulb filament, give a continuous spectrum?

What to expect. An incandescent bulb gives a smooth continuous spectrum. White LEDs (bulbs and phone flashlights) show a strong blue band plus a broad yellow band. A phone screen showing white splits into separate red, green and blue bands; showing pure red, only the red band remains. Colored LEDs each give one fairly narrow band. None of these is a true line spectrum: the lamps that show sharp lines (gas-discharge tubes and fluorescent lamps) are left out of this guide, so the reference charts stand in for them. Hydrogen's visible lines are red, blue-green and two violet.

Typical sources of error. A slit that is too wide blurs the bands; room light washing out the spectrum; a CD gives fainter, curved spectra than grating glasses.

Go further. Build a box spectroscope: tape a whole CD at an angle inside a cereal box, with a narrow slit at one end and a viewing hole at the other, and compare how sharp the bands look. Or compare colored LEDs and record the width of each band by color.

Lab 6. Charles's Law: Gas Volume and Temperature

CHEMISTRY Unit 7: Gases · About 1.5 hours

Purpose

Measure how the volume of a gas changes with temperature at nearly constant pressure, and extrapolate to estimate absolute zero.

Background

For a balloon that is close to a sphere, volume = $C^3 \div 59.2$, where C is the circumference (59.2 is 6 times pi squared). Kelvin temperature = °C + 273.

Materials

- A round balloon (non-latex if needed)
- String and a ruler, or a sewing tape measure
- A kitchen thermometer
- A freezer, a bowl of ice water and a bowl of warm tap water (about 45 °C)
- Tongs and a towel
- Graph paper

Safety

Adult present for the whole lab · **Wear:** nothing special · **Disposal:** Balloon in the trash.

- The warm bath should feel comfortably warm, not hot.
- Balloons are a choking hazard for young children.

Procedure

1. Blow the balloon up about two-thirds full so it has room to expand, and tie it.
2. Measure its circumference at the widest point at room temperature, and record the room temperature.
3. Put it in the freezer for 20 minutes. Record the freezer temperature and measure the balloon immediately after taking it out.
4. Hold it under the ice water with tongs for 10 minutes, then pat it dry and measure.
5. Hold it in the warm water for 10 minutes, then pat it dry and measure.
6. Calculate each volume. Graph volume (y-axis) against temperature in °C (x-axis), draw the best-fit line, and extend it to where volume would be zero.

Data

Condition	Temperature (°C)	Temperature (K)	Circumference (cm)	Volume (cm³)
Freezer				
Ice water				
Room				
Warm water				

Analysis questions

1. How does volume change with temperature? Is the relationship closer to direct proportion in °C or in K?
2. At what temperature does your line reach zero volume? Compare with −273 °C.
3. Explain your results using the kinetic-molecular theory.
4. The balloon's rubber pushes back as it stretches, so the pressure is not perfectly constant. How might that affect your results?
5. Why do car-tire pressures drop on a cold morning?

What to expect. Volume rises with temperature. The changes are modest (between a freezer and a warm bath, the circumference changes by only a few centimeters), so careful measuring matters. The line usually points toward somewhere near −273 °C, but balloon data often miss by 50 to 100 °C or more.

Typical sources of error. The balloon warms or cools while being measured; a rough circumference reading; the balloon is not a perfect sphere; rubber tension.

Go further. Seal the tip of a large plastic syringe, set the plunger at half volume, and repeat the baths, reading the volume from the syringe scale.

Lab 7. Paper Chromatography of Marker Inks

CHEMISTRY Unit 8: Liquids, Solids, Solutions, and Colloids · About 1 hour

Purpose

Separate the dyes in water-soluble markers, calculate Rf values, and explain the results with the attractions between the dyes, water and paper.

Materials

- White coffee filters (or paper towel), cut into strips about 2 cm by 12 cm
- Washable markers in several colors, including black and brown
- A pencil, a ruler and tape
- Tall glasses and water

Safety

Adult present for the whole lab · **Wear:** nothing special · **Disposal:** Strips in the trash; water down the sink.

- No special hazards. Use washable markers only; permanent markers need solvents that are not used in this guide.

Procedure

1. Draw a pencil line 2 cm from the bottom of each strip.
2. Put a small dot of one marker on the line, let it dry, and dot it again two or three times to concentrate it.
3. Tape the strip to a pencil laid across a glass, so the strip hangs into water about 1 cm deep. The water must start below the dot.
4. When the water has risen to about 1 cm from the top (10 to 20 minutes), remove the strip and mark the water's highest point (the solvent front) in pencil.
5. When dry, measure the distance from the pencil line to the center of each colored band, and to the solvent front.
6. Calculate Rf = distance traveled by the dye ÷ distance traveled by the water.

Data

Marker	Band color	Dye distance (cm)	Solvent distance (cm)	Rf

Analysis questions

1. Which markers were mixtures of dyes? How do you know?
2. Why is the starting line drawn in pencil rather than pen?
3. Explain why some dyes travel farther than others, in terms of their attraction to the water (mobile phase) and to the paper (stationary phase).
4. Would a dye's Rf change if the strip ran longer? Why is Rf more useful than distance?
5. Describe how chromatography could show which of three markers wrote a note.

What to expect. Black and brown washable markers usually separate into several colors, often blue, pink or red, and yellow. Some single-color markers stay as one band. Every Rf is between 0 and 1.

Typical sources of error. Dots that are too large; the water touching the dot directly; the strip touching the glass wall; uneven paper.

Go further. Repeat with salt water (1 teaspoon of salt in a cup of water) as the solvent and compare the Rf values.

Lab 8. Freezing-Point Depression: Ice and Salt

CHEMISTRY Unit 8: Liquids, Solids, Solutions, and Colloids · About 45 minutes

Purpose

Measure how dissolving salt lowers the temperature of an ice-water mixture, and connect the result to colligative properties.

Materials

- Ice (crushed works best)
- Table salt and table sugar
- 2 bowls or large cups
- A thermometer that reads to at least −20 °C
- A measuring cup, a tablespoon and a kitchen scale
- Gloves or a towel

Safety

Adult present for the whole lab · **Wear:** gloves · **Disposal:** Let the ice melt; salt and sugar water down the sink with running water.

- Salted ice gets cold enough to injure skin within minutes. Do not hold it in a bare hand or against skin; use gloves or a towel.
- Salt water stings cuts and eyes.

Procedure

1. Check the thermometer: stir it in a cup of ice and a little water for 2 minutes. It should read 0 °C. Note any difference and correct your later readings by it.
2. In each bowl, mix 2 cups of ice with 1/2 cup of cold water. Stir for 2 minutes and record the temperature.
3. In the first bowl, add 2 tablespoons of salt, stir for 2 minutes, and record the lowest steady temperature.
4. Add salt 2 tablespoons at a time, stirring and recording after each, up to 8 tablespoons. Weigh a level tablespoon of salt once so you can report grams.
5. In the second bowl, repeat with the same masses of sugar.
6. Graph temperature against grams added, one line for salt and one for sugar.

Data

Added (tablespoons)	Grams added	Salt bowl (°C)	Sugar bowl (°C)
0			
2			
4			
6			
8			

Analysis questions

1. Why did the salted mixture fall below 0 °C? Explain what happens as ice melts.
2. Freezing-point depression depends on the number of dissolved particles. Sodium chloride splits into two ions; sugar stays as one molecule and has a much larger molar mass. Use both facts to explain the difference between your two lines.
3. Why do road crews spread salt in winter, and why does it stop working in very cold weather?
4. Why did the temperature eventually stop dropping?

What to expect. Ice water reads about 0 °C. With plenty of salt the mixture usually reaches about −10 to −20 °C; the lowest possible for salt and ice is about −21 °C. Sugar lowers the temperature much less per gram.

Typical sources of error. Not stirring enough; reading before the temperature settles; kitchen thermometers that are less accurate below zero.

Go further. Freeze a sealed bag of juice in the salted ice and time it against one in plain ice. (This is the chemistry behind hand-cranked ice cream. Nothing from the lab is eaten.)

Lab 9. Reaction Rate: Temperature and Surface Area

CHEMISTRY Unit 9: Kinetics · About 1 hour

Purpose

Measure how temperature and particle size change the rate of a reaction, using effervescent tablets.

Background

Effervescent tablets hold a solid acid (often citric acid) and sodium bicarbonate. They react only once they dissolve in water, releasing carbon dioxide.

Materials

- Effervescent tablets, all the same brand (plain vitamin C or antacid type)
- 6 or more clear cups
- Water at three temperatures: ice-cold, room and warm tap water (about 45 °C)
- A thermometer, a timer, a measuring cup
- A zip-top bag and a spoon for crushing

Safety

Adult present for the whole lab · **Wear:** safety goggles · **Disposal:** Water down the sink; unused tablets back in their package, out of children's reach.

- Some effervescent tablets contain aspirin or other medicine. Nobody drinks the water; pour it down the sink. Keep the tablets away from young children.
- Wear eye protection; the fizz can splash.

Procedure

1. Pour 200 mL of water at each temperature into its own cup and record the temperature.
2. Drop a whole tablet into each cup and time how long it takes until the tablet has completely disappeared and the fizzing stops. Do two trials at each temperature.
3. At room temperature, compare a whole tablet, a tablet broken into quarters, and a tablet crushed to powder in the bag.
4. Calculate a rate for each as 1 ÷ time (per second).

Data

Condition	Temperature (°C)	Trial 1 (s)	Trial 2 (s)	Mean (s)	Rate (1/s)
Whole, ice-cold					
Whole, room					
Whole, warm					
Quarters, room					
Powder, room					

Analysis questions

1. Graph rate against temperature. Describe the pattern.
2. Use collision theory to explain the effect of temperature.
3. Why does crushing the tablet speed up the reaction?
4. For many reactions, a 10 °C rise roughly doubles the rate. Is that roughly true for your data?
5. What was hardest to judge, and how did it affect your times?

What to expect. Warm water is fastest and ice water several times slower. Powder reacts faster than quarters, and quarters faster than a whole tablet.

Typical sources of error. Judging the exact end; water temperatures drifting toward room temperature; tablets that are not identical.

Go further. Test water with a spoonful of sugar dissolved in it, or a cup with the same water but stirred constantly.

Lab 10. Red Cabbage pH Indicator

CHEMISTRY Unit 10: Equilibrium: Fundamentals, Acids and Bases, and Solubility · About 1 hour

Purpose

Make an acid-base indicator from red cabbage, build a color scale, and classify household liquids as acidic, neutral or basic.

Materials

- About 1/4 of a red cabbage, a knife (adult) and hot water from an electric kettle (adult)
- A strainer and a jar
- Clear cups or a white ice-cube tray, and a teaspoon
- Samples: vinegar, lemon juice, clear soda pop, tap water, milk, baking soda solution (1 teaspoon in 1/2 cup water), a crushed antacid tablet in 1/2 cup water, a few drops of dish soap in water
- Optional: inexpensive pH test strips

Safety

Adult present for the whole lab · **Wear:** safety goggles and gloves · **Disposal:** Samples down the sink with running water; cabbage in the compost or trash.

- An adult handles the knife and the hot water.
- Wear eye protection and disposable gloves (the juice stains).
- Test only the listed samples. Never test drain cleaner, oven cleaner, bleach, ammonia, toilet-bowl cleaner or any other cleaning product: they can burn skin and eyes, and some release toxic gases when mixed.
- Cabbage juice stains. Pour samples down the sink with running water; compost or throw away the cabbage.

Procedure

1. Chop the cabbage, cover it with hot water, and let it steep for 15 to 30 minutes. Strain the purple liquid into the jar and let it cool. This is the indicator.
2. Put 2 tablespoons of indicator in each cup.
3. Predict acidic, neutral or basic for each sample. Then add 1 teaspoon of each sample to its own cup and record the color.
4. Estimate each pH from the color scale below (or check with pH strips).
5. Neutralization: to a cup of indicator with 1 teaspoon of vinegar, add baking soda solution a few drops at a time, swirling, and count the drops until the color turns purple.

Data

Approximate cabbage-indicator colors

Color	Approximate pH
Red	about 1 to 3
Pink to red-purple	about 4 to 5
Purple	about 6 to 7
Blue	about 8
Blue-green to green	about 9 to 11
Yellow	about 12 and above (not tested here)

Results

Sample	Prediction	Color	Estimated pH	Strip pH (optional)

Analysis questions

1. What is an indicator, and why can a natural indicator only estimate pH?
2. A liquid at pH 3 has 100 times the concentration of hydronium ions of one at pH 5. Explain why, using what each pH unit means.
3. Which samples were acidic, and which basic? Did any surprise you?
4. Write the neutralization equation for vinegar and baking soda. What gas forms?
5. Tap water and distilled water are not always exactly pH 7. Suggest a reason.

What to expect. Vinegar, lemon juice and soda pop turn red or pink (about pH 2 to 4). Water and milk stay purple (about pH 6 to 7; milk is slightly acidic). Baking soda solution turns blue to blue-green (about pH 8 to 8.5), and the antacid is similar or a little greener. Soaps vary. The colors shift with how strong the indicator is.

Typical sources of error. Judging colors by eye; indicator strength; colored samples (use clear ones).

Go further. Make indicator paper by soaking coffee filter strips in the indicator and drying them, then test the same samples.

Lab 11. Titration: How Much Acid Is in Vinegar?

CHEMISTRY Unit 3: Composition, Solutions, and Stoichiometry; Unit 10: Equilibrium: Fundamentals, Acids and Bases, and Solubility · About 1.5 hours

Purpose

Use a titration with a weighed amount of baking soda and the cabbage indicator to measure the acetic acid in vinegar, and compare the result with the label.

Background

Baking soda and acetic acid react 1:1. If you know the moles of baking soda, the vinegar needed to use it all up contains the same number of moles of acetic acid. Measuring the vinegar by mass on the scale is more precise than reading a volume.

Materials

- Kitchen scale (0.01 g best, 0.1 g works)
- Baking soda and vinegar labeled 5% acidity
- Cabbage indicator from Lab 10
- A cup for the reaction and a small cup of vinegar with a dropper or needle-free syringe
- Water and a spoon

Safety

Adult present for the whole lab · **Wear:** safety goggles · **Disposal:** Everything down the sink with running water.

- Wear eye protection; vinegar stings eyes.
- Pour everything down the sink with running water afterward.

Procedure

1. Weigh 1.00 g of baking soda (2.0 g if your scale reads to 0.1 g) into the reaction cup. Add 50 mL of water and 2 tablespoons of indicator. It turns blue-green.
2. Put the small cup of vinegar with its dropper on the scale and record the mass.
3. Add vinegar to the reaction cup in small squirts, swirling. It fizzes. As the color moves toward purple, slow down to a few drops at a time, swirling for 30 seconds after each.
4. Stop at the first color that stays pink-purple for 30 seconds. This is the end point.
5. Weigh the vinegar cup and dropper again. The difference is the mass of vinegar used.
6. Calculate: moles of baking soda = mass ÷ 84.01; moles of acetic acid = the same; grams of acetic acid = moles × 60.05; volume of vinegar (mL) = mass of vinegar ÷ 1.01; grams of acetic acid per 100 mL = grams ÷ volume × 100.
7. Do three trials.

Data

Trial	Baking soda (g)	Vinegar cup before (g)	After (g)	Vinegar used (g)	Acetic acid (g per 100 mL)
1					
2					
3					
Mean					

Analysis questions

1. Why can you say the moles of acetic acid equal the moles of baking soda?
2. Why swirl after each addition? (Think about the carbon dioxide.)
3. How close were your three trials? How close was your mean to the label?
4. Adding too much vinegar past the end point is called overshooting. Would it make your result too high or too low?
5. What is the difference between the end point you saw and the equivalence point?

What to expect. Results usually land between about 4 and 6 g of acetic acid per 100 mL for vinegar labeled 5 percent. Overshooting is common and makes the result read low.

Typical sources of error. Overshooting; dissolved carbon dioxide blurring the color change; judging color by eye; scale resolution.

Go further. Test a different brand, or apple cider vinegar (its color makes the end point harder to see; record how you handled it).

Lab 12. A Lemon Battery

CHEMISTRY Unit 11: Thermodynamics and Electrochemistry · About 1 hour

Purpose

Build galvanic cells from lemons, zinc and copper, measure their voltage, and connect cells in series.

Background

Zinc is oxidized at the zinc electrode ($Zn \rightarrow Zn^{2+} + 2e^-$), and electrons flow through the wire to the copper. At the copper surface, hydrogen ions from the lemon juice gain electrons; the copper itself is not used up.

Materials

- 4 lemons (or potatoes)
- 4 galvanized (zinc-coated) nails or screws
- 4 pieces of thick copper wire or copper strip (U.S. pennies made since 1982 are mostly zinc and only thinly plated with copper, so wire is better)
- Alligator-clip leads
- An inexpensive digital multimeter
- A small red LED

Safety

Adult present for the whole lab · **Wear:** gloves · **Disposal:** Lemons in the trash, never eaten; keep the nails, wire and LED.

- Do not eat the lemons afterward; metal dissolves into the juice. Throw them away and wash hands.
- Nails and wire ends are sharp.
- Never connect anything in this lab to a wall outlet or household wiring.

Procedure

1. Roll each lemon on the table to free the juice. Push a nail and a copper wire into it about 3 cm apart, not touching.
2. Set the multimeter to DC volts. Touch the red lead to the copper and the black lead to the zinc. Record the voltage.
3. Connect two lemons in series: copper of the first to zinc of the second. Measure across the free copper and free zinc. Repeat with three and four lemons.
4. Connect the LED to the four-lemon battery (the longer leg goes to the copper end). Look at it in a dark room.
5. Try other combinations: a potato, a cup of salt water, or a steel paper clip in place of the zinc nail.

Data

Setup	Voltage (V)	Notes
1 lemon		
2 lemons in series		
3 lemons in series		
4 lemons in series		
4 lemons + LED		
Potato		
Salt water		
Steel clip and copper		

Analysis questions

1. Which metal is the anode, and which is oxidized? Which way do electrons flow in the wire?
2. What happens to the voltage when cells are added in series? Explain.
3. The standard zinc-copper cell gives 1.10 V. Why does a lemon cell give less?
4. Why is the LED dim even when the voltage is high enough to light it?
5. Galvanized nails are coated with zinc to protect the steel. Use this lab to explain how the zinc protects the iron.

What to expect. One lemon cell gives about 0.9 to 1.0 V. Cells in series add, so four lemons give roughly 3.5 to 4 V. The current is very small, so the LED glows only faintly and is easiest to see in the dark. Swapping zinc for steel lowers the voltage.

Typical sources of error. Electrodes touching inside the lemon; poor clip contact; electrodes pushed in to different depths.

Go further. Measure how the voltage of one cell changes as the electrodes are moved closer together or farther apart.

Lab report template

Print one copy per lab, or copy the headings into the notebook.

Student _____ Lab number and title _____ Date _____

Course unit _____

Question or purpose

Hypothesis (If … then … because …)

Variables: independent, dependent, and those kept the same

Materials and procedure (summarize; note any changes you made)

Safety steps followed

Data (attach tables and graphs) and calculations, with units

Results in words

Conclusion: claim, evidence from the data, reasoning

Sources of error, how each affected the result, and improvements

Hands-on minutes

Write-up minutes

Rubric score (of 20)

Parent initials

Parent grading rubric

Score each criterion from 4 to 1 and add them (20 points in all). How the score becomes a grade, and how lab grades combine with course work, is the parent's decision.

Criterion	4	3	2	1
Safety and procedure	Followed every safety rule without reminders; set up, carried out and cleaned up carefully; noted any change to the procedure.	Followed safety rules and the procedure with few reminders.	Needed several reminders; some steps rushed or skipped.	Unsafe or careless work, or procedure not followed.
Data	Complete, organized tables with units; repeated trials; raw data recorded as it was taken; clear observations.	Mostly complete, with units; minor gaps.	Incomplete or disorganized; units missing.	Little usable data.
Analysis	Calculations correct and shown with units; graphs titled, labeled and scaled well; all questions answered fully.	Minor calculation or graph errors; questions mostly answered.	Several errors; graph or questions incomplete.	Calculations, graph or questions largely missing.
Conclusion	Answers the purpose with a clear claim, evidence from the data and scientific reasoning; compares with the expected result.	Claim and evidence given; reasoning thin.	Claim stated with little evidence.	No real conclusion.
Error and improvement	Names specific sources of error, says how each would change the result, and proposes a better method.	Names specific sources of error.	Only general errors ("human error").	Not addressed.

Score sheet

Lab	Safety	Data	Analysis	Conclusion	Error	Total (20)
1. Density and Measurement Uncertainty						
2. Conservation of Mass: Open and Closed Systems						
3. Limiting Reactant: Baking Soda and Vinegar						
4. Specific Heat of a Metal						
5. Seeing Spectra with a CD Spectroscope						
6. Charles's Law: Gas Volume and Temperature						
7. Paper Chromatography of Marker Inks						
8. Freezing-Point Depression: Ice and Salt						
9. Reaction Rate: Temperature and Surface Area						
10. Red Cabbage pH Indicator						
11. Titration: How Much Acid Is in Vinegar?						

Lab	Safety	Data	Analysis	Conclusion	Error	Total (20)
12. A Lemon Battery						

Lab-hours log

Log every lab and investigation, including Go further and the student's own. Print as many copies as you need.

Date	Lab or investigation	Unit	Hands-on min	Write-up min	Running hours	Parent

Total lab hours _____ Parent signature _____ Date _____

Beyond the kitchen

Virtual labs. Free simulations from universities and science organizations cover gas laws, circuits, genetics, titrations and more, including experiments that are not safe at home. They are good practice and a good way to repeat a lab with many more trials. Use them alongside hands-on labs; whether a simulation counts toward a lab requirement is your call under your state's rules.

A formal lab course. If you want lab work with equipment a home cannot have, such as microscopes, analytical balances and fume hoods, look for lab classes through a homeschool co-op, your state or local homeschool association, a community college's dual-enrollment office, or a science center. StudyWalks does not endorse or partner with any provider; ask each one what is covered and how work is documented.

StudyWalks is a curriculum and study-tool provider, not a school. In a homeschool the parent decides the courses, the credit and the grades, and keeps the records. Check your own state's homeschool rules. This pack is free to print and copy for your family or homeschool group; please do not sell it or repost the file.

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