High School Chemistry

First edition | 9 units | 36 lessons | One credit with lab | A video for every lesson

Worksheets A and B with keys | Quick reference: **Chemistry** · **Periodic table** · **Organic chemistry**

Where this leads

Fields of study: Chemistry, Chemical engineering, Pharmacy, Medicine and nursing, Materials science, Food science, Environmental and water science, Agriculture.

Jobs and careers (trades, licenses and degree paths):

- DEGREE Chemist (industry, quality control, research)
- DEGREE Chemical engineer
- DEGREE Pharmacist
- DEGREE Physician, nurse or physician assistant
- DEGREE Materials or metallurgical engineer
- 2-YR / CERT Laboratory technician
- 2-YR / CERT Water and wastewater treatment operator
- TRADE Welder or metal finisher (plating, coatings)
- TRADE HVAC and refrigeration technician
- LICENSE Pest-control applicator
- FIELD Brewer, baker or food technologist

Key topics and where they're used:

- **The mole and stoichiometry**: pharmacy dosing, manufacturing, recipes scaled to a factory
- **Gases**: diving, HVAC, airbags, welding gases
- **Acids, bases and solutions**: water treatment, pools, soil testing, medicine
- **Energy and rates**: fuels, batteries, engines, food preservation

Paths are examples, not promises. Most careers need further training, an apprenticeship, a license or a degree.

FIRST EDITION

Course guide

This is a full-year, college-preparatory chemistry course with lab for students in grades 10–11. It covers the whole introductory core: measurement and significant

figures, atomic structure, electrons and the periodic table, bonding and molecular shape, chemical reactions, the mole and stoichiometry, gases, solutions, acids and bases, and energy, rates, equilibrium and nuclear chemistry. Every lesson has key terms, worked examples with units on every line, a practice set with answers, a lab, two printable worksheets (A and B) with separate answer keys, and a history note. Each unit ends with a 100-point test, and there is a midterm after Unit 5 and a final after Unit 9, all with printable keys for the parent.

Chemistry was discovered by people, one careful experiment at a time, and the course is built around those experiments: Boyle's doubts about the four elements, Lavoisier weighing everything before and after, Dalton's atoms, Thomson's cathode rays, Rutherford's gold foil, Mendeleev leaving gaps for elements no one had found, Sørensen inventing pH in a brewery laboratory. The labs use a kitchen scale, syringes, vinegar, baking soda, red cabbage and Epsom salt; no glassware set is required. Many of the founders of chemistry, Boyle, Faraday and Dalton among them, were men of faith who saw an orderly creation worth measuring, and the course says so where it is true.

> Videos were chosen lesson by lesson from trusted teaching channels: Tyler DeWitt (patient, clear introductions), Periodic Videos from the University of Nottingham (real demonstrations), Melissa Maribel (short worked problems), Khan Academy, The Organic Chemistry Tutor, Cognito, Bozeman Science, Science Buddies and Reactions from the American Chemical Society. Each video was matched to its lesson by topic, channel and length and confirmed to play embedded; as with any video, a parent preview is wise. Most run under 10 minutes; a few longer ones have a "when" note saying which part to watch. Watch with the notebook open and pause to work each example before the presenter does.

Start with Lesson 1.1 →

Who it is for

Written for students in grades 10–11 who have completed Algebra I (solving equations, ratios, scientific notation). Algebra II at the same time helps with

logarithms for pH (taught in place in Lesson 8.4) and the gas-law algebra. Biology is not required first. Students work fairly independently, reading each lesson and working its examples before checking answers, with a parent present for labs that use heat, flame or hot liquids.

How to use this course

1. **The weekly rhythm.** Each lesson takes four sessions of about 45 minutes: (1) warm-up, video and core ideas; (2) worked examples and the first practice problems; (3) the lab and lab notebook; (4) the rest of the practice set and Worksheet A, checked and corrected. Worksheet B is a second set for review or a retake. Each unit ends with a test day, and the midterm and final each have a review day. At five sessions a week the course fills about 32 weeks.
2. **Story, then chemistry, then use.** Read the core ideas in order. The story and "how they figured it out" explain why the idea exists; then comes the chemistry, then where it is used today.
3. **Units on every line.** Write the equation, substitute the numbers with their units, cancel units along with numbers, and box the answer with its unit and sensible significant figures.
4. **Practice and worksheets.** Problems are marked easy, medium and hard. Mark your work against the answers the same day and rework every miss. The printable worksheets have separate keys for the parent; numeric answers on worksheets were computed by program, not typed by hand.
5. **The lab notebook.** Use a bound notebook for the whole year. For every lab write the date, question, hypothesis with a reason, procedure, a data table with units, a graph where it fits, a comparison with the expected value (percent error) and a paragraph on sources of error. Record what actually happened: "a just weight is his delight" (Proverbs 11:1).
6. **Tests and exams.** Each unit test is worth 100 points and mixes short answers, calculations, a lab-data question and a short essay. Calculators and a periodic table are allowed; formulas are not given, so build a one-page formula sheet for each unit as you go.
7. **Safety.** Read the safety note before every lab. Wear splash goggles and nitrile gloves. Use only 3% hydrogen peroxide. Never mix bleach with ammonia, vinegar

or any acid. Keep isopropyl alcohol away from flames. Flame and hot-liquid labs need an adult present.

Scope and sequence

UNIT 1 · WEEKS 1–4

Matter and Measurement

How chemistry became an experimental science (Boyle), SI units, significant figures and dimensional analysis, density and Archimedes' crown, and separating mixtures with Tsvet's chromatography.

1.1 From Alchemy to Experiment: Boyle's The Sceptical Chymist (1661)

1.2 SI Units, Significant Figures and Dimensional Analysis: "A Just Weight Is His Delight"

1.3 Density and Archimedes' Crown

1.4 Mixtures, Separation and Purity: Tsvet's Chromatography (1903–1906)

UNIT 2 · WEEKS 4–7

Atoms

The atom discovered step by step: Lavoisier's conservation of mass, Proust and Dalton, Thomson's electron and Rutherford's nucleus, and isotopes and average atomic mass.

2.1 Lavoisier and the Conservation of Mass (Traité, 1789)

2.2 Proust's Definite Proportions and Dalton's Atomic Theory (1808)

2.3 Thomson's Electron (1897) and Rutherford's Gold Foil (1909–1911)

2.4 Isotopes, Atomic Mass and Aston's Mass Spectrograph (1919)

UNIT 3 · WEEKS 7–11

Electrons and the Periodic Table

Flame tests and spectra (Bunsen, Kirchhoff and Bohr), electron configurations, Mendeleev's table and its predicted elements, and periodic trends.

3.1 Flame, Spectra and Bohr's Atom: Bunsen and Kirchhoff (1859–1860)

3.2 Electron Configurations and Orbitals

3.3 Mendeleev's Table and the Missing Elements (1869; Gallium 1875)

3.4 Periodic Trends: Size, Ionization Energy and Pauling's Electronegativity

UNIT 4 · WEEKS 11–14

Chemical Bonds and Molecular Shape

Ionic and covalent bonds, Lewis dot structures, molecular shape and polarity (VSEPR), and the intermolecular forces that make water so unusual.

4.1 Ionic Bonds and Crystals: Salt of the Earth

4.2 Covalent Bonds and G. N. Lewis's Dot Structures (1916)

4.3 Molecular Shape: VSEPR and Polarity

4.4 Intermolecular Forces: Why Water Is So Strange

UNIT 5 · WEEKS 14–18

Chemical Reactions

Writing and balancing equations, five reaction types and Faraday's candle, precipitates and net ionic equations, and oxidation-reduction from rust to batteries.

5.1 Writing and Balancing Equations

5.2 Five Reaction Types and Faraday's Chemical History of a Candle (1848)

5.3 Precipitates, Solubility Rules and Net Ionic Equations

5.4 Oxidation and Reduction: Rust, Batteries and Davy's Electrolysis (1807)

UNIT 6 · WEEKS 18–21

The Mole and Stoichiometry

Avogadro's number and the mole, molar mass and empirical formulas, stoichiometry, and limiting reactants and percent yield with the Haber–Bosch process.

6.1 Avogadro and the Mole

6.2 Molar Mass, Percent Composition and Empirical Formulas

6.3 Stoichiometry: Chemistry's Recipe Math

6.4 Limiting Reactant and Percent Yield: Haber and Bosch (1909–1913)

UNIT 7 · WEEKS 21–25

Gases

Pressure and Boyle's law, Charles's and Gay-Lussac's laws, the ideal gas law and molar volume, and the kinetic-molecular theory with partial pressures and diffusion.

7.1 Pressure: Torricelli and Boyle's Law (1662)

7.2 Charles and Gay-Lussac: Balloons and Temperature (1783–1804)

7.3 The Ideal Gas Law and Molar Volume

7.4 Kinetic-Molecular Theory, Dalton's Partial Pressures and Graham's Diffusion

UNIT 8 · WEEKS 25–28

Solutions, Acids and Bases

Solutions and molarity, colligative properties (salt on ice and ice cream), acids and bases from Arrhenius to Brønsted–Lowry, and pH, indicators and titration.

8.1 Solutions, Solubility and Molarity

8.2 Colligative Properties: Salt on Ice and Homemade Ice Cream

8.3 Acids and Bases: Arrhenius (1884) to Brønsted and Lowry (1923)

8.4 pH, Indicators and Titration: Sørensen at the Carlsberg Laboratory (1909)

UNIT 9 · WEEKS 28–31

Energy, Rates, Equilibrium and the Nucleus

Thermochemistry and Hess's law, reaction rates and catalysts, equilibrium and Le Chatelier's principle, and nuclear chemistry with half-life.

9.1 Thermochemistry: Lavoisier and Laplace's Ice Calorimeter and Hess's Law (1840)

9.2 Reaction Rates and Catalysts

9.3 Equilibrium and Le Chatelier's Principle (1884)

9.4 Nuclear Chemistry: Becquerel (1896), the Curies (1898) and Half-Life

Weekly pacing calendar

WEEK	PLAN
1	• Day 1: Lesson 1.1 From Alchemy to Experiment: Boyle's The Sceptical Chymist (1661): warm-up, video and core ideas • Day 2: Lesson 1.1: worked examples and first practice problems • Day 3: Lesson 1.1: lab and lab notebook • Day 4: Lesson 1.1: practice set, Worksheet A and corrections • Day 5: Lesson 1.2 SI Units, Significant Figures and Dimensional Analysis: "A Just Weight Is His Delight": warm-up, video and core ideas *Five sessions a week of about 45 minutes. Families on a four-day week can combine the practice day of one lesson with the first day of the next.*
2	• Day 1: Lesson 1.2: worked examples and first practice problems • Day 2: Lesson 1.2: lab and lab notebook • Day 3: Lesson 1.2: practice set, Worksheet A and corrections • Day 4: Lesson 1.3 Density and Archimedes' Crown: warm-up, video and core ideas • Day 5: Lesson 1.3: worked examples and first practice problems
3	• Day 1: Lesson 1.3: lab and lab notebook

WEEK	PLAN
	• Day 2: Lesson 1.3: practice set, Worksheet A and corrections • Day 3: Lesson 1.4 Mixtures, Separation and Purity: Tsvet's Chromatography (1903–1906): warm-up, video and core ideas • Day 4: Lesson 1.4: worked examples and first practice problems • Day 5: Lesson 1.4: lab and lab notebook
4	• Day 1: Lesson 1.4: practice set, Worksheet A and corrections • Day 2: Unit 1 test (Worksheet B sets make a good review the day before) • Day 3: Lesson 2.1 Lavoisier and the Conservation of Mass (Traité, 1789): warm-up, video and core ideas • Day 4: Lesson 2.1: worked examples and first practice problems • Day 5: Lesson 2.1: lab and lab notebook
5	• Day 1: Lesson 2.1: practice set, Worksheet A and corrections • Day 2: Lesson 2.2 Proust's Definite Proportions and Dalton's Atomic Theory (1808): warm-up, video and core ideas • Day 3: Lesson 2.2: worked examples and first practice problems • Day 4: Lesson 2.2: lab and lab notebook • Day 5: Lesson 2.2: practice set, Worksheet A and corrections
6	• Day 1: Lesson 2.3 Thomson's Electron (1897) and Rutherford's Gold Foil (1909–1911): warm-up, video and core ideas • Day 2: Lesson 2.3: worked examples and first practice problems • Day 3: Lesson 2.3: lab and lab notebook • Day 4: Lesson 2.3: practice set, Worksheet A and corrections • Day 5: Lesson 2.4 Isotopes, Atomic Mass and Aston's Mass Spectrograph (1919): warm-up, video and core ideas
7	• Day 1: Lesson 2.4: worked examples and first practice problems • Day 2: Lesson 2.4: lab and lab notebook • Day 3: Lesson 2.4: practice set, Worksheet A and corrections • Day 4: Unit 2 test (Worksheet B sets make a good review the day before) • Day 5: Lesson 3.1 Flame, Spectra and Bohr's Atom: Bunsen and Kirchhoff (1859–1860): warm-up, video and core ideas
8	• Day 1: Lesson 3.1: worked examples and first practice problems • Day 2: Lesson 3.1: lab and lab notebook • Day 3: Lesson 3.1: practice set, Worksheet A and corrections

WEEK	PLAN
	• Day 4: Lesson 3.2 Electron Configurations and Orbitals: warm-up, video and core ideas • Day 5: Lesson 3.2: worked examples and first practice problems
9	• Day 1: Lesson 3.2: lab and lab notebook • Day 2: Lesson 3.2: practice set, Worksheet A and corrections • Day 3: Lesson 3.3 Mendeleev's Table and the Missing Elements (1869; Gallium 1875): warm-up, video and core ideas • Day 4: Lesson 3.3: worked examples and first practice problems • Day 5: Lesson 3.3: lab and lab notebook
10	• Day 1: Lesson 3.3: practice set, Worksheet A and corrections • Day 2: Lesson 3.4 Periodic Trends: Size, Ionization Energy and Pauling's Electronegativity: warm-up, video and core ideas • Day 3: Lesson 3.4: worked examples and first practice problems • Day 4: Lesson 3.4: lab and lab notebook • Day 5: Lesson 3.4: practice set, Worksheet A and corrections
11	• Day 1: Unit 3 test (Worksheet B sets make a good review the day before) • Day 2: Lesson 4.1 Ionic Bonds and Crystals: Salt of the Earth: warm-up, video and core ideas • Day 3: Lesson 4.1: worked examples and first practice problems • Day 4: Lesson 4.1: lab and lab notebook • Day 5: Lesson 4.1: practice set, Worksheet A and corrections
12	• Day 1: Lesson 4.2 Covalent Bonds and G. N. Lewis's Dot Structures (1916): warm-up, video and core ideas • Day 2: Lesson 4.2: worked examples and first practice problems • Day 3: Lesson 4.2: lab and lab notebook • Day 4: Lesson 4.2: practice set, Worksheet A and corrections • Day 5: Lesson 4.3 Molecular Shape: VSEPR and Polarity: warm-up, video and core ideas
13	• Day 1: Lesson 4.3: worked examples and first practice problems • Day 2: Lesson 4.3: lab and lab notebook • Day 3: Lesson 4.3: practice set, Worksheet A and corrections • Day 4: Lesson 4.4 Intermolecular Forces: Why Water Is So Strange: warm-up, video and core ideas • Day 5: Lesson 4.4: worked examples and first practice problems

WEEK	PLAN
14	• Day 1: Lesson 4.4: lab and lab notebook • Day 2: Lesson 4.4: practice set, Worksheet A and corrections • Day 3: Unit 4 test (Worksheet B sets make a good review the day before) • Day 4: Lesson 5.1 Writing and Balancing Equations: warm-up, video and core ideas • Day 5: Lesson 5.1: worked examples and first practice problems
15	• Day 1: Lesson 5.1: lab and lab notebook • Day 2: Lesson 5.1: practice set, Worksheet A and corrections • Day 3: Lesson 5.2 Five Reaction Types and Faraday's Chemical History of a Candle (1848): warm-up, video and core ideas • Day 4: Lesson 5.2: worked examples and first practice problems • Day 5: Lesson 5.2: lab and lab notebook
16	• Day 1: Lesson 5.2: practice set, Worksheet A and corrections • Day 2: Lesson 5.3 Precipitates, Solubility Rules and Net Ionic Equations: warm-up, video and core ideas • Day 3: Lesson 5.3: worked examples and first practice problems • Day 4: Lesson 5.3: lab and lab notebook • Day 5: Lesson 5.3: practice set, Worksheet A and corrections
17	• Day 1: Lesson 5.4 Oxidation and Reduction: Rust, Batteries and Davy's Electrolysis (1807): warm-up, video and core ideas • Day 2: Lesson 5.4: worked examples and first practice problems • Day 3: Lesson 5.4: lab and lab notebook • Day 4: Lesson 5.4: practice set, Worksheet A and corrections • Day 5: Unit 5 test (Worksheet B sets make a good review the day before)
18	• Day 1: Midterm review: formula sheet and missed problems • Day 2: Midterm exam • Day 3: Lesson 6.1 Avogadro and the Mole: warm-up, video and core ideas • Day 4: Lesson 6.1: worked examples and first practice problems • Day 5: Lesson 6.1: lab and lab notebook
19	• Day 1: Lesson 6.1: practice set, Worksheet A and corrections • Day 2: Lesson 6.2 Molar Mass, Percent Composition and Empirical Formulas: warm-up, video and core ideas • Day 3: Lesson 6.2: worked examples and first practice problems

WEEK	PLAN
	• Day 4: Lesson 6.2: lab and lab notebook • Day 5: Lesson 6.2: practice set, Worksheet A and corrections
20	• Day 1: Lesson 6.3 Stoichiometry: Chemistry's Recipe Math: warm-up, video and core ideas • Day 2: Lesson 6.3: worked examples and first practice problems • Day 3: Lesson 6.3: lab and lab notebook • Day 4: Lesson 6.3: practice set, Worksheet A and corrections • Day 5: Lesson 6.4 Limiting Reactant and Percent Yield: Haber and Bosch (1909–1913): warm-up, video and core ideas
21	• Day 1: Lesson 6.4: worked examples and first practice problems • Day 2: Lesson 6.4: lab and lab notebook • Day 3: Lesson 6.4: practice set, Worksheet A and corrections • Day 4: Unit 6 test (Worksheet B sets make a good review the day before) • Day 5: Lesson 7.1 Pressure: Torricelli and Boyle's Law (1662): warm-up, video and core ideas
22	• Day 1: Lesson 7.1: worked examples and first practice problems • Day 2: Lesson 7.1: lab and lab notebook • Day 3: Lesson 7.1: practice set, Worksheet A and corrections • Day 4: Lesson 7.2 Charles and Gay-Lussac: Balloons and Temperature (1783–1804): warm-up, video and core ideas • Day 5: Lesson 7.2: worked examples and first practice problems
23	• Day 1: Lesson 7.2: lab and lab notebook • Day 2: Lesson 7.2: practice set, Worksheet A and corrections • Day 3: Lesson 7.3 The Ideal Gas Law and Molar Volume: warm-up, video and core ideas • Day 4: Lesson 7.3: worked examples and first practice problems • Day 5: Lesson 7.3: lab and lab notebook
24	• Day 1: Lesson 7.3: practice set, Worksheet A and corrections • Day 2: Lesson 7.4 Kinetic-Molecular Theory, Dalton's Partial Pressures and Graham's Diffusion: warm-up, video and core ideas • Day 3: Lesson 7.4: worked examples and first practice problems • Day 4: Lesson 7.4: lab and lab notebook • Day 5: Lesson 7.4: practice set, Worksheet A and corrections

WEEK	PLAN
25	• Day 1: Unit 7 test (Worksheet B sets make a good review the day before) • Day 2: Lesson 8.1 Solutions, Solubility and Molarity: warm-up, video and core ideas • Day 3: Lesson 8.1: worked examples and first practice problems • Day 4: Lesson 8.1: lab and lab notebook • Day 5: Lesson 8.1: practice set, Worksheet A and corrections
26	• Day 1: Lesson 8.2 Colligative Properties: Salt on Ice and Homemade Ice Cream: warm-up, video and core ideas • Day 2: Lesson 8.2: worked examples and first practice problems • Day 3: Lesson 8.2: lab and lab notebook • Day 4: Lesson 8.2: practice set, Worksheet A and corrections • Day 5: Lesson 8.3 Acids and Bases: Arrhenius (1884) to Brønsted and Lowry (1923): warm-up, video and core ideas
27	• Day 1: Lesson 8.3: worked examples and first practice problems • Day 2: Lesson 8.3: lab and lab notebook • Day 3: Lesson 8.3: practice set, Worksheet A and corrections • Day 4: Lesson 8.4 pH, Indicators and Titration: Sørensen at the Carlsberg Laboratory (1909): warm-up, video and core ideas • Day 5: Lesson 8.4: worked examples and first practice problems
28	• Day 1: Lesson 8.4: lab and lab notebook • Day 2: Lesson 8.4: practice set, Worksheet A and corrections • Day 3: Unit 8 test (Worksheet B sets make a good review the day before) • Day 4: Lesson 9.1 Thermochemistry: Lavoisier and Laplace's Ice Calorimeter and Hess's Law (1840): warm-up, video and core ideas • Day 5: Lesson 9.1: worked examples and first practice problems
29	• Day 1: Lesson 9.1: lab and lab notebook • Day 2: Lesson 9.1: practice set, Worksheet A and corrections • Day 3: Lesson 9.2 Reaction Rates and Catalysts: warm-up, video and core ideas • Day 4: Lesson 9.2: worked examples and first practice problems • Day 5: Lesson 9.2: lab and lab notebook
30	• Day 1: Lesson 9.2: practice set, Worksheet A and corrections • Day 2: Lesson 9.3 Equilibrium and Le Chatelier's Principle (1884): warm-up, video and core ideas

WEEK	PLAN
	• Day 3: Lesson 9.3: worked examples and first practice problems • Day 4: Lesson 9.3: lab and lab notebook • Day 5: Lesson 9.3: practice set, Worksheet A and corrections
31	• Day 1: Lesson 9.4 Nuclear Chemistry: Becquerel (1896), the Curies (1898) and Half-Life: warm-up, video and core ideas • Day 2: Lesson 9.4: worked examples and first practice problems • Day 3: Lesson 9.4: lab and lab notebook • Day 4: Lesson 9.4: practice set, Worksheet A and corrections • Day 5: Unit 9 test (Worksheet B sets make a good review the day before)
32	• Day 1: Final review • Day 2: Final exam

Materials

Scientific calculator and a printed periodic table

With logarithms and scientific notation.

Bound lab notebook, pencil, ruler

For every lab, graph and corrected practice problem.

Safety: ANSI Z87.1 splash goggles and nitrile gloves

About $20. Worn in every wet lab.

Digital kitchen scale reading to 0.01 g (or 0.1 g)

About $15.

Digital probe thermometer

About $12, for calorimetry, gas and solution labs.

Measuring tools

60 mL syringes (4), plastic graduated cylinders (10, 50 and 100 mL), droppers or pipettes; about $25.

pH strips, a 9 V battery with an LED and clip leads

About $15, for acid-base and conductivity labs.

Household chemicals

Baking soda, vinegar, table salt, salt substitute (KCl), Epsom salt, washing soda, borax, calcium chloride ice melt, cream of tartar, 3% hydrogen peroxide, dry

yeast, steel wool, 91% isopropyl alcohol, antacid tablets, iodine tincture, vitamin C tablets, cornstarch and red cabbage; about $40–50.

From a science supplier

500 mL of 0.50 M sodium hydroxide solution and phenolphthalein indicator for the Unit 8 titration; about $15. Copper sulfate (root killer) for Units 3 and 5.

Optional

A diffraction-grating viewer or an old CD for spectra; a small beaker and flask set from a science supplier. Core lab total is about $125–165.

Grading

Grades come from the practice sets, the labs, the lab notebook, the nine unit tests, the midterm and the final. A calculation earns full credit when the equation, the substitution with units and the answer with its unit and sensible significant figures are all shown. A lab earns full credit for honest data carefully analyzed, even when the result disagrees with the expected value, as long as the student explains why.

PART	%
Practice sets and worksheets (completed and corrected)	**15**
Labs (procedure, data, analysis)	**25**
Lab notebook (organization, honesty, error discussion)	**10**
Unit tests	**30**
Midterm exam	**10**
Final exam	**10**

Scale: A 90–100, B 80–89, C 70–79, D 60–69, F below 60.

- Practice is graded for completion and correction: a problem done, checked and correctly reworked earns full credit.

- On tests, give partial credit for a correct equation and setup with units even if the arithmetic slips. A missing or wrong unit costs a point.
- For a high-school transcript this course is one full credit of laboratory chemistry.

A note to parents

You do not need to know chemistry to supervise this course well. Your jobs are to keep the rhythm, check the practice against the answers, ask the student to explain one worked example aloud each week, read the lab notebook, and be present for the labs that use a flame, hot liquids, dry ice or anything that could splash.

Expect the labs not to match the textbook exactly. Kitchen scales, impure household chemicals and heat lost to the air are always present, and finding them is half of what a chemist does. Praise careful measurement and honest records more than a perfect answer.

Many of the founders of chemistry were believing Christians who saw their work as studying an orderly creation: Boyle, Faraday, Dalton and others. The course says so plainly where it is true. Nuclear decay in Unit 9 is taught as measured arithmetic; radiometric dating is described as a method resting on stated assumptions, with a pointer to Earth Science 7 and High School Geology for the fuller discussion. Combustion and carbon dioxide are taught as chemistry, not policy.

UNIT 1

Matter and Measurement

Chemistry is the study of matter and how it changes. This first unit lays the foundation that every later unit stands on: how chemists moved from alchemy to careful experiment, how to measure honestly and report a measurement with the right number of significant figures, how to convert units with dimensional analysis, how density identifies a substance (Archimedes and the crown), and how mixtures are separated by filtering, distilling and chromatography.

The labs use a kitchen scale, a measuring cup and coffee filters. The habits you build here (a written procedure, a data table with units, an honest record of what

happened) are the habits of Robert Boyle, who insisted that chemistry must be built on experiments anyone can repeat.

PREREQUISITES

- Algebra I: solving one-step and two-step equations
- Scientific notation
- Ratios and proportions

MASTERY CHECKLIST

- ☐ I can explain how Boyle and the early chemists turned alchemy into an experimental science.
- ☐ I can count significant figures and round a calculated answer correctly.
- ☐ I can convert units with dimensional analysis, including squared and cubed units.
- ☐ I can measure and calculate density and use it to identify a substance.
- ☐ I can classify matter and choose a method to separate a mixture.

LESSONS IN THIS UNIT

1.1 From Alchemy to Experiment: Boyle's The Sceptical Chymist (1661)

4 sessions of 45 minutes

1.2 SI Units, Significant Figures and Dimensional Analysis: "A Just Weight Is His Delight"

4 sessions of 45 minutes

1.3 Density and Archimedes' Crown 4 sessions of 45 minutes

1.4 Mixtures, Separation and Purity: Tsvet's Chromatography (1903–1906)

4 sessions of 45 minutes

Unit 1 test

Show your work. For every calculation write the equation, substitute with units and give the answer with its unit and sensible significant figures. The test is worth 100 points.

1. **8 pts** **Lesson 1.1**

   Boyle doubted that everything is made of earth, water, air and fire. What kind of evidence did he say was needed to decide what the elements are?

2. **8 pts** **Lesson 1.1**

   Give one reason a scientist's faith encouraged careful science in Boyle's time.

3. **8 pts** **Lesson 1.2**

   Round 0.006 749 to two significant figures.

4. **8 pts** **Lesson 1.2**

   Convert 1.00 cubic foot to liters (1 ft = 30.48 cm).

5. **8 pts** **Lesson 1.2**

   Convert 55 miles per hour to meters per second (1 mi = 1609 m).

6. **7 pts** **Lesson 1.3**

   What instrument measures the density of a liquid directly?

7. **7 pts** **Lesson 1.3**

   What happens to the density of most substances as they are heated?

8. **7 pts** **Lesson 1.3**

   What is the mass of 24.9 cm^3 of zinc (density 7.13 g/cm^3)?

9. **7 pts** **Lesson 1.4**

   Explain why burning paper is a chemical change.

10. **7 pts** **Lesson 1.4**

    Brass is copper and zinc melted together. Is it an element, compound or mixture? Defend your answer.

LAB AND DATA

11. **10 pts**

A student measures a metal cylinder five times: 12.31 g, 12.29 g, 12.33 g, 12.30 g, 12.32 g. The accepted mass is 12.45 g. (a) Find the mean. (b) Are the measurements precise, accurate, both or neither? (c) Name one likely cause.

SHORT ESSAY

12. **15 pts**

In a paragraph, explain what Robert Boyle changed about the way chemistry was done, and why a written, repeatable procedure matters.

ANSWER KEY

1. Experimental evidence: substances that cannot be broken down further by experiment should be counted as elements, not ideas accepted from tradition.
2. Belief that God made an orderly, lawful world meant nature could be studied and its patterns discovered.
3. 0.0067
4. 28.3 L ($30.48^3 = 28\,317\ \text{cm}^3$).
5. $25\ \text{m/s}$
   $55\ \frac{\text{mi}}{\text{h}} \times \frac{1609\ \text{m}}{1\ \text{mi}} \times \frac{1\ \text{h}}{3600\ \text{s}} \cdot = 24.6\ \text{m/s}$, or 25 to two significant figures.
6. A hydrometer.
7. It decreases, because they expand while the mass stays the same.
8. 178 g
   $m = \rho V = (7.13)(24.9) = 178\ \text{g}$
9. New substances (carbon dioxide, water vapor, ash) form, with heat and light released.
10. A homogeneous mixture (an alloy); the copper-to-zinc ratio can vary and the parts are not chemically bonded in a fixed ratio.
11. (a) 12.31 g. (b) Precise (they agree within 0.02 g) but not accurate (about 0.14 g low). (c) A scale that is out of calibration (a systematic error); taring or calibrating with known masses would fix it.
12. Look for: Boyle (The Sceptical Chymist, 1661) rejected secret recipes and the four-element theory as untested; he insisted on public, repeatable experiments, careful records and reasoning from evidence; repeatability lets others check and build on results; a full answer mentions his air pump or gas law and his faith that nature is orderly because God made it.

Printable answer key (PDF, for the parent)

1.1 From Alchemy to Experiment: Boyle's The Sceptical Chymist (1661)

OBJECTIVES

- Describe what alchemists believed and what they actually contributed (apparatus, techniques).
- Explain why Boyle's The Sceptical Chymist marks the turn to experimental chemistry.
- Distinguish observation, hypothesis, experiment, law and theory.
- Set up a lab notebook the way the course requires.

> **Warm-up.** Light a candle (with an adult) and write down ten observations about it, using only your senses and a ruler. Then sort your list into observations (what you saw) and explanations (why you think it happened).

WATCH

Introduction to chemistry

Khan Academy · 7:45

https://www.youtube.com/watch?v=Rd4a1X3B61w

> **WHEN TO WATCH**
> After the core ideas, as an overview of what chemistry studies.

BACKUP VIDEO

The scientific method

Khan Academy · 11:48

https://www.youtube.com/watch?v=N6IAzlugWw0

Optional, after the lab, to review the steps of the scientific method.

KEY TERMS

matter

Anything that has mass and takes up space.

chemistry

The study of the composition, structure and properties of matter and the changes it undergoes.

hypothesis

A testable, specific explanation or prediction that an experiment can support or refute.

scientific law

A statement, often an equation, that describes a pattern observed over and over in nature, such as Boyle's law.

theory

A well-tested explanation of why a set of observations and laws hold, such as the atomic theory.

control

A part of an experiment kept the same, or a trial without the change, so that any effect can be traced to the variable being tested.

CORE IDEAS

The story. For over a thousand years the people who worked with furnaces, flasks and ores were alchemists. They hoped to turn lead into gold and find an elixir of life, and they kept their recipes secret, written in symbols and riddles. Yet they invented much of the apparatus chemists still use (the still, the water bath, the crucible) and learned to make acids, alcohol and pure metals. In 1661 the Irish-born scientist Robert Boyle published *The Sceptical Chymist*, a dialogue in which his spokesman doubts every claim that cannot be shown by experiment, including the old idea that everything is made of four elements (earth, water, air and fire).

How they figured it out. Boyle did not simply argue; he experimented, and he published exactly what he did so that others could repeat it. With his assistant Robert Hooke he built an air pump and showed that a candle goes out and a bell falls silent in a vacuum. His work on air led to Boyle's law (Lesson 7.1). He also helped found the Royal Society, whose motto, *Nullius in verba*, means "take nobody's word for it." Boyle was a devout Christian who saw nature as an orderly creation that could be studied, and he left money in his will for lectures defending the faith.

The chemistry. Science moves from *observation* to a testable *hypothesis*, then to an *experiment* with one variable changed and the others controlled. A pattern confirmed many times becomes a *law*, which says *what* happens (often as an equation). A *theory* explains *why* the laws hold, and it stays open to correction by new evidence. Matter is anything with mass and volume; chemistry studies what matter is made of and how it changes.

Where it is used. Every medicine, fertilizer, battery, paint and plastic was developed by this cycle of hypothesis and test. Good records matter: in industry and medicine a result that cannot be repeated cannot be trusted, and a lab notebook is a legal record. In this course your notebook is where you practice being honest with data.

> **Safety first.** An adult must be present whenever there is an open flame. Tie back long hair, keep paper away, and have water nearby.

LAB

Observing a Candle like Faraday and Boyle

You will make careful observations of a burning candle, form a hypothesis about why a candle goes out under a jar, and design a controlled test.

You will need

- A short candle on a saucer, matches or a lighter
- Three glass jars of different sizes
- A stopwatch, ruler, notebook
- Safety glasses

Steps

1. Set up your notebook: title, date, question, hypothesis, procedure, data table, results and discussion.
2. Light the candle. Record at least 15 observations (size and color of the flame, the wax pool, the wick, smoke when snuffed).
3. Write a hypothesis: how will the jar size affect how long the candle burns under it? Give a reason.
4. Cover the candle with each jar in turn, timing until it goes out. Let the jar air out between trials. Do three trials per jar.
5. Make a data table and find the average time for each jar. Graph time against jar volume.

Check your work

- ☐ Did you change only one variable (jar size)?
- ☐ Did you repeat each trial at least three times?
- ☐ Does your conclusion say whether the data support your hypothesis?

LAB NOTEBOOK

1. Your 15 candle observations, sorted into observations and inferences.
2. Data table (jar, volume, three times, average) and a graph of burn time against volume.

> **From history.** Boyle was the fourteenth child of the Earl of Cork. Besides chemistry he learned Hebrew and Greek to study the Bible, paid for translations of the Scriptures into Irish and other languages, and wrote *The Christian Virtuoso* (1690) to argue that studying nature and believing in God belong together.

PRACTICE

1. Easy What was the main difference between how alchemists and Boyle shared their work?

   ▼ **Show answer**

   Alchemists kept recipes secret and wrote in riddles; Boyle published exactly what he did so others could repeat and check it.

2. Easy What does the Royal Society motto *Nullius in verba* mean, and why is it a good motto for science?

   ▼ **Show answer**

   "Take nobody's word for it." Claims should be tested by experiment, not accepted because of who said them.

3. Medium Explain the difference between a scientific law and a theory, with an example of each.

   ▼ **Show answer**

   A law describes what happens, often as an equation (Boyle's law: PV is constant). A theory explains why (the kinetic-molecular theory explains Boyle's law).

COMMON MISTAKES

- Calling a guess a theory. In science a theory is a well-tested explanation, not a hunch.
- Changing two things at once (jar size and candle size), so the result cannot be traced to either.
- Recording what should have happened instead of what did happen.

REFERENCE GUIDE

An eight-page, color-coded General Chemistry guide: measurement and significant figures, atoms and electrons, periodic trends, bonding and molecular shape, intermolecular forces, naming, the mole and stoichiometry, reactions, solutions, gas laws, thermochemistry, equilibrium, acids and bases, and redox and nuclear change. The booklet prints on 11×17 paper and folds to letter size. Prints in color and reads clearly in black and white.

> **Make it a game · Observation or Inference?.** A parent reads statements about an everyday object ("the bread is brown," "the bread is stale because it was left out"). The student says "observation" or "inference" within three seconds. One point each; switch roles after ten.

> **Scripture.** "It is the glory of God to conceal a thing: but the honour of kings is to search out a matter." (Proverbs 25:2, KJV)
>
> Boyle and many early scientists saw research as searching out what God has made. Patient experiment is one way to honor the Creator.

UNIT 1 · MATTER AND MEASUREMENT · LESSON 1.2 OF 36 · 4 SESSIONS OF 45 MINUTES

1.2 SI Units, Significant Figures and Dimensional Analysis: "A Just Weight Is His Delight"

OBJECTIVES

- Use SI base units and metric prefixes from nano- to kilo-.
- Count significant figures and apply the rules for calculations.
- Distinguish accuracy from precision.
- Convert units, including squared and cubed units, by dimensional analysis.

> **Warm-up.** Measure the length of a pencil with a ruler marked only in centimeters, then with one marked in millimeters. Write each measurement with the last digit estimated. How many digits can you honestly write each time?

WATCH

Why are Significant Figures Important?

Tyler DeWitt · 7:45

https://www.youtube.com/watch?v=VAuslY-Uuf4

> **WHEN TO WATCH**
> After the core ideas on significant figures, before the worked examples.

BACKUP VIDEO

Unit Conversion with Multiple Conversion Factors

Melissa Maribel · 2:54

https://www.youtube.com/watch?v=t6fYYPrzhOU

Before Example 2, for a short demonstration of chained conversion factors.

KEY TERMS

SI

The International System of Units: meter, kilogram, second, kelvin, mole, ampere and candela, with decimal prefixes.

significant figures

The digits in a measurement that are known for certain plus one estimated digit.

accuracy

How close a measurement is to the true or accepted value.

precision

How close repeated measurements are to each other.

conversion factor

A ratio equal to 1, such as 1000 mL/1 L, used to change units.

dimensional analysis

Solving problems by multiplying by conversion factors so that unwanted units cancel.

CORE IDEAS

The story. Before the metric system, a “foot” or a “pound” varied from town to town, and merchants could cheat with false weights. Scripture condemns this: “A false

balance is abomination to the LORD: but a just weight is his delight” (Proverbs 11:1). In 1795 the French defined the meter and the gram on a decimal scale, and in 1875 seventeen nations signed the Treaty of the Meter. Since 2019 every SI unit, even the kilogram, is defined by fixed constants of nature rather than a metal object.

How they figured it out. Every measurement has limits. If a ruler is marked in millimeters, you can read 12.3 mm but estimate the last digit; 12.37 mm would claim more than you know. Chemists record all certain digits plus one estimated digit. These are the *significant figures*. Rules: nonzero digits count; zeros between nonzero digits count; leading zeros (0.0045) never count; trailing zeros count only if there is a decimal point (2.50 has three, 2500 has two unless written as 2.500×10^{3}).

The chemistry. When you multiply or divide, keep as many significant figures as the measurement with the *fewest*. When you add or subtract, keep as many *decimal places* as the measurement with the fewest. Exact numbers (12 eggs, 1000 mL in 1 L) never limit an answer. For conversions, multiply by factors equal to one, arranged so the old unit cancels: $5.0\ \text{km} \times \dfrac{1000\ \text{m}}{1\ \text{km}} = 5.0 \times 10^{3}\ \text{m}$. For squared or cubed units, square or cube the factor too: $1\ \text{m}^3 = (100\ \text{cm})^3 = 10^{6}\ \text{cm}^3$.

Where it is used. In 1983 an Air Canada jet ran out of fuel in flight because fuel was loaded in pounds when kilograms were needed; in 1999 NASA lost the Mars Climate Orbiter over a pound-second and newton-second mix-up. Nurses convert milligrams per kilogram of body weight every day. Units carried through every step are the best error check you have.

LAB

Calibrating a Kitchen Scale with Nickels

A U.S. nickel is minted at 5.000 g. You will use nickels as standard masses to test your scale's accuracy and precision, and measure a liquid's volume three ways.

You will need

- Kitchen scale reading to 0.01 g or 0.1 g
- 10 nickels
- Measuring cup, tablespoon, a 10 mL medicine syringe or graduated cylinder
- Water

Steps

1. Zero (tare) the scale. Weigh 1, 2, 4, 6, 8 and 10 nickels, recording every digit the scale shows.
2. Graph measured mass against expected mass (5.000 g per nickel). Draw the best-fit line and find its slope.
3. Weigh one nickel five times, removing it each time. Find the mean and the spread.
4. Measure 50 mL of water with the measuring cup, then weigh it (water is 1.00 g/mL at room temperature). Repeat with the syringe or cylinder.
5. Calculate the percent error for each method.

Check your work

- ☐ Does every number in your table have a unit?
- ☐ Did you record the scale reading exactly as shown, including zeros?
- ☐ Which volume tool was most accurate? Most precise?

LAB NOTEBOOK

1. The nickel data table and graph with best-fit slope.
2. A sentence on whether your scale is accurate, precise, both or neither.

KEY FORMULAS

$$\text{percent error} = \frac{|\text{measured} - \text{accepted}|}{\text{accepted}} \times 100\%$$

$$1\ \text{L} = 1000\ \text{mL} = 1000\ \text{cm}^3 \qquad 1\ \text{in} = 2.54\ \text{cm (exact)}$$

From history. The original kilogram, a platinum-iridium cylinder kept in a vault near Paris since 1889, was found to drift by about 50 micrograms compared with its copies. In 2019 the kilogram was redefined using Planck's constant, measured with an instrument called a Kibble balance.

WORKED EXAMPLES

Example 1. Calculate and round correctly: $(4.52\ \text{cm})(2.1\ \text{cm})$, then $12.47\ \text{g} + 3.2\ \text{g}$.

▼ **Show solution**

1. Multiplication: $4.52 \times 2.1 = 9.492$. The fewest significant figures is 2 (in 2.1), so round to $9.5\ \text{cm}^2$.
2. Addition: $12.47 + 3.2 = 15.67$. The fewest decimal places is one (3.2), so round to $15.7\ \text{g}$.

Answer: $9.5\ \text{cm}^2$ and $15.7\ \text{g}$

Example 2. A fish tank holds 25.0 gallons. How many liters and cubic centimeters is that? (1 gal = 3.785 L)

▼ **Show solution**

1. $25.0\ \text{gal} \times \dfrac{3.785\ \text{L}}{1\ \text{gal}} = 94.6\ \text{L}$
2. $94.6\ \text{L} \times \dfrac{1000\ \text{cm}^3}{1\ \text{L}} = 9.46 \times 10^4\ \text{cm}^3$

Answer: $94.6\ \text{L} = 9.46 \times 10^4\ \text{cm}^3$

PRACTICE

1. Easy Which is more precise: three measurements of 10.1, 10.2, 10.1 cm, or 9.5, 10.6, 10.0 cm?

   ▼ **Show answer**

   The first set; the values agree closely.

2. Medium The accepted density of water is 1.00 g/mL. A student measures 0.96 g/mL. Find the percent error.

   ▼ **Show answer**

   4%

3. Medium Multiply and give the answer with the correct significant figures: 61.04 cm × 6.9 cm.

   ▼ **Show answer**

   $4.2 \times 10^{2}\ \text{cm}^2$

4. Medium A road race is 3 miles long. Convert to kilometers (1 mi = 1.609 km) and then to meters.

   ▼ **Show answer**

   $4.83\ \text{km} = 4.83 \times 10^{3}\ \text{m}$

5. Hard Convert 75 miles per hour to meters per second (1 mi = 1609 m).

   ▼ **Show answer**

   $34\ \text{m/s}$

COMMON MISTAKES

- Counting leading zeros as significant (0.0050 has two, not four).
- Rounding in the middle of a calculation; keep extra digits and round only at the end.
- Flipping a conversion factor so the units multiply instead of cancel.

> **Make it a game · Prefix Race.** Write the prefixes nano-, micro-, milli-, centi-, kilo- and mega- on cards. A parent calls a conversion (“3.2 kilometers to meters”), and the student slaps the right prefix card and answers. Fastest correct answer wins the round.

> **Scripture.** “A false balance is abomination to the LORD: but a just weight is his delight.” (Proverbs 11:1, KJV)
>
> Honest measurement is a moral matter. Recording the digit you actually see, not the one you wish you saw, is a just weight.

UNIT 1 · MATTER AND MEASUREMENT · LESSON 1.3 OF 36 · 4 SESSIONS OF 45 MINUTES

1.3 Density and Archimedes’ Crown

OBJECTIVES

- Define density and calculate it from mass and volume.
- Measure the volume of an irregular solid by water displacement.
- Use density to identify substances and predict floating and sinking.
- Explain how Archimedes could test the king’s crown.

> **Warm-up.** Drop a can of regular soda and a can of diet soda into a sink full of water. Predict first: will they both sink? Explain what you see.

WATCH

Density: A Story of Archimedes and the Gold Crown

Tyler DeWitt · 9:10

https://www.youtube.com/watch?v=KMNwXUCXLdk

> **WHEN TO WATCH**
> Start of session 1, as the story of the lesson.

BACKUP VIDEO

Density Practice Problems

Tyler DeWitt · 8:56

https://www.youtube.com/watch?v=7tVebi3TSsg

After the worked examples, for more practice problems.

KEY TERMS

density

Mass per unit volume, $\rho = m/V$, usually in g/cm³ or g/mL.

volume

The amount of space an object occupies, in cm³, mL or L.

displacement

Finding an object's volume by the rise in water level when it is submerged.

buoyancy

The upward force a fluid exerts on an object in it, equal to the weight of fluid displaced.

intensive property

A property that does not depend on the amount of sample, such as density or boiling point.

extensive property

A property that depends on the amount of sample, such as mass or volume.

CORE IDEAS

The story. According to the Roman writer Vitruvius, King Hiero II of Syracuse gave a goldsmith a weighed amount of gold to make a crown. The crown weighed the right amount, but Hiero suspected that some gold had been replaced with silver. He asked Archimedes (about 250 B.C.) to find out without damaging the crown. Getting into a full bath, Archimedes saw the water overflow and realized he could measure the crown's volume by the water it displaced. He ran home shouting "Eureka!" ("I have found it!"). Historians think the bath detail may be legend, but the physics is sound.

How they figured it out. Silver is less dense than gold, so a crown with silver mixed in would take up more volume than pure gold of the same mass and would displace more water. Archimedes may actually have used a balance under water: hang the crown and an equal mass of pure gold on a balance and lower both into water; if the crown has more volume, the water pushes it up more and the balance tips.

The chemistry. Density is mass divided by volume: $\rho = m/V$. Water is about 1.00 g/mL; gold 19.3 g/cm³; aluminum 2.70 g/cm³. Density is an *intensive* property: a gold ring and a gold bar have the same density, so density helps identify substances. Objects less dense than a fluid float in it; denser objects sink. For an irregular solid, measure volume by displacement: the rise in water level in a graduated cylinder

equals the object's volume (1 mL = 1 cm^3). Density usually decreases as temperature rises because matter expands, but water is unusual: ice is less dense than liquid water, so lakes freeze from the top down.

Where it is used. Hydrometers measure the density of battery acid, antifreeze and fermenting juice. Geologists identify minerals by density, and recyclers separate plastics by floating them in liquids of chosen densities. Ships float because their hollow hulls give them an average density less than water.

> **Safety first.** Rubbing alcohol is flammable: no flames nearby. Do not taste anything in the lab. Pour the column out into the trash, not the drain, because of the honey and oil.

LAB

A Density Column and the Density of Pennies

You will build a density column from kitchen liquids, then measure the density of pre-1982 and post-1982 pennies to detect the change from copper to zinc.

You will need

- Honey, dish soap, water with food coloring, vegetable oil, rubbing alcohol (isopropyl)
- A tall clear glass, a funnel or spoon
- Small objects: grape, cherry tomato, cork, plastic button, paper clip
- 20 pennies dated before 1982 and 20 dated after 1982
- Scale, 100 mL graduated cylinder or narrow measuring cup

Steps

1. Measure 20 mL of each liquid and weigh it; calculate each density.
2. Pour the liquids slowly over the back of a spoon in order from densest to least dense. Record the order.
3. Predict where each small object will settle, then drop it in.
4. Weigh 20 old pennies, then find their volume by displacement. Repeat with 20 new pennies.
5. Calculate both densities and compare with copper (8.96 g/cm^3) and zinc (7.13 g/cm^3).

Check your work

- ☐ Did your measured densities predict the layer order?
- ☐ Is the newer penny density lower? By how much?
- ☐ Why use 20 pennies instead of one?

LAB NOTEBOOK

1. Table of liquids with mass, volume and density, and a labeled sketch of your column.
2. Penny densities with percent difference from copper.

KEY FORMULAS

$$\rho = \frac{m}{V} \qquad m = \rho V \qquad V = \frac{m}{\rho}$$

From history. Vitruvius told the crown story in *De Architectura*, about 200 years after Archimedes. Galileo, in 1586, argued that the overflow method would be too imprecise for a crown and proposed the underwater balance instead, writing a short work called *La Bilancetta* ("The Little Balance").

WORKED EXAMPLES

Example 1. A ring has a mass of 5.80 g. Dropped into a graduated cylinder, it raises the water from 20.0 mL to 20.3 mL. Could it be pure gold (19.3 g/cm^3)?

▼ **Show solution**

1. $V = 20.3 - 20.0 = 0.3\ \text{mL} = 0.3\ \text{cm}^3$
2. $\rho = 5.80 / 0.3 = 19\ \text{g/cm}^3$ (one significant figure in the volume limits the answer to about 20).
3. The result is consistent with gold, but the volume reading is too coarse to be sure; a narrower cylinder or an underwater balance would give a better test.

Answer: About 20 g/cm^3: consistent with gold, but not precise enough to prove it.

Example 2. What is the mass of 2.50 L of olive oil (density 0.918 g/mL)?

▼ **Show solution**

1. $2.50\ \text{L} = 2500\ \text{mL}$
2. $m = \rho V = (0.918\ \text{g/mL})(2500\ \text{mL}) = 2295\ \text{g}$
3. Three significant figures: 2.30×10^3 g (2.30 kg).

Answer: 2.30 kg

PRACTICE

1. Easy Why does ice float on water?

   ▼ **Show answer**

   Ice is less dense than liquid water (about 0.92 g/cm^3).

2. Easy A metal block has mass 103.4 g and volume 14.5 cm^3. Find its density and name the likely metal (aluminum 2.7, iron 7.87, copper 8.96, lead 11.3, gold 19.3, silver 10.5, zinc 7.13 g/cm^3).

   ▼ **Show answer**

   7.13 g/cm^3, zinc

3. Medium Is density an intensive or extensive property? Explain.

   ▼ **Show answer**

   Intensive; it does not depend on the amount of the sample.

4. Medium What is the mass of 28.6 cm^3 of iron (density 7.87 g/cm^3)?

   ▼ **Show answer**

   225 g

5. Hard A 53.3 g sample of an unknown metal is dropped into a graduated cylinder. The water rises from 30.9 mL to 37.7 mL. Find the density.

   ▼ **Show answer**

   7.84 g/cm^3

COMMON MISTAKES

- Forgetting that 1 mL equals 1 cm^3, so displacement volumes are already in cm^3.
- Dividing volume by mass instead of mass by volume.
- Thinking heavy objects always sink; it is density compared with the fluid that decides.

> **Make it a game · Sink or Float?.** Gather ten household objects. The student predicts "sink" or "float" for each in water, then in salt water. Score one point per correct prediction and explain each miss using density.

> **Scripture.** “Who hath measured the waters in the hollow of his hand, and meted out heaven with the span, and comprehended the dust of the earth in a measure, and weighed the mountains in scales, and the hills in a balance?” (Isaiah 40:12, KJV)
>
> The Creator measures waters and weighs mountains. Measuring mass and volume is a small way of thinking God’s thoughts after Him.

UNIT 1 · MATTER AND MEASUREMENT · LESSON 1.4 OF 36 · 4 SESSIONS OF 45 MINUTES

1.4 Mixtures, Separation and Purity: Tsvet’s Chromatography (1903–1906)

OBJECTIVES

- Classify matter as element, compound, homogeneous mixture or heterogeneous mixture.
- Distinguish physical from chemical properties and changes.
- Choose and explain a separation method: filtration, evaporation, distillation, chromatography, magnetism.
- Calculate and interpret R_f values in paper chromatography.

> **Warm-up.** Stir a spoonful of sand, salt and iron filings (or a few crushed paper clips) together. Before you read on, write a plan to get all three back separately.

WATCH

Candy Chromatography: What Colors Are in Your Candy? | Science Project

Science Buddies · 4:48

https://www.youtube.com/watch?v=GDROkh1zlr8

> **WHEN TO WATCH**
> Before the lab, to see candy chromatography done.

BACKUP VIDEO

Separating Mixtures with Distillation

Tyler DeWitt · 9:06

https://www.youtube.com/watch?v=XHwlke8MMcU

After the core ideas on distillation.

KEY TERMS

element

A pure substance made of only one kind of atom that cannot be broken down by chemical means.

compound

A pure substance of two or more elements chemically combined in a fixed ratio.

homogeneous mixture

A mixture with uniform composition throughout, also called a solution, such as salt water or air.

heterogeneous mixture

A mixture whose parts can be seen or are not uniform, such as sand in water.

distillation

Separating liquids by boiling and re-condensing the vapor, based on different boiling points.

chromatography

Separating the components of a mixture as a solvent carries them at different speeds through a stationary material.

CORE IDEAS

The story. In 1903 the Russian-Italian botanist Mikhail Tsvet poured a green extract of leaves into a glass column packed with powdered chalk and washed it through with solvent. The green separated into bands of yellow, orange and different greens as it moved down the column. He had separated the pigments of plants, including the chlorophylls and carotenes. In 1906 he named the method *chromatography*, "color writing." For decades few noticed; today chromatography is in every chemistry, forensics and drug-testing lab.

How they figured it out. Tsvet realized that each pigment clings to the chalk with a different strength. Pigments that cling weakly travel fast with the solvent; those that cling strongly lag behind. In paper chromatography the paper is the stationary phase and water or alcohol is the mobile phase. The ratio of the distance a spot travels to the distance the solvent travels is the R_f value, characteristic of that substance in that solvent.

The chemistry. *Pure substances* are elements (one kind of atom: gold, oxygen) or compounds (elements combined in a fixed ratio: water, salt). *Mixtures* combine substances physically in any ratio; they are homogeneous (solutions, air, brass) or heterogeneous (granite, salad, muddy water). *Physical* properties (color, density, melting point) can be measured without changing the substance; *chemical* properties (flammability, reactivity) show up only when it changes into something new. Mixtures are separated by physical differences: particle size (filtration), boiling point (distillation), solubility (dissolving and evaporating), attraction to paper (chromatography) and magnetism.

Where it is used. Fractional distillation separates crude oil into gasoline, diesel and asphalt; evaporation ponds harvest sea salt; gas chromatography detects drugs in athletes and pesticides in food. Purity matters in medicine: a pharmaceutical must be tested to show it contains only what the label says.

> **Safety first.** Do not eat candies that have been handled in the lab. Wash hands afterward.

LAB

Marker and Candy Chromatography

You will separate the dyes in washable markers and colored candy coatings and calculate R_f values.

You will need

- White coffee filters cut into strips 2 cm × 10 cm
- Washable markers (black, brown, green), colored candies (such as M&Ms or Skittles)
- Toothpicks, a glass, water, a pencil and ruler
- Optional: 1% salt water as a second solvent

Steps

1. Draw a pencil line 2 cm from the bottom of each strip (pencil does not run).
2. Put a small, dark marker dot on the line. For candy, wet the candy, rub it on a toothpick and dab the color on the line several times.
3. Hang each strip so the bottom dips in 1 cm of water but the dot stays above the water.
4. Remove the strip when the solvent is about 1 cm from the top; mark the solvent front with pencil. Let it dry.
5. Measure each spot's distance and calculate R_f. Compare the same color from different brands.

Check your work

☐ Did any single color turn out to be a mixture of dyes?

☐ Do the same dyes give the same

R_f on different strips?

☐ Why must the starting dot stay above the water?

LAB NOTEBOOK

1. Tape your dried chromatograms into the notebook and label every spot with its R_f.

2. A classification chart: ten kitchen items sorted into element, compound, homogeneous mixture, heterogeneous mixture.

KEY FORMULAS

$$R_f = \frac{\text{distance traveled by spot}}{\text{distance traveled by solvent}}$$

From history. Tsvet published in Russian and German during the upheavals before the Russian Revolution, and his method was almost forgotten until the 1930s. Archer Martin and Richard Synge won the 1952 Nobel Prize for partition chromatography, building directly on his idea.

WORKED EXAMPLES

Example 1. In a chromatogram the solvent front moved 8.0 cm. A blue spot moved 5.6 cm and a yellow spot moved 2.0 cm. Find each R_f.

▼ **Show solution**

1. Blue: $R_f = 5.6/8.0 = 0.70$
2. Yellow: $R_f = 2.0/8.0 = 0.25$
3. The blue dye is more attracted to the solvent (or less to the paper) than the yellow.

Answer: Blue 0.70; yellow 0.25

Example 2. Plan the separation of sand, salt and iron filings.

▼ **Show solution**

1. Pull out the iron filings with a magnet (magnetism).
2. Add water and stir; the salt dissolves, the sand does not.
3. Filter: sand stays on the filter paper; salt water passes through.
4. Evaporate the water from the filtrate to recover the salt.

Answer: Magnet, dissolve, filter, evaporate.

PRACTICE

1. Easy Classify: (a) table salt, (b) air, (c) gold, (d) granite.

   ▼ **Show answer**

   (a) compound, (b) homogeneous mixture, (c) element, (d) heterogeneous mixture.

2. Easy Is melting butter a physical or chemical change?

   ▼ **Show answer**

   Physical; it is still butter.

3. Medium What property of liquids does distillation use to separate them?

   ▼ **Show answer**

   Their different boiling points.

COMMON MISTAKES

- Calling salt water a compound; it is a mixture (a solution) because the ratio can vary.
- Measuring the solvent distance from the bottom of the paper instead of from the starting line.
- Confusing a physical change (melting, dissolving) with a chemical change (burning, rusting).

> **Make it a game · Sort It Fast.** A parent names a material (milk, air, copper wire, sugar, salad dressing, diamond). The student shouts element, compound, homogeneous or heterogeneous. Argue the close calls (milk looks uniform but is a colloid, heterogeneous under a microscope).

> **Scripture.** "And he shall sit as a refiner and purifier of silver: and he shall purify the sons of Levi, and purge them as gold and silver, that they may offer unto the LORD an offering in righteousness." (Malachi 3:3, KJV)
>
> A refiner of silver separates the pure metal from dross by heat. The image of refining runs through Scripture as a picture of making something pure.

UNIT 2

Atoms

Nobody has ever seen an atom with the naked eye, yet chemists were sure atoms existed long before microscopes could image them. This unit follows the evidence: Lavoisier's balance showed that mass is conserved; Proust found that compounds have fixed compositions; Dalton explained both with indivisible atoms; then Thomson found the electron inside the atom and Rutherford found the nucleus. The unit ends with isotopes and the mass spectrograph that weighs atoms one at a time.

Each step is a model that explained the evidence of its day and was later refined. Learning why each model was believed, and why it was changed, is as important as learning the final picture.

PREREQUISITES

- Unit 1: significant figures, density and dimensional analysis
- Scientific notation
- Percents

MASTERY CHECKLIST

- ☐ I can state the law of conservation of mass and use it in calculations.
- ☐ I can explain the laws of definite and multiple proportions and Dalton's atomic theory.
- ☐ I can describe Thomson's and Rutherford's experiments and what each showed.
- ☐ I can find protons, neutrons and electrons from isotope notation.
- ☐ I can calculate average atomic mass from isotope abundances.

LESSONS IN THIS UNIT

2.1 Lavoisier and the Conservation of Mass (Traité, 1789) — 4 sessions of 45 minutes

2.2 Proust's Definite Proportions and Dalton's Atomic Theory (1808)

4 sessions of 45 minutes

2.3 Thomson's Electron (1897) and Rutherford's Gold Foil (1909–1911)

4 sessions of 45 minutes

2.4 Isotopes, Atomic Mass and Aston's Mass Spectrograph (1919)

4 sessions of 45 minutes

Unit 2 test

Show your work. For every calculation write the equation, substitute with units and give the answer with its unit and sensible significant figures. The test is worth 100 points.

1. **7 pts** **Lesson 2.1**

   What is the difference between a reactant and a product?

2. **7 pts** **Lesson 2.1**

   Name Lavoisier's great textbook and its year.

3. **7 pts** **Lesson 2.1**

   10.93 g of steel wool burns completely in air, and the product weighs 13.83 g. What mass of oxygen combined with the iron?

4. **7 pts** **Lesson 2.2**

   What is percent composition?

5. **7 pts** **Lesson 2.2**

   Why do the four points of Dalton's theory still matter today?

6. **7 pts** **Lesson 2.2**

   Water is always 11.2% hydrogen by mass. How many grams of hydrogen are in 52.1 g of water? How many grams of oxygen?

7. **7 pts** **Lesson 2.3**

   What is an alpha particle?

8. **7 pts** **Lesson 2.3**

   A neutral atom has 11 protons. How many electrons does it have?

9. **7 pts** **Lesson 2.4**

Bromine's two isotopes have masses 78.92 u and 80.92 u and its average is 79.90 u. Estimate the abundances.

10. **6 pts** **Lesson 2.4**

What is the mass number of an atom with 17 protons and 18 neutrons?

11. **6 pts** **Lesson 2.4**

Natural gallium has two isotopes: mass 68.926 u (60.11%) and mass 70.925 u (39.89%). Calculate the average atomic mass.

LAB AND DATA

12. **10 pts**

In a sealed bag, baking soda and vinegar had a total mass of 152.36 g before reacting and 152.35 g after. In an open cup the masses were 151.90 g before and 149.62 g after. (a) Find the change in each. (b) Explain the difference. (c) Does the open-cup result disprove conservation of mass?

SHORT ESSAY

13. **15 pts**

Explain how Rutherford's gold-foil results forced a change from Thomson's "plum pudding" model of the atom. Describe what was expected, what was observed, and what Rutherford concluded.

ANSWER KEY

1. Reactants are the starting substances; products are the new substances formed.
2. *Traité élémentaire de chimie*, 1789.
3. 2.90 g
   Mass of oxygen = mass of product − mass of iron = 13.83 − 10.93 = 2.90 g
4. The percent by mass of each element in a compound.
5. Atoms combining in whole-number ratios and being rearranged (not created or destroyed) in reactions remain the basis of chemical formulas and equations.
6. 5.84 g H, 46.3 g O
   $52.1 \times 0.112 = 5.84$ g H · Oxygen = 52.1 − 5.84 = 46.3 g
7. A helium nucleus (two protons and two neutrons) emitted by some radioactive atoms.
8. 11
9. About 51% bromine-79 and 49% bromine-81.
10. 35
11. 69.72 u
    $(68.926)(0.6011) + (70.925)(0.3989) = 69.72$
12. (a) Sealed: −0.01 g (within scale error); open: −2.28 g. (b) In the open cup the carbon dioxide gas escaped into the air. (c) No; the missing mass left as gas. In a closed system mass is conserved.
13. Look for: Thomson's model (positive sphere with electrons embedded) predicted alpha particles would pass nearly straight through; Geiger and Marsden observed most passed through but about 1 in 8000 bounced back at large angles; Rutherford concluded the positive charge and most mass sit in a tiny, dense nucleus with mostly empty space around it; quoting his "15-inch shell" remark is a plus.

Printable answer key (PDF, for the parent)

2.1 Lavoisier and the Conservation of Mass (Traité, 1789)

OBJECTIVES

- State the law of conservation of mass and apply it to reactions.
- Explain how Lavoisier used careful weighing to overturn the phlogiston theory.
- Distinguish open and closed systems in an experiment.

> **Warm-up.** Weigh a sealed plastic bottle of soda. Shake it, then open it slowly and let it fizz for ten minutes, then cap and weigh it again. Predict and explain the result.

WATCH

Law of Conservation of Mass - Fundamental Chemical Laws, Chemistry

The Organic Chemistry Tutor · 3:14

https://www.youtube.com/watch?v=eBTNzScLUg4

> **WHEN TO WATCH**
> After the story, as a summary of the law.

BACKUP VIDEO

GCSE Chemistry - Conservation of Mass

Cognito · 4:33

https://www.youtube.com/watch?v=YZi5DnMX0XM

Before the lab.

KEY TERMS

law of conservation of mass

In a chemical reaction mass is neither created nor destroyed: total mass of products equals total mass of reactants.

reactant

A starting substance in a chemical reaction.

product

A substance formed in a chemical reaction.

closed system

A system in which no matter enters or leaves, such as a sealed container.

phlogiston

A substance once imagined to be released when things burn; Lavoisier showed it does not exist.

combustion

Rapid reaction of a substance with oxygen, releasing heat and light.

CORE IDEAS

The story. In the 1700s chemists explained burning by *phlogiston*, an invisible "fire substance" supposedly released by anything that burned. But metals got heavier when they burned or rusted. Antoine Lavoisier, a French tax official who spent his fortune on precise balances, set out to weigh everything. Heating tin in sealed glass vessels, he found the total mass did not change, but when he opened the vessel air rushed in and the tin had gained mass. Something from the air had combined with the metal.

How they figured it out. In 1774 the English minister Joseph Priestley heated mercury oxide and collected a gas in which a candle burned brightly and a mouse lived longer; Carl Scheele in Sweden had found it slightly earlier. Lavoisier repeated these experiments with his balance and named the gas oxygen. He showed that burning and rusting are combinations with oxygen, not losses of phlogiston. In his *Traité élémentaire de chimie* (1789) he stated the principle: in every operation an equal quantity of matter exists before and after. His wife, Marie-Anne Paulze Lavoisier, kept his laboratory records and drew the engraved plates of his apparatus.

The chemistry. Mass is conserved in every chemical reaction: the atoms are rearranged, not created or destroyed. If a reaction seems to gain or lose mass, look for a gas entering or leaving. In a *closed system* (a sealed bag or flask) the total mass stays constant within the error of the scale. This law is why chemical equations must be balanced (Lesson 5.1).

Where it is used. Engineers use mass balances to design chemical plants and water treatment; nutritionists use them to track what the body takes in and gives off. A car

burns about 3 kg of gasoline for every 9 kg of carbon dioxide it exhausts, because oxygen from the air adds mass to the products.

> **Safety first.** Goggles on. The bag can pop if over-filled; use the amounts given. Vinegar stings eyes.

LAB

Conservation of Mass: Sealed Bag versus Open Cup

You will react baking soda and vinegar in a sealed zip bag and in an open cup and compare the mass changes.

You will need

- Baking soda, vinegar, two zip-top bags, a small cup that fits inside a bag, a plastic cup
- Scale (0.01 g if possible)
- Goggles

Steps

1. Put 1 tablespoon of baking soda in a zip bag. Put 30 mL of vinegar in a small cup and stand it upright in the bag. Seal the bag and weigh the whole thing.
2. Without opening, tip the cup so the vinegar mixes with the baking soda. When the fizzing stops, weigh again.
3. Repeat with the same amounts in an open plastic cup: weigh the cup of vinegar and the baking soda together before and after mixing.
4. Record all masses and the changes.

Check your work

- ☐ Was the mass change in the sealed bag within your scale's error?
- ☐ Where did the missing mass go in the open cup?
- ☐ Could the bag's puffing up have affected the reading? (Buoyancy of the larger bag in air is a tiny real effect.)

LAB NOTEBOOK

1. Data table for both trials with masses before, after and change.
2. A labeled before-and-after sketch of the sealed bag.

KEY FORMULAS

$$m_{\text{reactants}} = m_{\text{products}}$$

> **From history.** Lavoisier was guillotined in 1794 during the Terror because he had been a tax collector. The judge's reply "The Republic has no need of scientists" is probably a legend, but the mathematician Lagrange truly said: "It took them only an instant to cut off that head, and a hundred years may not produce another like it."

WORKED EXAMPLES

Example 1. When 10.0 g of calcium carbonate is heated, it forms calcium oxide and carbon dioxide. If 5.6 g of calcium oxide remains, what mass of carbon dioxide was given off?

▼ **Show solution**

1. Conservation of mass: $10.0\ \text{g} = 5.6\ \text{g} + m_{\text{CO}_2}$
2. $m_{\text{CO}_2} = 10.0 - 5.6 = 4.4\ \text{g}$

Answer: 4.4 g

Example 2. A burning candle loses mass, but rusting steel wool gains mass. Explain both with conservation of mass.

▼ **Show solution**

1. Candle: wax + oxygen → carbon dioxide + water vapor; the gaseous products escape, so the candle side loses mass.
2. Steel wool: iron + oxygen → iron oxide; the solid product keeps the oxygen it gained from the air, so it weighs more.
3. In a sealed jar neither total mass would change.

Answer: Gases leave the candle; oxygen joins the iron.

PRACTICE

1. **Easy** State the law of conservation of mass.

   ▼ **Show answer**

   In a chemical reaction mass is neither created nor destroyed; reactants and products have equal total mass.

2. **Easy** In a sealed bag, 19.71 g of baking soda reacts with 22.09 g of vinegar. Afterward the bag holds liquid, a dissolved salt and 1.39 g of carbon dioxide gas. What is the total mass of everything in the bag after the reaction?

   ▼ **Show answer**

   41.80 g

3. **Medium** Why did metals gaining mass when they burned trouble the phlogiston theory?

   ▼ **Show answer**

   If burning released phlogiston, the metal should have become lighter, not heavier.

4. **Medium** 14.86 g of steel wool burns completely in air, and the product weighs 21.00 g. What mass of oxygen combined with the iron?

   ▼ **Show answer**

   6.14 g

5. **Medium** 7.19 g of steel wool burns completely in air, and the product weighs 9.87 g. What mass of oxygen combined with the iron?

   ▼ **Show answer**

   2.68 g

COMMON MISTAKES

- Concluding that mass was destroyed when a gas simply escaped.
- Forgetting the mass of the container when weighing before and after.
- Thinking mass and weight change during a reaction; only the arrangement of atoms changes.

Make it a game · Where Did It Go?. A parent describes a change (a log burns to ash, bread dough rises, ice melts in a sealed jar). The student says whether the visible mass goes up, down or stays the same, and explains where the matter went. Two points for a correct reason.

Scripture. “The thing that hath been, it is that which shall be; and that which is done is that which shall be done: and there is no new thing under the sun.” (Ecclesiastes 1:9, KJV)

Chemistry rearranges what already exists; no reaction makes matter from nothing. Only God creates out of nothing.

UNIT 2 · ATOMS · LESSON 2.2 OF 36 · 4 SESSIONS OF 45 MINUTES

2.2 Proust’s Definite Proportions and Dalton’s Atomic Theory (1808)

OBJECTIVES

- State the laws of definite and multiple proportions.
- List the main points of Dalton’s atomic theory and the parts later revised.
- Use percent composition to test whether two samples are the same compound.

Warm-up. Make two batches of trail mix: one with 2 cups of nuts to 1 cup of raisins and one with 3 cups of nuts to 1 cup of raisins. Are they the same mixture? Now compare water from a river and from a glacier. Is the water itself different?

WATCH

Dalton's Atomic Theory

The Organic Chemistry Tutor · 6:27

https://www.youtube.com/watch?v=rbmPpgASUAY

> **WHEN TO WATCH**
> After the core ideas.

BACKUP VIDEO

Law of Definite Proportions Chemistry Practice Problems

The Organic Chemistry Tutor · 10:31

https://www.youtube.com/watch?v=ly0ywRdVG_M

Optional, for more practice on definite proportions.

KEY TERMS

law of definite proportions

A given compound always contains the same elements in the same proportion by mass, wherever it comes from (Proust, 1799).

law of multiple proportions

When two elements form more than one compound, the masses of one that combine with a fixed mass of the other are in small whole-number ratios (Dalton).

atom

The smallest particle of an element that keeps the element's chemical identity.

atomic theory

Dalton's explanation that matter is made of atoms that combine in whole-number ratios and rearrange in reactions.

percent composition

The percent by mass of each element in a compound.

CORE IDEAS

The story. Around 1799 the French chemist Joseph Proust, working in Spain, analyzed copper carbonate made in his lab and found in nature. Every sample had the same proportions of copper, carbon and oxygen by mass. His rival Claude Berthollet argued that compounds could vary in composition; their debate lasted years, and careful analysis proved Proust right for true compounds. A compound has a fixed recipe.

How they figured it out. John Dalton, a Quaker schoolteacher in Manchester who also studied weather and his own color blindness, asked why compounds have fixed proportions. In his *New System of Chemical Philosophy* (1808) he proposed that each element is made of identical atoms with a characteristic mass, and compounds form when atoms combine in small whole-number ratios. That explained Proust's law, and it predicted a new one: carbon monoxide and carbon dioxide have oxygen masses in a 1 : 2 ratio for the same mass of carbon, because one has one oxygen atom per carbon and the other has two. Dalton drew atoms as circles with symbols and published the first table of relative atomic weights.

The chemistry. Dalton's theory: (1) elements are made of tiny particles called atoms; (2) atoms of one element are alike in mass and properties and differ from atoms of other elements; (3) atoms combine in simple whole-number ratios to form compounds; (4) in reactions atoms are rearranged, never created or destroyed. Later discoveries revised parts: atoms are divisible into protons, neutrons and electrons (Lessons 2.3–2.4), and atoms of the same element can differ in mass (isotopes). But the core, whole-number combining of atoms, still stands.

Where it is used. Quality control in medicine and food relies on fixed composition: pure aspirin is always 60.0% carbon by mass, so a sample that is not has impurities. Forensic chemists identify substances by their composition.

LAB

Whole-Number Ratios with Paper Clips and Washers

A model of Dalton's reasoning: you will "react" two kinds of hardware in fixed ratios and discover that the masses combine in whole-number patterns.

You will need

- Large paper clips and steel washers (or two kinds of beads or coins)
- Scale
- Zip bags

Steps

1. Weigh 10 paper clips ("element X") and 10 washers ("element Y") to find the mass of one of each.
2. Make "compound XY": 5 bags, each with 1 clip and 1 washer. Weigh each bag and compute the percent of mass that is X.
3. Make "compound XY_2" (1 clip, 2 washers) and "X_2Y" bags. Compute percent X for each.
4. For a fixed mass of X, compare the masses of Y in XY and XY_2.

Check your work

- ☐ Is the percent composition the same for every bag of the same "compound"?
- ☐ Is the Y-mass ratio between XY_2 and XY a whole number?
- ☐ How is this model like and unlike real atoms?

LAB NOTEBOOK

1. Data table of bag masses and percent X for each "compound."
2. Dalton's four points written in your own words, with one revision for each that later science made.

KEY FORMULAS

$$\%\ \text{element} = \frac{\text{mass of element}}{\text{mass of compound}} \times 100\%$$

From history. Dalton was color-blind and wrote the first scientific paper on it (1794); color blindness was long called "Daltonism." He asked that his eyes be examined after his death, and in 1995 DNA from his preserved eye confirmed he lacked the green-sensitive pigment.

WORKED EXAMPLES

Example 1. Sample A of a white powder contains 2.30 g sodium and 3.55 g chlorine. Sample B contains 4.60 g sodium and 7.10 g chlorine. Could both be the same compound?

▼ **Show solution**

1. A: Na fraction $= 2.30/(2.30 + 3.55) = 0.393$, so 39.3% Na.
2. B: $4.60/(4.60 + 7.10) = 0.393$, so 39.3% Na.
3. Same percent composition, consistent with the same compound (sodium chloride).

Answer: Yes; both are 39.3% sodium.

Example 2. Nitrogen forms two oxides. In the first, 7.00 g N combines with 8.00 g O; in the second, 7.00 g N combines with 16.0 g O. Show this fits the law of multiple proportions.

▼ **Show solution**

1. Same mass of nitrogen (7.00 g) in both.
2. Oxygen ratio: $16.0/8.00 = 2.00$, a small whole number.
3. The compounds are NO and NO_2.

Answer: Ratio 2 : 1

PRACTICE

1. Easy State the law of definite proportions.

   ▼ **Show answer**

   A compound always contains the same elements in the same proportion by mass.

2. Medium How did Dalton's atomic theory explain the law of definite proportions?

   ▼ **Show answer**

   Compounds form from atoms of fixed mass combining in fixed whole-number ratios, so the mass proportions are always the same.

3. Medium Water is always 11.2% hydrogen by mass. How many grams of hydrogen are in 15.2 g of water? How many grams of oxygen?

   ▼ **Show answer**

   1.70 g H, 13.5 g O

4. Medium Water is always 11.2% hydrogen by mass. How many grams of hydrogen are in 16.7 g of water? How many grams of oxygen?

   ▼ **Show answer**

   1.87 g H, 14.8 g O

5. Hard Carbon forms two oxides. In the first, 3.25 g of carbon combines with 4.33 g of oxygen; in the second, 3.25 g of carbon combines with 8.66 g of oxygen. Find the ratio of the oxygen masses and say which law it illustrates.

   ▼ **Show answer**

   2 : 1, the law of multiple proportions (CO and CO_2)

COMMON MISTAKES

- Thinking definite proportions means equal numbers of atoms; it means a fixed ratio by mass.
- Comparing oxygen masses for different masses of carbon; fix the mass of one element first.
- Treating Dalton's theory as wrong because parts were revised; its core is still the basis of chemistry.

> **Make it a game · Same or Different?.** A parent gives two sample analyses (masses of each element). The student computes percent composition and says "same compound" or "different" within one minute. Five rounds.

> **Scripture.** “He telleth the number of the stars; he calleth them all by their names.” (Psalm 147:4, KJV)
>
> Order and number are written into creation. Dalton’s small whole numbers are part of that order.

UNIT 2 · ATOMS · LESSON 2.3 OF 36 · 4 SESSIONS OF 45 MINUTES

2.3 Thomson’s Electron (1897) and Rutherford’s Gold Foil (1909–1911)

OBJECTIVES

- Describe the cathode-ray experiment and what it showed about the electron.
- Describe the gold-foil experiment and how it led to the nuclear model.
- Compare the charge and mass of the proton, neutron and electron.
- Explain how scientific models change with new evidence.

> **Warm-up.** Rub a balloon on your hair and hold it near a thin stream of tap water. What happens? What does this tell you about charges in ordinary matter?

WATCH

Discovery of the Electron: Cathode Ray Tube Experiment

Tyler DeWitt · 11:08

https://www.youtube.com/watch?v=Rb6MguN0Uj4

> **WHEN TO WATCH**
> After the story of Thomson.

BACKUP VIDEO

Discovery of the Nucleus: Rutherford's Gold Foil Experiment

Tyler DeWitt · 15:59

https://www.youtube.com/watch?v=dNp-vP17asI

After the story of Rutherford (watch from the beginning; about 16 minutes).

KEY TERMS

electron

A tiny negatively charged particle found outside the nucleus, about 1/1836 the mass of a proton.

cathode ray

A beam of electrons traveling from the negative electrode in a vacuum tube.

nucleus

The tiny, dense, positively charged center of an atom containing protons and neutrons.

proton

A positively charged particle in the nucleus; the number of protons is the atomic number.

neutron

A neutral particle in the nucleus with about the same mass as a proton (discovered by Chadwick, 1932).

alpha particle

A positively charged particle (a helium nucleus) given off by some radioactive atoms.

CORE IDEAS

The story. In the 1890s physicists sent electricity through glass tubes with most of the air pumped out. A glowing beam, a "cathode ray," streamed from the negative electrode. Was it a wave of light or a stream of particles? At the Cavendish Laboratory in Cambridge, J. J. Thomson bent the beam with electric and magnetic fields in 1897. It bent toward the positive plate, so it was negatively charged, and from how much it bent he found its charge-to-mass ratio: about 1800 times that of a hydrogen ion. The particles were far lighter than any atom and were the same no matter what metal the electrodes were made of. Thomson had found the first subatomic particle, the electron.

How they figured it out. Since atoms are neutral, Thomson pictured the electrons embedded in a sphere of positive charge, like plums in a pudding. Ernest Rutherford, a New Zealander working in Manchester, tested this. His assistants Hans Geiger and Ernest Marsden fired alpha particles at gold foil a few hundred atoms thick and counted the flashes on a screen in a dark room. Most went straight through, as expected, but about 1 in 8000 bounced back at large angles. Rutherford said it was "as if you fired a 15-inch shell at a piece of tissue paper and it came back and hit you." In 1911 he concluded that the atom's positive charge and almost all its

mass are packed into a tiny nucleus, about 1/100 000 the diameter of the atom, with electrons moving in mostly empty space around it.

The chemistry. The atom has three main particles. Protons (charge +1, mass about 1 u) and neutrons (charge 0, mass about 1 u) make up the nucleus. Electrons (charge −1, mass about 1/1836 u) occupy the space around it. A neutral atom has equal numbers of protons and electrons. The number of protons, the *atomic number*, defines the element. In 1913 Henry Moseley showed by X-rays that atomic number, not mass, is the true basis of the periodic table.

Where it is used. Old television sets and computer monitors were cathode-ray tubes: an electron beam steered by magnets painted the picture. Electron microscopes, X-ray machines and the electron beams that sterilize medical supplies all descend from Thomson's tube.

LAB

Probing a Hidden Shape: A Marble "Gold Foil" Model

Like Rutherford, you will learn the shape of something you cannot see by rolling marbles at it and watching how they bounce.

You will need

- A large piece of cardboard or a tray propped on four books to make a low "roof"
- A cardboard cutout shape (circle, triangle or square, 10–15 cm) taped under the roof by a helper
- 10 marbles, a sheet of paper and pencil

Steps

1. A helper hides a shape under the cardboard without showing you.
2. Roll marbles under the roof from all four sides, one at a time. Record where each enters and where it comes out.
3. Draw the paths on paper. Where marbles deflect, mark a point of the hidden edge.
4. After 50 rolls, sketch your guess of the shape, then lift the cover.
5. Repeat with a smaller shape. How many more rolls do you need to find it?

Check your work

- ☐ How did you infer something you could not see?
- ☐ Why did the smaller target take more rolls to find?
- ☐ How is this like the gold-foil experiment? How is it different?

LAB NOTEBOOK

1. Your marble-path diagram and your shape guess beside the real shape.
2. A table comparing proton, neutron and electron (charge, mass, location, discoverer).

From history. Rutherford trained an astonishing number of future Nobel laureates, including Chadwick (the neutron) and Bohr. Thomson's son, G. P. Thomson, later won a Nobel Prize for showing that electrons also behave as waves, so father and son proved the electron is a particle and a wave.

SEE IT: INTERACTIVE.

Quantum Explorer: Size and Pauli

Zoom from a whole hydrogen atom down to the nucleus at true scale.

The atom is about 63,000 times wider than its nucleus, which is why almost every alpha particle went straight through Rutherford's gold foil. Opens in a new tab on the Size and Pauli tab. Worked answers for the parent: answer key (PDF).

WORKED EXAMPLES

Example 1. If an atom were enlarged so its nucleus were the size of a marble (1 cm across), about how big would the whole atom be? (The atom is about 100 000 times wider than its nucleus.)

▼ **Show solution**

1. $1\ \text{cm} \times 100\,000 = 100\,000\ \text{cm}$
2. $= 1000\ \text{m} = 1\ \text{km}$

Answer: About 1 km across: a marble in the middle of a space as wide as a small town.

Example 2. Why did most alpha particles go straight through the gold foil, but a few bounced back?

▼ **Show solution**

1. Most of the atom is empty space, so most alpha particles met nothing.
2. A few came very close to a nucleus, where the concentrated positive charge repelled the positive alpha particle strongly.
3. Because the nucleus is tiny, such close approaches were rare (about 1 in 8000).

Answer: Mostly empty space with a tiny, dense, positive nucleus.

PRACTICE

1. Easy What is the charge of an electron, a proton and a neutron?

   ▼ **Show answer**

   Electron −1, proton +1, neutron 0.

2. Medium How did Thomson show that cathode rays are negatively charged?

   ▼ **Show answer**

   The beam bent toward the positive plate (and away from the negative) in an electric field.

3. Medium Describe Thomson's "plum pudding" model.

   ▼ **Show answer**

   A sphere of positive charge with negative electrons embedded throughout.

COMMON MISTAKES

- Thinking the nucleus fills most of the atom; it is about 1/100 000 of its diameter.
- Thinking electrons have about the same mass as protons; they are about 1836 times lighter.
- Saying Rutherford found the neutron; Chadwick found it in 1932.

> **Make it a game · Particle Cards.** Make cards: proton, neutron, electron. A parent reads a clue ("charge −1," "found by Chadwick," "in the nucleus, no charge"), and the student holds up the right card. One point each, ten clues.

> **Scripture.** "Through faith we understand that the worlds were framed by the word of God, so that things which are seen were not made of things which do appear." (Hebrews 11:3, KJV)
>
> The visible world is made of things that do not appear. Chemists reason carefully from evidence to what they cannot see.

2.4 Isotopes, Atomic Mass and Aston's Mass Spectrograph (1919)

OBJECTIVES

- Use isotope notation to find protons, neutrons and electrons.
- Explain why atomic masses on the periodic table are not whole numbers.
- Calculate average atomic mass from isotope masses and abundances.
- Describe how a mass spectrometer separates atoms by mass.

> **Warm-up.** A bag holds 75 marbles that weigh 35 g each and 25 that weigh 37 g each. What is the average mass of a marble? Why is it closer to 35 than to 37?

WATCH

What are Isotopes?

Tyler DeWitt · 7:56

https://www.youtube.com/watch?v=EboWeWmh5Pg

> **WHEN TO WATCH**
> After the core ideas on isotope notation.

BACKUP VIDEO

How to Calculate Atomic Mass Practice Problems

Tyler DeWitt · 6:11

https://www.youtube.com/watch?v=ULRsJYhQmlo

Before the worked examples.

KEY TERMS

isotopes

Atoms of the same element (same number of protons) with different numbers of neutrons.

mass number

The total number of protons and neutrons in a nucleus.

atomic mass unit

One-twelfth the mass of a carbon-12 atom, symbol u.

average atomic mass

The weighted average mass of an element's naturally occurring isotopes, as printed on the periodic table.

mass spectrometer

An instrument that ionizes atoms or molecules and separates them by mass in a magnetic or electric field.

ion

An atom or group of atoms with a net electric charge from gaining or losing electrons.

CORE IDEAS

The story. If Dalton was right that all atoms of an element have the same mass, why is chlorine's atomic mass 35.45, not a whole number? Frederick Soddy, studying radioactive elements, suggested in 1913 that one element could have atoms of different masses; he called them *isotopes* ("same place" in the periodic table). In 1919 Francis Aston at Cambridge built the mass spectrograph and proved it: neon split into two beams of masses 20 and 22, and chlorine into 35 and 37. Aston went on to find 212 natural isotopes and won the 1922 Nobel Prize.

How they figured it out. In a mass spectrometer, atoms are given a charge, accelerated, and sent through a magnetic field. Lighter ions bend more, heavier ions bend less, so each isotope lands in a different place. The height of each peak shows how abundant the isotope is. Once neutrons were discovered (1932), the cause was clear: isotopes have the same number of protons but different numbers of neutrons.

The chemistry. Isotope notation writes the mass number A above and the atomic number Z below the symbol: ${}^{37}_{17}\mathrm{Cl}$. Neutrons $= A - Z$. Isotopes of an element have the same chemistry, because chemistry depends on electrons, but different masses. The periodic-table mass is the weighted average: multiply each isotope's mass by its fractional abundance and add. Chlorine is 75.78% chlorine-35 and 24.22% chlorine-37, so its average is 35.45 u, closer to 35.

Where it is used. Mass spectrometers test athletes for drugs, find the age of artifacts, detect explosives at airports and identify proteins. Isotopes are tracers in medicine: iodine-131 treats thyroid disease, and heavy-water and oxygen-18 studies track how the body uses water.

LAB

“Beanium”: Isotopes in a Bag of Beans

You will model an element whose isotopes are three kinds of beans, then calculate its average atomic mass just as a mass spectrometer would.

You will need

- A bag with a mix of three kinds of dry beans (for example kidney, pinto, navy), about 100 total
- Scale reading to 0.01 g or 0.1 g
- Calculator

Steps

1. Sort the beans into the three “isotopes.” Count each kind.
2. Weigh each group and find the average mass of one bean of each kind.
3. Calculate the percent abundance of each isotope.
4. Calculate the weighted average mass of one “beanium atom.”
5. Check: total mass of all beans ÷ total number of beans should give the same answer.

Check your work

- ☐ Do your two methods agree?
- ☐ Why is the average closest to the most abundant isotope?
- ☐ Why are real isotope masses measured, not just counted like beans?

LAB NOTEBOOK

1. Beanium data table and both calculations of average mass.
2. A sketch of a mass spectrometer showing heavier and lighter ions bending.

KEY FORMULAS

$$\text{neutrons} = A - Z$$

$$\text{average atomic mass} = \sum(\text{isotope mass}) \times (\text{fractional abundance})$$

> **From history.** Aston was also a skilled amateur athlete and a keen photographer who built his own instruments. Soddy, who named isotopes, later worried about the military use of atomic energy and wrote on economics and the responsibilities of science.

WORKED EXAMPLES

Example 1. How many protons, neutrons and electrons are in $^{56}_{26}\mathrm{Fe}^{3+}$?

▼ **Show solution**

1. Protons $= Z = 26$
2. Neutrons $= 56 - 26 = 30$
3. The 3+ charge means 3 electrons lost: $26 - 3 = 23$ electrons

Answer: 26 p, 30 n, 23 e

Example 2. Copper is 69.15% copper-63 (62.930 u) and 30.85% copper-65 (64.928 u). Find its average atomic mass.

▼ **Show solution**

1. $(62.930)(0.6915) = 43.516$
2. $(64.928)(0.3085) = 20.030$
3. Sum $= 63.546$, so 63.55 u

Answer: 63.55 u

PRACTICE

1. Easy What do all isotopes of carbon have in common? How do they differ?

   ▼ **Show answer**

   Same number of protons (6); different numbers of neutrons.

2. Easy Write the symbol for the isotope with 8 protons and 10 neutrons.

   ▼ **Show answer**

   $^{18}_{8}\mathrm{O}$ (oxygen-18)

3. **Easy** How many protons, neutrons and electrons are in a neutral atom of carbon-14 (atomic number 6)?

   ▼ **Show answer**

   6 protons, 8 neutrons, 6 electrons

4. **Easy** How many protons, neutrons and electrons are in a neutral atom of iron-56 (atomic number 26)?

   ▼ **Show answer**

   26 protons, 30 neutrons, 26 electrons

5. **Medium** Natural chlorine has two isotopes: mass 34.969 u (75.78%) and mass 36.966 u (24.22%). Calculate the average atomic mass.

   ▼ **Show answer**

   35.45 u

COMMON MISTAKES

- Using percent (75.78) instead of the fraction (0.7578) in the weighted average.
- Confusing mass number (a whole count) with atomic mass (a measured average).
- Thinking isotopes have different chemistry; they react the same way because they have the same electrons.

REFERENCE GUIDE

An eight-page, color-coded Organic Chemistry guide: carbon bonding and hybridization, structural formulas, functional groups, naming, isomers and chirality, acids and bases with pKa values, reaction types and mechanisms, spectroscopy (IR and NMR), and the molecules of life. The booklet prints on 11×17 paper and folds to letter size. Prints in color and reads clearly in black and white.

> **Make it a game · Isotope Bingo.** Make bingo cards with neutron counts (6–20). A parent calls an isotope (“carbon-14,” “sodium-23,” “chlorine-37”), the student finds its neutron number and marks it. First bingo wins.

> **Scripture.** “A just weight and balance are the LORD’s: all the weights of the bag are his work.” (Proverbs 16:11, KJV)
>
> Accurate weights belong to the Lord. Aston’s instrument weighed atoms one at a time.

UNIT 3

Electrons and the Periodic Table

Why do sodium and potassium both explode in water, while neon does nothing at all? The answer lies in the electrons. This unit begins with the colored flames and bright-line spectra that let Bunsen and Kirchhoff discover new elements, and with Niels Bohr’s idea that electrons occupy fixed energy levels. It then teaches electron configurations, Mendeleev’s periodic table and the predictions that made it famous, and the trends in size, ionization energy and electronegativity that let chemists predict how an element will behave.

The flame-test lab is the most dramatic of the year and requires an adult at the stove or burner.

PREREQUISITES

- Unit 2: protons, neutrons, electrons and atomic number
- Scientific notation and simple algebra

MASTERY CHECKLIST

- ☐ I can explain why each element gives a characteristic flame color and line spectrum.
- ☐ I can write electron configurations and count valence electrons for elements 1–36.
- ☐ I can tell how Mendeleev arranged the elements and why his predictions mattered.
- ☐ I can predict trends in atomic radius, ionization energy and electronegativity across a period and down a group.

LESSONS IN THIS UNIT

3.1 Flame, Spectra and Bohr’s Atom: Bunsen and Kirchhoff (1859–1860)

4 sessions of 45 minutes

3.2 Electron Configurations and Orbitals 4 sessions of 45 minutes

3.3 Mendeleev’s Table and the Missing Elements (1869; Gallium 1875)

4 sessions of 45 minutes

3.4 Periodic Trends: Size, Ionization Energy and Pauling’s Electronegativity

4 sessions of 45 minutes

Unit 3 test

Show your work. For every calculation write the equation, substitute with units and give the answer with its unit and sensible significant figures. The test is worth 100 points.

1. **8 pts** **Lesson 3.1**
   Why is a continuous spectrum produced by a hot solid but a line spectrum by a hot gas?

2. **8 pts** **Lesson 3.1**
   Which law connects the speed, wavelength and frequency of light?

3. **8 pts** **Lesson 3.1**
   Light from a flame test has wavelength 656 nm. Find its frequency. ($c = 3.00 \times 10^8$ m/s)

4. **8 pts** **Lesson 3.2**
   What is the difference between Bohr’s orbits and orbitals?

5. **8 pts** **Lesson 3.2**
   How many valence electrons does neon have?

6. **7 pts** **Lesson 3.2**

   How many valence electrons does potassium (Z = 19) have?

7. **7 pts** **Lesson 3.3**

   What element is named after Mendeleev?

8. **7 pts** **Lesson 3.3**

   Predict the formula of the compound of rubidium and chlorine, and explain.

9. **7 pts** **Lesson 3.4**

   Where on the periodic table are the elements with the largest atoms?

10. **7 pts** **Lesson 3.4**

    Why do noble gases have very high ionization energies?

LAB AND DATA

11. **10 pts**

    In a flame test, a solution of an unknown salt gives a bright crimson-red flame. Lithium gives crimson, sodium yellow-orange, potassium lilac and copper green. (a) Identify the likely metal. (b) Why does each metal give a different color? (c) Why can a trace of sodium spoil the test?

SHORT ESSAY

12. **15 pts**

    Explain why Mendeleev's periodic table was accepted by chemists. Include his gaps and predictions and what happened when gallium was discovered.

ANSWER KEY

1. In a hot solid atoms are packed and interact, giving all wavelengths; in a thin gas isolated atoms emit only their own line wavelengths.
2. $c = \lambda f$
3. 4.57×10^{14} Hz

   $f = c/\lambda = (3.00 \times 10^{8})/(656 \times 10^{-9}) = 4.57 \times 10^{14}$ Hz
4. Orbits are fixed circular paths; orbitals are regions of probability from wave mechanics.
5. 8
6. 1

   Configuration $1s^2 2s^2 2p^6 3s^2 3p^6 4s^1$; count the electrons in the highest shell (n = 4).
7. Mendelevium (element 101).
8. RbCl; rubidium is in group 1 like sodium (NaCl) and forms a 1+ ion.
9. The lower left (cesium, francium).
10. Their outer shells are full and the electrons are held tightly by a large effective nuclear charge.
11. (a) Lithium. (b) Each element has its own set of electron energy levels; electrons excited by the flame fall back and emit photons whose energies (colors) match the gaps. (c) Sodium's yellow is so intense that even a trace (for example from fingers) can hide other colors.
12. Look for: Mendeleev arranged elements by atomic mass and recurring properties (1869); he left gaps and predicted properties of undiscovered elements (eka-aluminum, eka-silicon, eka-boron); gallium (1875), scandium (1879) and germanium (1886) matched his predictions closely, including gallium's density after he urged a re-measurement; successful predictions are strong evidence for a scientific idea.

Printable answer key (PDF, for the parent)

3.1 Flame, Spectra and Bohr's Atom: Bunsen and Kirchhoff (1859–1860)

OBJECTIVES

- Relate wavelength, frequency and energy of light.
- Explain emission spectra with Bohr's energy levels.
- Describe how Bunsen and Kirchhoff used the spectroscope to discover cesium and rubidium.
- Identify metals by flame test.

> **Warm-up.** Look at a streetlight, a candle flame and a phone screen through a diffraction-grating viewer or at the reflection on an old CD. Sketch the colors you see from each.

WATCH

Coloured Flames - Periodic Table of Videos

Periodic Videos · 5:40

https://www.youtube.com/watch?v=O3ZuimXdObE

> **WHEN TO WATCH**
> Before the lab, to see real flame colors.

BACKUP VIDEO

Emission and Absorption Spectra

Bozeman Science · 5:18

https://www.youtube.com/watch?v=1uPyq63aRvg

After the core ideas, to review emission and absorption spectra.

KEY TERMS

wavelength

The distance between successive crests of a wave, λ, measured for light in nanometers.

frequency

The number of wave crests passing a point per second, f, in hertz.

photon

A packet of light energy, $E = hf$.

emission spectrum

The pattern of bright lines of specific wavelengths given off by excited atoms of an element.

energy level

One of the fixed energies an electron may have in an atom (Bohr, 1913).

ground state

The lowest-energy arrangement of an atom's electrons.

CORE IDEAS

The story. In the 1850s Robert Bunsen in Heidelberg designed a burner whose flame was nearly colorless, perfect for studying the colors salts give when heated. His friend Gustav Kirchhoff suggested sending the light through a prism. Each element produced its own set of sharp bright lines, a fingerprint. In 1860–1861, examining mineral water from Dürkheim, they saw two blue lines no known element made and named the new element cesium (Latin *caesius*, sky-blue). Soon after, a dark red line led them to rubidium.

How they figured it out. Why only certain colors? In 1913 Niels Bohr proposed that an electron in an atom can have only certain energies, called energy levels. Heat or electricity bumps an electron to a higher level (an excited state); when it falls back, it gives off a photon whose energy equals the difference between the levels. Since every element has a different set of levels, every element gives off a different set of colors. Bohr's model calculated hydrogen's visible lines exactly. Later quantum theory replaced his circular orbits with orbitals (Lesson 3.2), but energy levels remain.

The chemistry. Light is a wave with speed $c = 3.00 \times 10^{8}$ m/s: $c = \lambda f$. Its energy comes in photons, $E = hf$, with Planck's constant $h = 6.626 \times 10^{-34}$ J·s. Shorter wavelength means higher frequency and more energy: violet photons carry more energy than red. In a flame test, lithium gives crimson, sodium bright yellow-orange (589 nm), potassium lilac, calcium orange-red, strontium red and copper green-blue.

Where it is used. Fireworks get their colors from metal salts (strontium red, barium green, copper blue). Sodium and mercury streetlights glow with their emission lines. Astronomers interpret the dark lines in sunlight and starlight as the

fingerprints of elements; helium was first identified in the solar spectrum in 1868 before it was found on earth.

> **Safety first.** Adult hands-on: only an adult handles the flame. Goggles on, hair tied back, sleeves rolled. Never use alcohol or other flammable liquids. Copper sulfate is toxic: do not touch with bare hands; wash up afterward. Keep a fire extinguisher at hand.

LAB

Flame Tests and a CD Spectroscope

You will identify metals by the color they give a flame and look at spectra of household lights.

You will need

- Wooden splints or bamboo skewers soaked overnight in water
- Table salt (NaCl), salt substitute (KCl), borax, cream of tartar, calcium chloride ice melt, and (optional) copper sulfate root killer
- A gas stove burner or propane torch, held by an adult
- An old CD or a diffraction-grating viewer, a cardboard tube

Steps

1. Lay out each salt on a labeled saucer. Wet a soaked splint tip and dip it in the salt.
2. With an adult, hold the tip in the edge of the flame. Record the color. Use a fresh splint for each salt.
3. Repeat with a mixture of NaCl and KCl. Look at it through blue cobalt glass or blue cellophane if you have it (it hides sodium yellow).
4. Build a spectroscope: tape a CD at an angle in the end of a tube with a narrow slit at the other end. Look at a fluorescent bulb, an LED bulb, an incandescent bulb and a sodium streetlight.
5. Sketch each spectrum: continuous rainbow or separate bright lines?

Check your work

- ☐ Which salts gave the clearest colors?
- ☐ Why did sodium tend to mask the others?
- ☐ Which light sources gave line spectra?

LAB NOTEBOOK

1. Flame-test data table (salt, metal, color) and your identification of an unknown.
2. Sketches of four spectra from your CD spectroscope.

KEY FORMULAS

$$c = \lambda f \qquad E = hf = \frac{hc}{\lambda}$$

$$h = 6.626 \times 10^{-34}\ \text{J}\cdot\text{s} \qquad c = 3.00 \times 10^{8}\ \text{m/s}$$

From history. Kirchhoff's banker reportedly asked what use it was to find gold in the sun if one could not bring it down. When Kirchhoff later received a medal and a gold prize for the work, he is said to have handed it over and replied, "Here is gold from the sun." The story may be embroidered, but the science is real.

SEE IT: INTERACTIVE.

Quantum Explorer: Energy levels and light

Move hydrogen's electron between Bohr's energy levels and watch its bright lines appear.

Each drop gives off one photon of one exact wavelength, and only the drops that end on level 2 are visible: the Balmer lines at 656, 486, 434 and 410 nm. That one-element, one-set-of-lines fingerprint is what Bunsen and Kirchhoff used to find new elements. Opens in a new tab on the Energy levels and light tab. Worked answers for the parent: answer key (PDF).

WORKED EXAMPLES

Example 1. Sodium's yellow light has wavelength 589 nm. Find its frequency and photon energy.

▼ **Show solution**

1. $\lambda = 589 \times 10^{-9}\ \text{m}$
2. $f = c/\lambda = (3.00 \times 10^{8})/(589 \times 10^{-9}) = 5.09 \times 10^{14}\ \text{Hz}$
3. $E = hf = (6.626 \times 10^{-34})(5.09 \times 10^{14}) = 3.37 \times 10^{-19}\ \text{J}$

Answer: 5.09×10^{14} Hz; 3.37×10^{-19} J

Example 2. Which carries more energy per photon: red light (700 nm) or violet light (400 nm)? By what factor?

▼ **Show solution**

1. $E \propto 1/\lambda$
2. Ratio $= 700/400 = 1.75$

Answer: Violet, 1.75 times as much

PRACTICE

1. Easy Which has the higher frequency: red light or blue light?

   ▼ **Show answer**

   Blue light.

2. Medium In Bohr's model, why does an excited atom give off only certain colors?

   ▼ **Show answer**

   Electrons can have only certain energies; when they fall between levels they emit photons whose energies equal the level differences.

3. Medium Light from a flame test has wavelength 589 nm. Find its frequency. ($c = 3.00 \times 10^8$ m/s)

   ▼ **Show answer**

   5.09×10^{14} Hz

4. Hard Find the energy of one photon of 520 nm light. ($h = 6.626 \times 10^{-34}$ J·s, $c = 3.00 \times 10^8$ m/s)

   ▼ **Show answer**

   3.82×10^{-19} J

5. Hard Find the energy of one photon of 656 nm light. ($h = 6.626 \times 10^{-34}$ J·s, $c = 3.00 \times 10^8$ m/s)

   ▼ **Show answer**

   3.03×10^{-19} J

COMMON MISTAKES

- Thinking longer wavelength means more energy; it is the reverse.
- Forgetting to convert nanometers to meters (10^{-9}).
- Saying the electron gives off light when it jumps up; it absorbs energy going up and emits light coming down.

REFERENCE GUIDE

A two-page, landscape Periodic Table of the Elements, colored by the sublevel being filled, with a guide to reading the table, common ions, electron configurations and constants. Prints in color and reads clearly in black and white.

> **Make it a game · Name That Color.** A parent shows a flame color (picture or description) and the student names the metal. Then reverse: the parent names the metal and the student gives the color and one use (fireworks, streetlights).

> **Scripture.** "Who coverest thyself with light as with a garment: who stretchest out the heavens like a curtain:" (Psalm 104:2, KJV)
>
> Light is God's garment in the Psalm. Spectra show that even light carries the order of the elements He made.

UNIT 3 · ELECTRONS AND THE PERIODIC TABLE · LESSON 3.2 OF 36 · 4 SESSIONS OF 45 MINUTES

3.2 Electron Configurations and Orbitals

OBJECTIVES

- Describe s, p and d orbitals and how many electrons each holds.
- Write electron configurations using the Aufbau principle, the Pauli principle and Hund's rule.
- Relate configuration to position on the periodic table and count valence electrons.

> **Warm-up.** A hotel fills rooms from the ground floor up, two guests per room, and guests take empty rooms before sharing. Using that rule, how would 7 guests fill a floor with one single room and then a floor with three double rooms?

WATCH

Electron Configuration - Basic introduction

The Organic Chemistry Tutor · 10:19

https://www.youtube.com/watch?v=NIwcDnFjj98

> **WHEN TO WATCH**
> After the core ideas, before the worked examples.

BACKUP VIDEO

How to Write the Electron Configuration for an Element in Each Block

Melissa Maribel · 7:23

https://www.youtube.com/watch?v=iFN9agJVea4

After the worked examples, for block-by-block practice.

KEY TERMS

orbital
A region around the nucleus where an electron is likely to be found; each orbital holds at most two electrons.

electron configuration
The arrangement of an atom's electrons among its orbitals, such as $1s^2 2s^2 2p^4$.

Aufbau principle
Electrons fill the lowest-energy orbitals first.

Pauli exclusion principle
No more than two electrons can occupy one orbital, and they must have opposite spins.

Hund's rule
Electrons fill orbitals of equal energy one at a time before pairing.

valence electrons
The electrons in the outermost shell, which take part in bonding.

CORE IDEAS

The story. Bohr's orbits worked beautifully for hydrogen but failed for atoms with more electrons. In the 1920s Louis de Broglie proposed that electrons behave as waves, and Erwin Schrödinger wrote an equation for those waves in 1926. The

solutions, called orbitals, are not paths but regions where the electron is likely to be found. Wolfgang Pauli added the rule that no two electrons in an atom can be in exactly the same state, and Friedrich Hund found how electrons spread out among orbitals of equal energy.

How they figured it out. Orbitals come in shapes named s (a sphere), p (a dumbbell; three per shell) and d (five per shell, more complex). An s sublevel holds 2 electrons, p holds 6, d holds 10. Energy levels (shells) are numbered n = 1, 2, 3 and so on. The order of filling is 1s, 2s, 2p, 3s, 3p, 4s, 3d, 4p, 5s. Notice 4s fills before 3d because it is slightly lower in energy for those elements.

The chemistry. To write a configuration, place electrons in order of increasing energy: oxygen (8 electrons) is $1s^2 2s^2 2p^4$. Noble-gas shorthand puts the core in brackets: sodium is $[\mathrm{Ne}]\, 3s^1$. The periodic table is a map of configurations: the s-block (groups 1–2), p-block (groups 13–18), d-block (transition metals). Elements in the same group have the same number of valence electrons, which is why they act alike: lithium, sodium and potassium each have one valence electron and react violently with water.

Where it is used. Configuration explains colors of gemstones (d electrons in chromium make ruby red), magnetism of iron, and why semiconductors such as silicon conduct a little. Chemists use valence electrons to predict bonding in every later unit.

LAB

Build the Periodic Table by Blocks

You will color a blank periodic table by s, p and d blocks and use it to write configurations without memorizing the filling order.

You will need

- A blank periodic table (print one), four colored pencils
- A list of 12 elements chosen by a parent

Steps

1. Color groups 1–2 (and helium) as the s-block, groups 13–18 as the p-block, and groups 3–12 as the d-block.
2. Label each row with its shell number. Note that the d-block in row 4 is the 3d sublevel.
3. For each of the 12 elements, read the configuration by walking across the table from hydrogen.
4. Check three of them by counting electrons.

Check your work

☐ Does each configuration add up to the atomic number?

☐ Do elements in the same column have the same ending (such as ns^1)?

LAB NOTEBOOK

1. Your colored block table.
2. Configurations of the 12 elements with valence electrons circled.

> **From history.** Pauli was famous for his sharp criticism: of one muddled paper he said, “It is not even wrong.” He predicted the neutrino in 1930 to save conservation of energy in radioactive decay; it was detected in 1956.

Another video. Orbital - Chem Definition. A short extra on the same idea, after the core reading.



For parents: Periodic Videos, 2:44. A short definition of an atomic orbital from the University of Nottingham chemists. Periodic Videos is an approved channel. Periodic Videos · 2:44

SEE IT: INTERACTIVE.

Quantum Explorer: Orbitals

See s, p and d orbitals as clouds of probability computed from Schrödinger's equation.

Pick 2p or 3d and count the nodes, where the electron is never found. Then open the Size and Pauli tab and step through the first 18 elements as the boxes fill two at a time with opposite spins, spreading out across a set before they pair (Hund's rule). Opens in a new tab. Worked answers for the parent: answer key (PDF).

WORKED EXAMPLES

Example 1. Write the full configuration and noble-gas shorthand for chlorine (Z = 17), and give its valence electrons.

▼ **Show solution**

1. Fill in order: $1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^5$ (2 + 2 + 6 + 2 + 5 = 17).
2. Shorthand: $[\text{Ne}]\ 3s^2\ 3p^5$
3. Valence electrons are in n = 3: 2 + 5 = 7.

Answer: $1s^2 2s^2 2p^6 3s^2 3p^5$; 7 valence electrons

Example 2. Draw the orbital diagram for nitrogen's 2p electrons and explain with Hund's rule.

▼ **Show solution**

1. Nitrogen: $1s^2\ 2s^2\ 2p^3$.
2. The three 2p electrons go one into each of the three 2p orbitals, all with the same spin: ↑ ↑ ↑.
3. Hund's rule: electrons spread out before pairing, which lowers repulsion.

Answer: One electron in each 2p orbital, spins parallel.

PRACTICE

1. Easy How many electrons can one orbital hold?

   ▼ **Show answer**

   Two.

2. Easy How many electrons fill an s, a p and a d sublevel?

   ▼ **Show answer**

   2, 6 and 10.

3. Easy How many valence electrons does nitrogen (Z = 7) have?

   ▼ **Show answer**

   5

4. **Medium** Write the full electron configuration of lithium (Z = 3).

   ▼ **Show answer**

   $1s^2 2s^1$

5. **Medium** Write the full electron configuration of sodium (Z = 11).

   ▼ **Show answer**

   $1s^2 2s^2 2p^6 3s^1$

COMMON MISTAKES

- Filling 3d before 4s.
- Putting more than two electrons in one orbital or more than six in a p sublevel.
- Pairing electrons in p orbitals before each has one (breaking Hund's rule).

> **Make it a game · Electron Seating Chart.** Draw boxes for 1s, 2s, 2p, 3s, 3p on a sheet. The parent names an element up to argon, and the student places coins (electrons) by the rules. A wrong placement loses the round.

> **Scripture.** "Let all things be done decently and in order." (1 Corinthians 14:40, KJV)
>
> Electrons fill orbitals in a definite order. Order in the smallest things reflects the God whose word asks that all things be done in order.

UNIT 3 · ELECTRONS AND THE PERIODIC TABLE · LESSON 3.3 OF 36 · 4 SESSIONS OF 45 MINUTES

3.3 Mendeleev's Table and the Missing Elements (1869; Gallium 1875)

OBJECTIVES

- Describe how Mendeleev organized the elements and why he left gaps.
- Explain how the discoveries of gallium, scandium and germanium confirmed his table.
- Use the periodic table: groups, periods, metals, nonmetals, metalloids and named families.

> **Warm-up.** Shuffle a deck of cards with a few cards removed. Lay them out by suit and value. Can you name the missing cards? How?

WATCH

Mendeleev's Business Card - Periodic Table of Videos

Periodic Videos · 4:05

https://www.youtube.com/watch?v=GuCR_Webyi8

> **WHEN TO WATCH**
> After the story.

BACKUP VIDEO

Gallium (beating heart) - Periodic Table of Videos

Periodic Videos · 7:12

https://www.youtube.com/watch?v=N6ccRvKKwZQ

After the worked example on gallium.

KEY TERMS

periodic law

When the elements are arranged by atomic number, their properties repeat in a regular pattern.

group

A vertical column of the periodic table; elements in a group have similar properties.

period

A horizontal row of the periodic table.

metal

An element that is shiny, conducts heat and electricity, is malleable and tends to lose electrons.

nonmetal

An element that is usually dull, brittle as a solid, a poor conductor, and tends to gain or share electrons.

metalloid

An element with properties between metals and nonmetals, such as silicon and germanium.

CORE IDEAS

The story. By the 1860s about 63 elements were known, and chemists noticed families: lithium, sodium and potassium; chlorine, bromine and iodine. Dmitri Mendeleev, a Russian professor writing a textbook, made a card for each element with its atomic mass and properties and arranged and rearranged them. In 1869 he published a table in which properties repeated in columns. His boldness was in what he left out: where no known element fit, he left a gap and predicted that an element would be found to fill it.

How they figured it out. Mendeleev predicted "eka-aluminum" with atomic mass about 68, a density about 5.9 g/cm^3, and a low melting point. In 1875 the French chemist Paul-Émile Lecoq de Boisbaudran found a new element by its spectrum and named it gallium. He first measured its density as 4.7; Mendeleev wrote that he should purify it and measure again, and the new value was 5.9. Scandium (1879) and germanium (1886) also matched predictions closely. Mendeleev even swapped tellurium and iodine against their masses because their properties demanded it; Moseley's atomic numbers (1913) proved him right.

The chemistry. The modern table is ordered by atomic number. Columns (groups 1–18) share valence-electron patterns and properties: group 1 alkali metals, group 2 alkaline earth metals, group 17 halogens, group 18 noble gases. Rows are periods. Metals lie to the left of the staircase line, nonmetals to the right, and metalloids along it. The lanthanides and actinides are pulled out below to keep the table narrow.

Where it is used. Chemists use the table to predict: a new drug might replace one atom with its group neighbor; semiconductor engineers swap silicon for germanium or gallium arsenide; geologists hunting for an element look where its group neighbors are found.

LAB

Card-Sort Periodic Table

Like Mendeleev, you will arrange cards of "mystery elements" by mass and properties, find the gap, and predict the missing card.

You will need

- 20 index cards prepared from a list (mass, color, state, formula of oxide and chloride for the first 20 elements, with names hidden and one card removed)
- Table space

Steps

1. Lay out the cards in order of increasing mass.
2. Start a new row whenever the properties begin to repeat (for example a very reactive metal follows a gas that does not react).
3. Find the gap and predict the missing element's mass, state and oxide formula.
4. Reveal the names and compare with the real table.

Check your work

- ☐ Where did properties repeat?
- ☐ How close was your prediction?
- ☐ Why are hydrogen and helium hard to place?

LAB NOTEBOOK

1. A photo or sketch of your card arrangement.
2. Your prediction for the missing card beside the real values.

> **From history.** Mendeleev said the arrangement came to him in a dream, but his notes show months of work sorting cards. Element 101, mendelevium, is named for him. The German chemist Lothar Meyer published a similar table in 1870, but he did not make bold predictions.

> **Try the lab.** WeduLab PeriodicLab: a free 3D periodic table. Press play on the history of discovery and watch the elements light up in the year they were found; Mendeleev knew only 63 when he left his gaps in 1869. Free, with no account. Opens in a new tab.

WORKED EXAMPLES

Example 1. Mendeleev predicted eka-silicon to have atomic mass 72, density 5.5 g/cm^3 and to form an oxide EsO_2. Germanium was found to have mass 72.6, density 5.35 g/cm^3 and oxide GeO_2. Find the percent difference for density.

▼ **Show solution**

1. $|5.5 - 5.35| / 5.35 \times 100 = 2.8\%$

Answer: About 3%: a remarkably close prediction.

Example 2. Use the table to classify and describe element 19.

▼ **Show solution**

1. Element 19 is potassium, in group 1, period 4.
2. It is an alkali metal: soft, shiny, very reactive with water, one valence electron.
3. It forms K^+ ions in compounds such as KCl.

Answer: Potassium, an alkali metal

PRACTICE

1. Easy How did Mendeleev arrange the elements?

   ▼ **Show answer**

   By increasing atomic mass, in rows and columns so that elements with similar properties lined up.

2. Medium Why did Mendeleev leave gaps in his table?

   ▼ **Show answer**

   No known element fit the pattern there; he predicted undiscovered elements would fill them.

3. **Medium** What happened when gallium's density was first measured?

▼ **Show answer**

It came out 4.7 g/cm^3; Mendeleev said it should be about 5.9; after purification it was 5.9.

COMMON MISTAKES

- Saying Mendeleev ordered by atomic number; he used atomic mass, adjusted by properties.
- Confusing groups (columns) and periods (rows).
- Thinking all metalloids are metals.

> **Make it a game · Element Detective.** The parent describes an element by its place and family ("period 3, a halogen"). The student names it. Then switch: the student describes, the parent guesses.

> **Scripture.** "Lift up your eyes on high, and behold who hath created these things, that bringeth out their host by number: he calleth them all by names by the greatness of his might, for that he is strong in power; not one faileth." (Isaiah 40:26, KJV)
>
> God calls the stars by name and none is missing. Mendeleev's gaps were filled because the elements follow an order not invented by men.

UNIT 3 · ELECTRONS AND THE PERIODIC TABLE · LESSON 3.4 OF 36 · 4 SESSIONS OF 45 MINUTES

3.4 Periodic Trends: Size, Ionization Energy and Pauling's Electronegativity

OBJECTIVES

- Describe and explain trends in atomic radius across a period and down a group.
- Describe and explain trends in ionization energy.
- Use Pauling's electronegativity values to compare atoms.
- Explain trends with nuclear charge and shielding.

> **Warm-up.** Two magnets attract a paper clip. One is near the clip; the other is far away behind a book. Which pulls harder, and why? Keep this picture in mind for electrons and the nucleus.

WATCH

Ionization Energy and Atomic Radius

Tyler DeWitt · 9:33

https://www.youtube.com/watch?v=Mmti4kKDcqA

> **WHEN TO WATCH**
> After the core ideas on radius and ionization energy.

BACKUP VIDEO

Periodic Trends Practice Problems: Atomic Radius

Melissa Maribel · 6:15

https://www.youtube.com/watch?v=pZH6M2oq5rA

With the practice set.

KEY TERMS

atomic radius
Half the distance between the nuclei of two bonded atoms of the same element; a measure of atom size.

ionization energy
The energy needed to remove the outermost electron from a gaseous atom.

electronegativity
The ability of an atom in a bond to attract shared electrons (Pauling scale, fluorine = 4.0).

effective nuclear charge
The net positive pull felt by an outer electron after inner electrons shield part of the nuclear charge.

shielding
The blocking of nuclear attraction on outer electrons by inner electrons.

cation
A positively charged ion formed when an atom loses electrons.

CORE IDEAS

The story. Once electron configurations were understood, chemists could explain why properties repeat. In 1932 Linus Pauling, studying the energies of chemical bonds, assigned each element a number for how strongly it pulls on shared electrons: electronegativity. Fluorine got the highest value, 4.0; cesium and francium the lowest, about 0.7. Pauling won the Nobel Prize in Chemistry in 1954 for his work on the chemical bond.

How they figured it out. Two competing effects control the trends. Across a period, protons are added to the nucleus while electrons go into the same shell, so the outer electrons feel a stronger pull and are drawn closer. Down a group, each new shell is farther out and inner electrons shield the outer ones, so the pull weakens. Measurements of bond lengths and the energy needed to pull electrons off atoms confirmed the predictions.

The chemistry. *Atomic radius* decreases across a period (left to right) and increases down a group. *Ionization energy* increases across a period and decreases down a group, so metals on the lower left lose electrons easily and are the most reactive metals. *Electronegativity* follows ionization energy: it rises toward the upper right (fluorine) and falls toward the lower left (cesium). Noble gases have very high ionization energies and form almost no compounds. Successive ionization energies jump sharply when an electron must be taken from a full inner shell, which shows the shell structure.

Where it is used. Trends predict reactivity: cesium is so reactive it is stored under oil; fluorine attacks almost everything. Electronegativity differences predict whether a bond is ionic or covalent (Unit 4) and which end of a molecule is negative, which governs how drugs fit their targets.

LAB

Graphing Periodic Trends

You will graph atomic radius and first ionization energy against atomic number for elements 1–36 and find the repeating pattern.

You will need

- A data table of atomic radius and ionization energy for elements 1–36 (from a reference or the periodic table app)
- Spreadsheet or graph paper

Steps

1. Enter atomic number, radius and first ionization energy for elements 1–36.
2. Make two line graphs against atomic number.
3. Mark the alkali metals and noble gases on each graph.
4. Describe the repeating pattern and explain one peak and one valley.

Check your work

- ☐ Where are the peaks of ionization energy?
- ☐ Where are the peaks of atomic radius?
- ☐ Why do the two graphs look like mirror images?

LAB NOTEBOOK

1. Both graphs with the alkali metals and noble gases labeled.
2. A paragraph explaining the trends with nuclear charge and shielding.

> **From history.** Pauling is the only person to have won two unshared Nobel Prizes (Chemistry 1954, Peace 1962). His book *The Nature of the Chemical Bond* (1939) is among the most cited science books ever written.

SEE IT: INTERACTIVE.

Quantum Explorer: Size and Pauli

Step through the first 18 elements and watch the orbital boxes fill.

Each box holds two electrons with opposite spins (the Pauli exclusion principle), and each new row of the periodic table begins when electrons start a new shell, the root of the trends in size and ionization energy in this lesson. Opens in a new tab on the Size and Pauli tab. Worked answers for the parent: answer key (PDF).

WORKED EXAMPLES

Example 1. Rank Na, Mg and Cl by atomic radius and explain.

▼ **Show solution**

1. All are in period 3, with outer electrons in n = 3.
2. Nuclear charge rises from Na (11) to Mg (12) to Cl (17), pulling the electrons closer.
3. Largest to smallest: Na > Mg > Cl.

Answer: Na > Mg > Cl

Example 2. Magnesium's first three ionization energies are 738, 1451 and 7733 kJ/mol. Explain the jump.

▼ **Show solution**

1. Mg is $[\mathrm{Ne}]\,3s^2$.
2. The first two electrons come from the 3s shell.
3. The third must come from the full, inner 2p shell, much closer to the nucleus, so far more energy is needed.

Answer: Two valence electrons, then an inner shell

PRACTICE

1. Easy Which is larger: a potassium atom or a sodium atom?

▼ **Show answer**

Potassium.

2. Easy Which element has the highest electronegativity?

   ▼ **Show answer**

   Fluorine (4.0).

3. Medium Why does atomic radius decrease across a period?

   ▼ **Show answer**

   Protons are added while electrons enter the same shell, so a stronger nuclear pull draws the electrons closer.

COMMON MISTAKES

- Thinking atoms get bigger across a period because they have more electrons.
- Mixing up trends for ionization energy and atomic radius.
- Assigning electronegativity to noble gases, which rarely bond.

> **Make it a game · Trend Arrows.** The parent names two elements. The student points an arrow (with a hand or a pencil) toward the larger atom, then toward the higher ionization energy, then toward the higher electronegativity. Three points per round.

> **Scripture.** “Where wast thou when I laid the foundations of the earth? declare, if thou hast understanding.” (Job 38:4, KJV)
>
> God asks Job where he was when the earth’s foundations were laid. The patterns of the elements were set long before chemists graphed them.

UNIT 4

Chemical Bonds and Molecular Shape

Atoms rarely stand alone. This unit asks how they hold together and why the way they hold together matters. Students meet ionic crystals such as table salt, the shared electron pairs that G. N. Lewis drew as dots in 1916, the shapes that VSEPR predicts, and the weaker attractions between molecules that make water boil at 100 °C instead of −80 °C. By the end, a student can look at a formula and predict whether the substance is a brittle, high-melting crystal, a gas, or a liquid that climbs a paper towel.

PREREQUISITES

- Unit 3: valence electrons, electronegativity and periodic trends

MASTERY CHECKLIST

- ☐ I can explain how ionic bonds form and predict ionic formulas from charges.
- ☐ I can draw Lewis structures for simple molecules, including double and triple bonds.
- ☐ I can predict molecular shape and polarity with VSEPR.
- ☐ I can rank substances by intermolecular forces and explain boiling points, surface tension and solubility.

LESSONS IN THIS UNIT

4.1	**Ionic Bonds and Crystals: Salt of the Earth**	4 sessions of 45 minutes
4.2	**Covalent Bonds and G. N. Lewis's Dot Structures (1916)**	4 sessions of 45 minutes
4.3	**Molecular Shape: VSEPR and Polarity**	4 sessions of 45 minutes
4.4	**Intermolecular Forces: Why Water Is So Strange**	4 sessions of 45 minutes

Unit 4 test

Show your work. For every calculation write the equation, substitute with units and give the answer with its unit and sensible significant figures. The test is worth 100 points.

1. **10 pts** **Lesson 4.1**
   Write the formula of copper(II) hydroxide and explain the parentheses.

2. **10 pts** **Lesson 4.1**
   What electronegativity difference generally marks an ionic bond?

3. **10 pts** **Lesson 4.2**

   Electronegativities: Na 0.9, Cl 3.0, C 2.5, H 2.1. Classify Na–Cl and C–H bonds.

4. **9 pts** **Lesson 4.2**

   Which atom is never placed in the center of a Lewis structure?

5. **9 pts** **Lesson 4.3**

   Who proposed the tetrahedral carbon atom, and what did it explain?

6. **9 pts** **Lesson 4.3**

   Is CH_2Cl_2 polar? Explain.

7. **9 pts** **Lesson 4.4**

   Who proposed the hydrogen bond, and where?

8. **9 pts** **Lesson 4.4**

   Which has stronger intermolecular forces: water or vegetable oil? Use the penny test result.

LAB AND DATA

9. **10 pts**

   You test four white solids: one melts easily in a spoon over a candle and does not conduct when dissolved; one does not melt and its solution conducts; one does not melt and does not dissolve; one dissolves but its solution does not conduct. Which is likely table salt, sugar, paraffin wax and sand? Explain.

SHORT ESSAY

10. **15 pts**

    Water is strange: it is a liquid at room temperature though its molecules are lighter than oxygen gas, and its solid floats. Explain both facts using molecular shape, polarity and hydrogen bonding, and give one reason these properties matter for life on earth.

ANSWER KEY

1. $Cu(OH)_2$; two hydroxide ions are needed, and the parentheses show the subscript applies to the whole ion.
2. Greater than about 1.7.
3. Na–Cl: difference 2.1, ionic. C–H: 0.4, essentially nonpolar covalent.
4. Hydrogen.
5. Van ’t Hoff and Le Bel (1874); it explained mirror-image forms of some molecules.
6. Yes; the tetrahedral shape has unlike atoms around carbon, so the C–Cl dipoles do not cancel.
7. Latimer and Rodebush in 1920, in G. N. Lewis’s Berkeley laboratory.
8. Water: it held more drops because of its higher surface tension.
9. Melts easily, no conduction: paraffin wax (nonpolar covalent, weak forces). High melting, solution conducts: table salt (ionic). Does not melt or dissolve: sand (network covalent SiO_2). Dissolves but does not conduct: sugar (polar covalent molecules, no ions).
10. Look for: bent shape (104.5°) and polar O–H bonds make water polar; hydrogen bonds between molecules require much energy to separate, so a high boiling point; in ice hydrogen bonds lock molecules in an open hexagonal lattice, less dense than liquid, so ice floats; lakes freeze from the top, protecting fish; high specific heat moderates climate; water dissolves many substances needed by cells.

Printable answer key (PDF, for the parent)

UNIT 4 · CHEMICAL BONDS AND MOLECULAR SHAPE · LESSON 4.1 OF 36 · 4 SESSIONS OF 45 MINUTES

4.1 Ionic Bonds and Crystals: Salt of the Earth

OBJECTIVES

- Explain how atoms form ions to reach a noble-gas configuration.
- Predict ionic charges from group numbers and write formulas of ionic compounds.
- Name binary ionic compounds and common polyatomic ions.

- Relate the crystal lattice to the properties of ionic solids.

> **Warm-up.** Look at grains of table salt under a magnifying glass. Describe their shape. Why might every grain be a little cube?

WATCH

Ionic Bonding Introduction

Tyler DeWitt · 7:20

https://www.youtube.com/watch?v=Qf07-8Jhhpc

> **WHEN TO WATCH**
> After the core ideas.

BACKUP VIDEO

Sodium Explodes in Water...Why Doesn't Salt?!?

Tyler DeWitt · 7:45

https://www.youtube.com/watch?v=CArZeDTwVh4

With the naming practice.

KEY TERMS

ion

An atom or group of atoms with a net electric charge.

crystal lattice

A repeating three-dimensional arrangement of particles in a solid.

polyatomic ion

A charged group of covalently bonded atoms, such as SO_4^{2-}.

ionic bond

The electrostatic attraction between oppositely charged ions.

octet rule

Atoms tend to gain, lose or share electrons to have eight valence electrons.

formula unit

The simplest ratio of ions in an ionic compound, such as NaCl.

CORE IDEAS

The story. Salt was so valuable in the ancient world that Roman soldiers were sometimes paid an allowance to buy it, a word that survives in “salary.” Humphry Davy’s electrolysis (Unit 5) showed salts are made of a metal and a nonmetal held by electrical forces. In 1913 William Henry Bragg and his son Lawrence fired X-rays through a salt crystal and found no NaCl molecules at all: just a checkerboard of sodium and chlorine particles, each surrounded by six of the other.

How they figured it out. Sodium ($[\mathrm{Ne}]\,3s^1$) has a low ionization energy and loses its one valence electron to become $\mathrm{Na^+}$ with neon’s configuration. Chlorine ($[\mathrm{Ne}]\,3s^2 3p^5$) has a high electronegativity and gains one electron to become $\mathrm{Cl^-}$ with argon’s configuration. The oppositely charged ions attract in all directions, building a lattice. Generally, a bond is ionic when the electronegativity difference is greater than about 1.7.

The chemistry. Group 1 metals form 1+ ions, group 2 form 2+, aluminum 3+; group 15 nonmetals form 3−, group 16 form 2−, group 17 form 1−. The formula balances the charges: $\mathrm{Mg^{2+}}$ and $\mathrm{Cl^-}$ give $\mathrm{MgCl_2}$; $\mathrm{Al^{3+}}$ and $\mathrm{O^{2-}}$ give $\mathrm{Al_2O_3}$ (crisscross the charges and reduce). Names give the metal, then the nonmetal ending in *-ide*: sodium chloride, magnesium oxide. Transition metals with more than one charge get a Roman numeral: iron(III) oxide. Common polyatomic ions: nitrate $\mathrm{NO_3^-}$, sulfate $\mathrm{SO_4^{2-}}$, carbonate $\mathrm{CO_3^{2-}}$, hydroxide $\mathrm{OH^-}$, ammonium $\mathrm{NH_4^+}$, phosphate $\mathrm{PO_4^{3-}}$.

Where it is used. Ionic solids have high melting points (NaCl melts at 801 °C), are hard but brittle (a shift lines up like charges, which repel, and the crystal cracks), and conduct electricity when melted or dissolved because the ions can move. Electrolytes in sports drinks, road salt, baking soda, chalk and the calcium phosphate of bone are all ionic.

> **Safety first.** Adult handles hot water. Alum is safe in kitchen quantities but do not eat the crystals.

LAB

Growing Salt and Alum Crystals

You will grow crystals from saturated solutions and compare their shapes.

You will need

- Table salt, alum (spice aisle), distilled water
- Two clean jars, a saucepan, string, a pencil, a magnifier

Steps

1. With an adult, heat 1 cup of water and stir in salt until no more dissolves. Pour into a jar, leaving any undissolved salt behind.
2. Do the same with alum in the second jar.
3. Hang a string from a pencil into each jar. Cover loosely and leave undisturbed for a week.
4. Observe daily with a magnifier and sketch the crystal shapes.

Check your work

- ☐ What shape do the salt crystals have? The alum crystals?
- ☐ How does the shape reflect the lattice?

LAB NOTEBOOK

1. Daily crystal sketches with dates.
2. An explanation of why salt forms cubes.

> **From history.** Lawrence Bragg was 25 when he shared the 1915 Nobel Prize in Physics with his father, still the youngest science laureate. X-ray crystallography later revealed the structures of penicillin, vitamin B_{12} and DNA.

WORKED EXAMPLES

Example 1. Write the formula and name for the compound of calcium and nitrogen.

▼ **Show solution**

1. Calcium, group 2: Ca^{2+}. Nitrogen, group 15: N^{3-}.
2. Least common multiple of 2 and 3 is 6: three Ca^{2+} (+6) and two N^{3-} (−6).
3. Formula Ca_3N_2; name calcium nitride.

Answer: Ca_3N_2, calcium nitride

Example 2. Write the formula of aluminum sulfate.

▼ **Show solution**

1. Al^{3+} and SO_4^{2-}.
2. Two Al (+6) and three sulfate (−6).
3. Use parentheses for more than one polyatomic ion.

Answer: $Al_2(SO_4)_3$

PRACTICE

1. Easy What charge does a magnesium ion have?

   ▼ **Show answer**

   2+

2. Easy Write the formula of potassium bromide.

   ▼ **Show answer**

   KBr

3. Medium Write the formula of aluminum oxide.

   ▼ **Show answer**

   Al_2O_3

COMMON MISTAKES

- Writing ionic charges in the formula (Na^+Cl^- instead of NaCl).

- Forgetting parentheses around repeated polyatomic ions.
- Calling an ionic compound a molecule.

> **Make it a game · Ion Crisscross.** Write ion cards (Na^+, Ca^{2+}, Al^{3+}, Cl^-, O^{2-}, N^{3-}, SO_4^{2-}). Draw one positive and one negative; the first to write the correct formula and name scores.

> **Scripture.** "Ye are the salt of the earth: but if the salt have lost his savour, wherewith shall it be salted? it is thenceforth good for nothing, but to be cast out, and to be trodden under foot of men." (Matthew 5:13, KJV)
>
> Salt preserved and flavored food. Jesus used it to describe what His people are to be in the world.

UNIT 4 · CHEMICAL BONDS AND MOLECULAR SHAPE · LESSON 4.2 OF 36 · 4 SESSIONS OF 45 MINUTES

4.2 Covalent Bonds and G. N. Lewis's Dot Structures (1916)

OBJECTIVES

- Explain covalent bonding as the sharing of electron pairs.
- Draw Lewis structures for molecules with single, double and triple bonds.
- Distinguish polar and nonpolar covalent bonds using electronegativity.

> **Warm-up.** Two students each have one textbook, but each needs two to finish an assignment. How could they both get access to two books without buying more?

WATCH

How To Draw Lewis Structures

The Organic Chemistry Tutor · 11:50

https://www.youtube.com/watch?v=NFZtjSeT3XE

> **WHEN TO WATCH**
> After the story.

BACKUP VIDEO

GCSE Chemistry - Covalent Bonding - Formation | Drawing Covalent Bonds

Cognito · 5:57

https://www.youtube.com/watch?v=7IkYm7ZgiAw

After drawing the first structures.

KEY TERMS

covalent bond
A bond in which two atoms share a pair of electrons.

lone pair
A pair of valence electrons not shared in a bond.

polar covalent bond
A bond with unequal sharing, giving partial charges δ+ and δ−.

Lewis structure
A diagram showing valence electrons as dots and shared pairs as lines.

double bond
Two shared pairs of electrons between two atoms.

molecule
A neutral group of atoms held together by covalent bonds.

CORE IDEAS

The story. In 1902 Gilbert Newton Lewis, teaching at Harvard, sketched atoms as cubes with electrons at the corners to explain chemistry to his students. In 1916, at Berkeley, he published the idea that a chemical bond is a *shared pair of electrons*, and drew the dots that every chemistry student has drawn since. Irving Langmuir

popularized the octet rule. Lewis was nominated for the Nobel Prize 41 times and never won.

How they figured it out. Ionic bonding could not explain H_2, Cl_2 or methane: two identical atoms cannot transfer an electron to each other. Lewis proposed they share. Each hydrogen in H_2 "counts" both shared electrons, reaching helium's stable pair. Each chlorine in Cl_2 shares one pair and keeps three lone pairs, completing an octet.

The chemistry. To draw a Lewis structure: (1) count all valence electrons; (2) place the least electronegative atom in the center (never hydrogen); (3) connect atoms with single bonds; (4) fill octets on outer atoms with lone pairs; (5) place leftover electrons on the central atom; (6) if the center lacks an octet, turn lone pairs into double or triple bonds. O_2 has a double bond and N_2 a triple bond, which is why nitrogen gas is so unreactive. When electronegativities differ by about 0.4 to 1.7, the bond is polar: in H–Cl the chlorine end is $\delta-$.

Where it is used. Covalent molecules make up most of the substances in living things, plastics and fuels. The strong N≡N triple bond is why fertilizer had to be made by the Haber process (Unit 6). Diamond and quartz are covalent networks, among the hardest substances known.

LAB

Gumdrop and Toothpick Molecules

You will build models from Lewis structures; next lesson you will use them for shapes.

You will need

- Gumdrops or modeling clay in four colors, toothpicks
- A list of molecules: H_2, Cl_2, H_2O, NH_3, CH_4, CO_2, N_2, HCN

Steps

1. Draw the Lewis structure of each molecule.
2. Build a model: one toothpick per shared pair, two or three bent toothpicks for double or triple bonds.
3. Mark lone pairs with small paper flags.
4. Photograph or sketch each model.

Check your work

☐ Did every atom (except H) reach an octet?

☐ Which molecules have multiple bonds?

LAB NOTEBOOK

1. Lewis structures for the eight molecules.
2. Photos or sketches of the models.

> **From history.** Lewis died in his Berkeley laboratory in 1946 while working with hydrogen cyanide. He trained or hired many Nobel laureates, including Glenn Seaborg and Harold Urey.

WORKED EXAMPLES

Example 1. Draw the Lewis structure of carbon dioxide.

▼ **Show solution**

1. Valence electrons: C 4 + O 6 × 2 = 16.
2. Carbon in the center: O–C–O uses 4; outer octets use 12; 0 left; carbon has only 4.
3. Move one lone pair from each O into a double bond: O=C=O. Each atom has an octet, each O has two lone pairs.

Answer: O=C=O with two lone pairs on each oxygen

Example 2. Is the C–O bond polar? Electronegativities: C 2.5, O 3.5.

▼ **Show solution**

1. Difference 1.0, between 0.4 and 1.7.
2. Polar covalent, with O δ−.

Answer: Polar covalent

PRACTICE

1. Easy What is a covalent bond?

   ▼ **Show answer**

   A shared pair of electrons between two atoms.

2. Easy How many valence electrons are in a water molecule?

   ▼ **Show answer**

   8

3. Medium Draw (describe) the Lewis structure of ammonia.

   ▼ **Show answer**

   N in the center with three single bonds to H and one lone pair.

COMMON MISTAKES

- Miscounting valence electrons, especially for ions (add for negative charge, subtract for positive).
- Putting hydrogen in the center or giving it more than one bond.
- Leaving the central atom short of an octet instead of forming a multiple bond.

REFERENCE GUIDE

An eight-page, color-coded Organic Chemistry guide: carbon bonding and hybridization, structural formulas, functional groups, naming, isomers and chirality, acids and bases with pKa values, reaction types and mechanisms, spectroscopy (IR and NMR), and the molecules of life. The booklet prints on 11×17 paper and folds to letter size. Prints in color and reads clearly in black and white.

> **Make it a game · Dot Count Race.** The parent writes a formula; the student counts the total valence electrons in under 10 seconds. Then draw the structure for a bonus point.

> **Scripture.** “Two are better than one; because they have a good reward for their labour.” (Ecclesiastes 4:9, KJV)
>
> Two atoms sharing electrons are more stable than either alone, a small picture of the Preacher’s word.

UNIT 4 · CHEMICAL BONDS AND MOLECULAR SHAPE · LESSON 4.3 OF 36 · 4 SESSIONS OF 45 MINUTES

4.3 Molecular Shape: VSEPR and Polarity

OBJECTIVES

- Use VSEPR theory to predict molecular shapes and bond angles.
- Decide whether a molecule is polar from its bonds and shape.
- Connect shape and polarity to properties.

> **Warm-up.** Tie two, three and four balloons together at their necks. What arrangements do they take on their own?

WATCH

VSEPR Theory - Basic Introduction

The Organic Chemistry Tutor · 13:10

https://www.youtube.com/watch?v=DBrq31w8vC4

> **WHEN TO WATCH**
> After the balloon demonstration.

BACKUP VIDEO

VSEPR Theory: Common Mistakes

Tyler DeWitt · 9:32

https://www.youtube.com/watch?v=8Tl_bDWCAmo

With the shape table.

KEY TERMS

VSEPR

Valence Shell Electron Pair Repulsion: electron groups around a central atom spread as far apart as possible.

linear

A shape with atoms in a straight line, 180°.

trigonal planar

Three groups in a flat triangle, 120°.

tetrahedral

Four groups pointing to the corners of a tetrahedron, 109.5°.

bent

A V-shaped molecule, such as water (104.5°).

polar molecule

A molecule with a positive and a negative end because its bond dipoles do not cancel.

CORE IDEAS

The story. In 1874 Jacobus van ’t Hoff and Joseph Le Bel independently proposed that carbon’s four bonds point to the corners of a tetrahedron, explaining why some molecules exist in mirror-image forms. The general rule for predicting shape,

VSEPR, was worked out by Ronald Gillespie and Ronald Nyholm in London in 1957, and Gillespie spent his career at McMaster University teaching it.

How they figured it out. Electron groups (bonds or lone pairs) are negative and repel one another, so they spread as far apart as possible. Two groups go to opposite sides (180°), three to the corners of a triangle (120°), four to a tetrahedron (109.5°). A double or triple bond counts as one group. Lone pairs occupy space but are invisible in the shape, and they push a little harder, squeezing bond angles.

The chemistry. Count groups on the central atom: CO_2 has 2, linear. BF_3 has 3, trigonal planar. CH_4 has 4, tetrahedral. NH_3 has 4 groups with 1 lone pair: trigonal pyramidal, 107°. H_2O has 4 groups with 2 lone pairs: bent, 104.5°. A molecule is polar if it has polar bonds *and* a shape that does not cancel them. CO_2 has polar bonds but is linear and symmetric, so it is nonpolar. Water is bent, so it is polar.

Where it is used. Shape determines how molecules fit together: enzymes recognize their targets by shape (Biology Unit 2), drugs are designed to fit receptors, and the bent shape of water is the root of its strange properties (Lesson 4.4). Polar substances dissolve polar substances: “like dissolves like.”

LAB

Balloon VSEPR and the Bending Stream

You will model electron groups with balloons and test liquids for polarity with a charged comb.

You will need

- Balloons (8), string
- A plastic comb or balloon rubbed on wool
- A faucet, a burette or a cup with a pinhole; water and vegetable oil (adult supervision)

Steps

1. Tie 2, 3 and 4 balloons together and observe the shapes. Measure the angles roughly with a protractor.
2. Use one smaller balloon to stand for a lone pair: how does the shape change?
3. Charge a comb by rubbing it on your hair or wool. Hold it near a thin stream of water. Observe.
4. Repeat with a thin stream of vegetable oil from a cup with a pinhole, over a tray.

Check your work

- ☐ Which stream bent toward the comb? Why?
- ☐ Which balloon shapes matched linear, trigonal planar and tetrahedral?

LAB NOTEBOOK

1. Photos or sketches of the balloon shapes with angles.
2. Observations of the water and oil streams with an explanation.

> **From history.** Van ’t Hoff’s tetrahedral carbon was ridiculed by the eminent chemist Hermann Kolbe as fantasy. Van ’t Hoff went on to win the first Nobel Prize in Chemistry in 1901.

SEE IT: INTERACTIVE.

Twelve Pentagons

See the C_{60} buckyball, 60 carbon atoms on the corners of a soccer-ball shape, and count its bonds and rings with Euler's rule.

The In the real world tab also shows virus shells built on the same 12-pentagon plan. Opens in a new tab. Worked answers for the parent: answer key (PDF).

WORKED EXAMPLES

Example 1. Predict the shape and bond angle of PCl_3.

▼ **Show solution**

1. P has 5 valence electrons; three bonds to Cl and one lone pair: 4 groups.
2. Electron geometry tetrahedral; with one lone pair the shape is trigonal pyramidal.
3. Angle a little less than 109.5° (measured: about 100°).

Answer: Trigonal pyramidal, about 100°

Example 2. Is CCl_4 polar?

▼ **Show solution**

1. C–Cl bonds are polar (difference 0.5).
2. Four identical bonds point to the corners of a tetrahedron and cancel.

Answer: Nonpolar

PRACTICE

1. Easy What shape is CH_4, and what is its bond angle?

   ▼ **Show answer**

   Tetrahedral, 109.5°.

2. Easy What shape is CO_2?

   ▼ **Show answer**

   Linear.

3. **Medium** Why is water bent rather than linear?

▼ **Show answer**

Oxygen has two bonds and two lone pairs; the four groups are arranged tetrahedrally, and the lone pairs make the shape bent.

COMMON MISTAKES

- Ignoring lone pairs when counting groups.
- Calling a molecule polar because it has polar bonds even when the shape cancels them.
- Counting a double bond as two groups.

REFERENCE GUIDE

An eight-page, color-coded Organic Chemistry guide: carbon bonding and hybridization, structural formulas, functional groups, naming, isomers and chirality, acids and bases with pKa values, reaction types and mechanisms, spectroscopy (IR and NMR), and the molecules of life. The booklet prints on 11×17 paper and folds to letter size. Prints in color and reads clearly in black and white.

> **Make it a game · Shape Charades.** The parent names a molecule; the student uses arms (and a friend's arms, or pencils) to show the shape and calls out the bond angle.

> **Scripture.** "I will praise thee; for I am fearfully and wonderfully made: marvellous are thy works; and that my soul knoweth right well." (Psalm 139:14, KJV)
>
> Even the shape of a water molecule is part of how we are fearfully and wonderfully made, since our bodies are mostly water.

4.4 Intermolecular Forces: Why Water Is So Strange

OBJECTIVES

- Name and compare London dispersion forces, dipole–dipole forces and hydrogen bonds.
- Use intermolecular forces to explain boiling points, surface tension, capillary action and solubility.
- Explain why ice floats.

> **Warm-up.** Gently lay a paper clip flat on the surface of a cup of water using a fork. Why does it float though steel is denser than water? Then add a drop of dish soap. What happens?

WATCH

Hydrogen Bonding and Common Mistakes

Tyler DeWitt · 9:00

https://www.youtube.com/watch?v=PyC5r2mB4d4

> **WHEN TO WATCH**
> After the warm-up.

BACKUP VIDEO

Intermolecular Forces and Boiling Point

Tyler DeWitt · 5:20

https://www.youtube.com/watch?v=4aSHovdnCyY

After the hydrogen-bond discussion.

KEY TERMS

intermolecular force

An attraction between molecules, much weaker than a chemical bond within a molecule.

London dispersion force

A weak attraction from momentary uneven electron distributions; present in all molecules, larger for bigger molecules.

dipole–dipole force

Attraction between the positive end of one polar molecule and the negative end of another.

hydrogen bond

A strong dipole attraction between H bonded to N, O or F and a lone pair on another N, O or F.

surface tension

The tendency of a liquid surface to resist stretching because molecules at the surface are pulled inward.

capillary action

The rise of a liquid in a narrow tube because of attraction to the tube's walls.

CORE IDEAS

The story. Water should be a gas. Molecules of similar size, such as methane (CH_4, boiling point −164 °C) and hydrogen sulfide (H_2S, −60 °C), are gases at room temperature. Yet water boils at 100 °C. In 1920 Wendell Latimer and Worth Rodebush, in G. N. Lewis's Berkeley laboratory, proposed that a hydrogen atom between two oxygen atoms acts as a bridge: the hydrogen bond. Pauling popularized the term.

How they figured it out. Johannes van der Waals showed in 1873 that real gases do not behave ideally because their molecules attract. Fritz London explained in 1930 that even nonpolar molecules attract slightly when their electron clouds momentarily shift. Comparing boiling points across families showed that water, ammonia and hydrogen fluoride boil far higher than the trend predicts: evidence for an extra, stronger attraction.

The chemistry. Weakest to strongest: *London dispersion* (all molecules; increases with size, which is why iodine is solid and fluorine is gas), *dipole–dipole* (polar molecules), *hydrogen bonding* (H on N, O or F). Stronger forces mean higher boiling points, higher surface tension and lower evaporation rates. In ice each water molecule hydrogen-bonds to four others in an open hexagonal lattice that is about 9% less dense than liquid water, so ice floats.

Where it is used. Because ice floats, lakes freeze from the top down and fish survive the winter. Capillary action lifts water in paper towels and in plant stems. Hydrogen bonds hold the two strands of DNA together (Biology Unit 6) and give proteins their

shapes. Water's high specific heat moderates coastal climates and keeps our body temperature steady.

> **Safety first.** Rubbing alcohol is flammable: no flames nearby. Do not drink any liquid.

LAB

Drops on a Penny and Evaporation Races

You will compare surface tension and evaporation for water, rubbing alcohol and oil.

You will need

- Three clean pennies, three droppers
- Water, isopropyl rubbing alcohol (70% or 91%), vegetable oil
- Paper towel strips, three cotton swabs, a timer

Steps

1. Count how many drops of each liquid fit on a penny before it spills. Repeat three times and average.
2. Swab a stripe of each liquid side by side on a dark countertop or chalkboard and time how long until each disappears.
3. Dip the end of paper towel strips into water and oil; measure the height reached after 5 minutes.

Check your work

- ☐ Which liquid held the most drops? Why?
- ☐ Which evaporated fastest? What does this say about its intermolecular forces?

LAB NOTEBOOK

1. Data table: drops on a penny (3 trials, average), evaporation times, capillary heights.
2. A conclusion ranking the liquids by intermolecular forces.

> **From history.** Water striders, small insects, walk on water by spreading their weight over long legs covered in water-repelling hairs. Biologists and engineers have built robot water striders to study how they jump.

WORKED EXAMPLES

Example 1. Rank CH_4, HCl and H_2O by boiling point and explain.

▼ **Show solution**

1. CH_4: nonpolar, London forces only.
2. HCl: polar, dipole–dipole.
3. H_2O: hydrogen bonding.

Answer: $CH_4 < HCl < H_2O$

Example 2. Why does iodine (I_2) exist as a solid while chlorine (Cl_2) is a gas, though both are nonpolar?

▼ **Show solution**

1. Both have only London forces.
2. Iodine is much larger, with more electrons that can shift, so its London forces are stronger.

Answer: Larger molecule, stronger London forces

PRACTICE

1. Easy Which intermolecular force is present in all molecules?

   ▼ **Show answer**

   London dispersion forces.

2. Easy Why does ice float on water?

   ▼ **Show answer**

   Hydrogen bonds hold ice in an open lattice that is less dense than liquid water.

3. **Medium** What is a hydrogen bond?

▼ **Show answer**

An attraction between H bonded to N, O or F and a lone pair on another N, O or F atom.

COMMON MISTAKES

- Confusing intermolecular forces with the bonds inside a molecule; boiling water does not break O–H bonds.
- Calling any bond with hydrogen a hydrogen bond.
- Forgetting that London forces act in all molecules.

> **Make it a game · Force Ladder.** Write substances on cards (methane, ammonia, water, iodine, neon, ethanol, HCl). The student lines them up from weakest to strongest intermolecular forces and names the main force for each.

> **Scripture.** "Out of whose womb came the ice? and the hoary frost of heaven, who hath gendered it? The waters are hid as with a stone, and the face of the deep is frozen." (Job 38:29–30, KJV)
>
> God asks Job about the ice and the hoary frost. Water hiding as stone above the deep is exactly what floating ice does.

UNIT 5

Chemical Reactions

A reaction is a rearrangement of atoms: nothing is created and nothing lost, as Lavoisier showed. This unit teaches students to write and balance equations, to sort reactions into five types, to predict precipitates with solubility rules, and to follow electrons in oxidation and reduction. Michael Faraday's Christmas lectures on a candle and Humphry Davy's electrolysis of potash, which gave the world potassium, anchor the history.

Labs use household chemicals: baking soda and vinegar, steel wool, copper wire and Epsom salt. One demonstration (Davy's electrolysis) is watch-only.

PREREQUISITES

- Unit 4: ionic formulas, polyatomic ions and naming

MASTERY CHECKLIST

- ☐ I can balance chemical equations by inspection.
- ☐ I can classify a reaction as synthesis, decomposition, single replacement, double replacement or combustion and predict products.
- ☐ I can use solubility rules to predict precipitates and write net ionic equations.
- ☐ I can assign oxidation numbers and identify what is oxidized and reduced.

LESSONS IN THIS UNIT

5.1 Writing and Balancing Equations — 4 sessions of 45 minutes

5.2 Five Reaction Types and Faraday’s Chemical History of a Candle (1848)

4 sessions of 45 minutes

5.3 Precipitates, Solubility Rules and Net Ionic Equations — 4 sessions of 45 minutes

5.4 Oxidation and Reduction: Rust, Batteries and Davy’s Electrolysis (1807)

4 sessions of 45 minutes

Unit 5 test

Show your work. For every calculation write the equation, substitute with units and give the answer with its unit and sensible significant figures. The test is worth 100 points.

1. **9 pts** **Lesson 5.1**

   Write a balanced equation: aluminum and oxygen form aluminum oxide.

2. **9 pts** **Lesson 5.1**

   Balance: $\mathrm{Ca(OH)_2 + H_3PO_4 \rightarrow Ca_3(PO_4)_2 + H_2O}$.

3. **9 pts** **Lesson 5.1**

   Balance: $\mathrm{C_3H_8 + O_2 \rightarrow CO_2 + H_2O}$

4. **8 pts** **Lesson 5.2**

   What reaction type inflates an airbag? Give the equation.

5. **8 pts** **Lesson 5.2**

   Complete: $\mathrm{Mg + O_2 \rightarrow}$ (balance and classify).

6. **8 pts** **Lesson 5.3**

   Is lead(II) iodide soluble? What color is it?

7. **8 pts** **Lesson 5.3**

   Why does hard water form soap scum?

8. **8 pts** **Lesson 5.4**

   What is the oxidation number of O in $\mathrm{H_2O_2}$?

9. **8 pts** **Lesson 5.4**

   Is rusting oxidation or reduction of iron?

LAB AND DATA

10. **10 pts**

    Copper wire is placed in silver nitrate solution. Over an hour the solution turns blue and silvery crystals grow on the wire. (a) Write the balanced equation. (b) Classify the reaction. (c) Which element is oxidized and which reduced?

SHORT ESSAY

11. **15 pts**

Faraday said there is no better door into natural philosophy than the physics and chemistry of a candle. Using at least three ideas from this unit, explain what happens when a candle burns.

ANSWER KEY

1. $4Al + 3O_2 \rightarrow 2Al_2O_3$
2. $3Ca(OH)_2 + 2H_3PO_4 \rightarrow Ca_3(PO_4)_2 + 6H_2O$
3. $C_3H_8 + 5O_2 \rightarrow 3CO_2 + 4H_2O$

   Balance one element at a time, leaving H and O (or a free element) for last; then count every atom on both sides.
4. Decomposition: $2NaN_3 \rightarrow 2Na + 3N_2$.
5. $2Mg + O_2 \rightarrow 2MgO$, synthesis.
6. No; bright yellow.
7. Calcium and magnesium ions form insoluble compounds with soap.
8. −1
9. Oxidation.
10. (a) $Cu + 2AgNO_3 \rightarrow Cu(NO_3)_2 + 2Ag$. (b) Single replacement (also redox). (c) Copper is oxidized (0 → +2); silver is reduced (+1 → 0).
11. Look for: wax (a hydrocarbon) melts, rises up the wick by capillary action, vaporizes; combustion with oxygen gives carbon dioxide and water (balanced equation such as $C_{25}H_{52} + 38O_2 \rightarrow 25CO_2 + 26H_2O$); the reaction is exothermic; mass is conserved though the candle shrinks (products are gases); carbon is oxidized; soot from incomplete combustion.

Printable answer key (PDF, for the parent)

UNIT 5 · CHEMICAL REACTIONS · LESSON 5.1 OF 36 · 4 SESSIONS OF 45 MINUTES

5.1 Writing and Balancing Equations

OBJECTIVES

- Translate word equations into formula equations with state symbols.
- Balance equations by adjusting coefficients, never subscripts.

- Explain balancing as conservation of atoms and mass.

> **Warm-up.** A recipe makes 12 cookies from 2 cups flour and 1 egg. How much flour and how many eggs for 36 cookies? What stays the same and what changes?

WATCH

Introduction to Balancing Chemical Equations

Tyler DeWitt · 12:23

https://www.youtube.com/watch?v=yA3TZJ2em6g

> **WHEN TO WATCH**
> After the core ideas.

BACKUP VIDEO

Balancing Chemical Equations Step by Step Practice Problems

Melissa Maribel · 6:23

https://www.youtube.com/watch?v=jy6F0Lbvjm8

After the first worked example.

KEY TERMS

chemical equation
A symbolic statement of a reaction showing reactants, products and their amounts.

reactant
A starting substance in a reaction, written on the left.

product
A substance formed in a reaction, written on the right.

coefficient
A number in front of a formula giving how many units take part.

subscript
A small number within a formula giving how many atoms of an element are in one unit.

state symbol
A label (s), (l), (g) or (aq) showing a substance's physical state.

CORE IDEAS

The story. Lavoisier's law of conservation of mass (Unit 2) made equations possible: if atoms are neither created nor destroyed, the left side must match the right. Jöns Jacob Berzelius gave chemistry its letter symbols in 1813, and by the mid-1800s chemists wrote reactions much as we do now.

How they figured it out. An equation is a bookkeeping statement. Hydrogen burns in oxygen to form water: $H_2 + O_2 \rightarrow H_2O$ has two O on the left and one on the right. Changing the subscript to make H_2O_2 would describe hydrogen peroxide, a different substance. Instead, change coefficients: $2H_2 + O_2 \rightarrow 2H_2O$.

The chemistry. Steps: (1) write correct formulas; (2) count atoms of each element on each side; (3) balance elements that appear in only one substance on each side first, leaving H and O for later and free elements for last; (4) treat an unchanged polyatomic ion as a unit; (5) reduce coefficients to the lowest whole numbers; (6) check. State symbols: (s) solid, (l) liquid, (g) gas, (aq) dissolved in water. An arrow means "yields"; a triangle above it means heat.

Where it is used. Balanced equations are the recipes of industry: how much iron ore for a ton of steel, how much oxygen a rocket engine must carry, how much baking soda a recipe needs. Unit 6 turns these coefficients into masses.

LAB

Conservation of Mass in a Bag

You will run baking soda and vinegar in a sealed bag on a kitchen scale and write the balanced equation.

You will need

- Kitchen scale (0.1 g if possible), zip-top bag, small cup
- Baking soda, vinegar

Steps

1. Put 1 tablespoon of baking soda in the bag. Pour 30 mL of vinegar into the small cup and stand it upright in the bag.
2. Squeeze out air, seal, and weigh everything.
3. Tip the cup to mix. When fizzing stops, weigh again.
4. Open the bag, let the gas escape, and weigh again.

Check your work

☐ Did the mass change in the sealed bag? After opening?

☐ Write the balanced equation:

$$\mathrm{NaHCO_3} + \mathrm{HC_2H_3O_2} \rightarrow \mathrm{NaC_2H_3O_2} + \mathrm{H_2O} + \mathrm{CO_2}$$

LAB NOTEBOOK

1. Three masses and an explanation of each.
2. The balanced equation with state symbols.

> **From history.** Berzelius, who invented our element symbols, also discovered cerium, selenium and thorium and coined the words “protein,” “catalysis” and “polymer.”

Another video. Magnesium Burns in Carbon Dioxide (reaction only). A short clip with no narration: watch it twice, then describe what happened in your own words.



For parents: Periodic Videos, 0:37. A wordless 37-second clip: burning magnesium keeps burning inside a block of dry ice (carbon dioxide). Periodic Videos is an approved channel.
Periodic Videos · 0:37

WORKED EXAMPLES

Example 1. Balance: $\mathrm{CH_4 + O_2 \rightarrow CO_2 + H_2O}$.

▼ **Show solution**

1. C: 1 = 1.
2. H: 4 on the left, so 2 $\mathrm{H_2O}$.
3. O: right side has 2 + 2 = 4, so 2 $\mathrm{O_2}$.
4. Check: C 1, H 4, O 4 on each side.

Answer: $\mathrm{CH_4 + 2O_2 \rightarrow CO_2 + 2H_2O}$

Example 2. Balance: $\mathrm{Al} + \mathrm{CuSO_4} \rightarrow \mathrm{Al_2(SO_4)_3} + \mathrm{Cu}$.

▼ **Show solution**

1. Treat sulfate as a unit: 3 on the right, so 3 $\mathrm{CuSO_4}$.
2. Then 3 Cu on the right.
3. 2 Al on the left.

Answer: $2\mathrm{Al} + 3\mathrm{CuSO_4} \rightarrow \mathrm{Al_2(SO_4)_3} + 3\mathrm{Cu}$

PRACTICE

1. Easy Why can't you change subscripts to balance an equation?

   ▼ **Show answer**

   It would change the identity of the substance.

2. Easy What does (aq) mean?

   ▼ **Show answer**

   Dissolved in water (aqueous).

3. Medium Balance: $\mathrm{CH_4 + O_2 \rightarrow CO_2 + H_2O}$

   ▼ **Show answer**

   $\mathrm{CH_4 + 2O_2 \rightarrow CO_2 + 2H_2O}$

4. Medium Balance: $\mathrm{Mg + HCl \rightarrow MgCl_2 + H_2}$

   ▼ **Show answer**

   $\mathrm{Mg + 2HCl \rightarrow MgCl_2 + H_2}$

5. Medium Balance: $\mathrm{Al + O_2 \rightarrow Al_2O_3}$

   ▼ **Show answer**

   $\mathrm{4Al + 3O_2 \rightarrow 2Al_2O_3}$

COMMON MISTAKES

- Changing subscripts to balance.
- Forgetting diatomic elements: $\mathrm{H_2}$, $\mathrm{N_2}$, $\mathrm{O_2}$, $\mathrm{F_2}$, $\mathrm{Cl_2}$, $\mathrm{Br_2}$, $\mathrm{I_2}$.
- Leaving coefficients that could be reduced (4, 2, 4 should be 2, 1, 2).

Make it a game · Coefficient Bingo. The parent writes an unbalanced equation; the student lays coins as coefficients. The parent checks with a quick atom count. Three correct in a row wins.

Scripture. "Just balances, just weights, a just ephah, and a just hin, shall ye have: I am the LORD your God, which brought you out of the land of Egypt." (Leviticus 19:36, KJV)

Just balances and just weights were commanded in the law. A balanced equation is an honest account of every atom.

UNIT 5 · CHEMICAL REACTIONS · LESSON 5.2 OF 36 · 4 SESSIONS OF 45 MINUTES

5.2 Five Reaction Types and Faraday's Chemical History of a Candle (1848)

OBJECTIVES

- Classify reactions as synthesis, decomposition, single replacement, double replacement or combustion.
- Predict products for each type.
- Use the activity series to decide whether a single replacement occurs.

Warm-up. Light a candle (with an adult). List every change you observe in two minutes. Which are physical and which chemical?

WATCH

Types of Chemical Reactions

Tyler DeWitt · 12:54

https://www.youtube.com/watch?v=aMU1RaRulSo

> **WHEN TO WATCH**
> After the warm-up, for Faraday's candle.

BACKUP VIDEO

Types of Chemical Reactions

Melissa Maribel · 3:00

https://www.youtube.com/watch?v=z0187f8gqfU

After the activity series.

KEY TERMS

synthesis

A reaction in which two or more substances combine into one: A + B → AB.

decomposition

A reaction in which one substance breaks into two or more: AB → A + B.

single replacement

An element replaces another in a compound: A + BC → AC + B.

double replacement

Two compounds exchange ions: AB + CD → AD + CB.

combustion

A rapid reaction with oxygen, usually producing heat and light; hydrocarbons give CO_2 and H_2O.

activity series

A list of metals in order of reactivity; a metal replaces any metal below it.

CORE IDEAS

The story. Michael Faraday, a blacksmith's son who became the greatest experimenter of his age, began the Royal Institution's Christmas Lectures for young people in 1825. His 1848 series, published as *The Chemical History of a Candle*, showed

children that a single candle displayed nearly every law of chemistry. He collected the water from a flame, showed that the black soot was carbon, and proved the flame needed air. Faraday was a devout member of a small Christian congregation, the Sandemanians, and served as an elder.

How they figured it out. Chemists sorted thousands of reactions by pattern. In synthesis, elements combine: $2Mg + O_2 \rightarrow 2MgO$. In decomposition, heat or electricity breaks a compound: $2H_2O_2 \rightarrow 2H_2O + O_2$. In single replacement a more active metal displaces a less active one: $Zn + CuSO_4 \rightarrow ZnSO_4 + Cu$. In double replacement ions swap partners, often forming a precipitate, gas or water. Combustion of a hydrocarbon gives carbon dioxide and water.

The chemistry. Activity series (most to least active): K, Na, Ca, Mg, Al, Zn, Fe, Ni, Sn, Pb, (H), Cu, Ag, Au. A metal replaces any metal (or hydrogen in acids) below it. So zinc dissolves in hydrochloric acid releasing hydrogen, but copper does not, and gold is untouched by most acids. A double replacement goes forward only if it forms a precipitate (Lesson 5.3), a gas or water.

Where it is used. Combustion powers engines and furnaces. Decomposition of sodium azide inflates car airbags in milliseconds. Single replacement (the thermite reaction of aluminum with iron oxide) welds railroad rails. Double replacement neutralizes stomach acid with antacids.

> **Safety first.** Goggles on. Copper sulfate is toxic and an eye irritant; adult handles it, gloves on, wash hands. Hydrogen peroxide foam is safe to touch but keep it from eyes. Only an adult uses flames.

LAB

Five Reactions at the Kitchen Table

You will carry out one example of each reaction type and classify it.

You will need

- Steel wool, vinegar, a jar (synthesis: rusting)
- 3% hydrogen peroxide, a pinch of dry yeast (decomposition)
- Iron nail, copper sulfate solution (single replacement; adult handles)
- Epsom salt and washing soda solutions (double replacement)
- A candle and a cold glass (combustion)

Steps

1. Soak steel wool in vinegar for 1 minute, squeeze, and put it in a jar with a thermometer. Watch for warming and rust (synthesis of iron oxide).
2. Add yeast to 50 mL hydrogen peroxide in a tall glass. Test the foam with a glowing splint (adult).
3. Place a clean nail in copper sulfate solution for 20 minutes. Observe the coating.
4. Mix Epsom salt solution with washing soda solution. Observe.
5. Hold a cold, dry glass over a candle flame for a few seconds. Observe the inside.

Check your work

- ☐ Write and classify an equation for each reaction.
- ☐ Which reactions gave off heat?

LAB NOTEBOOK

1. Observations for all five reactions.
2. Five balanced equations, each labeled with its type.

> **From history.** Faraday refused a knighthood and declined to help the British government develop poison gas for the Crimean War, saying he could but would not.

Another video. Thermite Reaction (reaction only). A short clip with no narration: watch it twice, then describe what happened in your own words.



For parents: Periodic Videos, 0:27. A wordless 27-second clip of the thermite reaction (molten iron and sparks). Periodic Videos is an approved channel. Periodic Videos · 0:27

WORKED EXAMPLES

Example 1. Classify and complete: $\mathrm{Mg} + \mathrm{HCl} \rightarrow$

▼ **Show solution**

1. Mg is above H in the activity series, so it replaces hydrogen.
2. Products: $\mathrm{MgCl_2}$ and $\mathrm{H_2}$.
3. Balance: $\mathrm{Mg} + 2\mathrm{HCl} \rightarrow \mathrm{MgCl_2} + \mathrm{H_2}$.

Answer: Single replacement: $\mathrm{Mg} + 2\mathrm{HCl} \rightarrow \mathrm{MgCl_2} + \mathrm{H_2}$

Example 2. Classify and complete the combustion of ethanol, $\mathrm{C_2H_5OH}$.

▼ **Show solution**

1. Products are $\mathrm{CO_2}$ and $\mathrm{H_2O}$.
2. C: 2 $\mathrm{CO_2}$. H: 6, so 3 $\mathrm{H_2O}$. O on right 4 + 3 = 7; ethanol supplies 1, so 3 $\mathrm{O_2}$.

Answer: $\mathrm{C_2H_5OH} + 3\mathrm{O_2} \rightarrow 2\mathrm{CO_2} + 3\mathrm{H_2O}$

PRACTICE

1. **Easy** What type of reaction is $A + B \rightarrow AB$?

   ▼ **Show answer**

   Synthesis.

2. **Easy** What are the products of the complete combustion of a hydrocarbon?

   ▼ **Show answer**

   Carbon dioxide and water.

3. **Medium** Will copper react with zinc sulfate solution? Explain.

   ▼ **Show answer**

   No; copper is below zinc in the activity series.

COMMON MISTAKES

- Predicting copper will replace zinc (check the activity series).
- Forgetting oxygen is diatomic in combustion.
- Calling any reaction with oxygen a synthesis; hydrocarbon burning is combustion.

> **Make it a game · Type Sort.** Write 12 equations on cards. The student sorts them into five piles by type in under 2 minutes.

> **Scripture.** “The heavens declare the glory of God; and the firmament sheweth his handywork.” (Psalm 19:1, KJV)
>
> Faraday believed the creation spoke of its Maker. The heavens declare God’s glory, and so does a candle flame studied carefully.

5.3 Precipitates, Solubility Rules and Net Ionic Equations

OBJECTIVES

- Use solubility rules to predict whether an ionic compound dissolves.
- Predict precipitates in double replacement reactions.

- Write complete and net ionic equations and identify spectator ions.

> **Warm-up.** Stir a spoon of Epsom salt into one cup of water and a spoon of chalk dust into another. What differences do you see after 5 minutes?

WATCH

Solubility Rules and How to Use a Solubility Table

Melissa Maribel · 7:35

https://www.youtube.com/watch?v=KOrq5i0rXUA

> **WHEN TO WATCH**
> After the solubility rules.

BACKUP VIDEO

How to Write Complete Ionic Equations and Net Ionic Equations

Melissa Maribel · 9:03

https://www.youtube.com/watch?v=MJy5CNzd6sY

With the net ionic practice.

KEY TERMS

soluble
Able to dissolve to a significant extent in a solvent.

electrolyte
A substance whose solution conducts electricity because it contains ions.

net ionic equation
An equation showing only the particles that actually react.

precipitate
An insoluble solid formed when two solutions are mixed.

spectator ion
An ion present on both sides of an ionic equation, unchanged.

dissociation
The separation of an ionic compound into its ions in water.

CORE IDEAS

The story. Analytical chemists of the 1700s and 1800s, such as Torbern Bergman in Sweden, identified the contents of mineral waters and ores by adding reagents and watching which solids fell out. A white curd with silver nitrate meant chloride; a white solid with barium meant sulfate. These tests, still used in schools, rest on a handful of solubility rules.

How they figured it out. Arrhenius showed in 1884 (Unit 8) that salts dissolve as separate ions. When solutions of silver nitrate and sodium chloride mix, the four ions Ag^+, NO_3^-, Na^+, Cl^- meet. Silver chloride is insoluble, so Ag^+ and Cl^- lock together as a solid; sodium and nitrate stay in solution unchanged.

The chemistry. Solubility rules: (1) all compounds of group 1 ions and ammonium are soluble; (2) all nitrates and acetates are soluble; (3) chlorides, bromides and iodides are soluble except with Ag^+, Pb^{2+} and Hg_2^{2+}; (4) sulfates are soluble except with Ba^{2+}, Pb^{2+}, Ca^{2+} (slightly) and Sr^{2+}; (5) carbonates, phosphates and hydroxides are insoluble except with group 1 and ammonium (and hydroxides of Ba^{2+}, Ca^{2+} slightly). To write a net ionic equation: write the balanced molecular equation, split soluble compounds (aq) into ions, cancel spectators.

Where it is used. Hard water contains calcium and magnesium; with soap or carbonate it forms scum and scale. Water treatment plants remove phosphate by precipitating it. Kidney stones are precipitates of calcium oxalate or phosphate. Barium sulfate is so insoluble that patients drink it safely for X-ray studies.

> **Safety first.** Goggles on. Calcium chloride warms when it dissolves and irritates skin; wash hands. Washing soda is caustic: avoid eyes.

LAB

Kitchen Precipitates

You will mix pairs of household solutions and predict, then observe, which form precipitates.

You will need

- Epsom salt ($MgSO_4$), washing soda (Na_2CO_3), baking soda, table salt, calcium chloride (ice melt or pickling crisp)
- Clear cups, distilled water, spoons, labels

Steps

1. Make a teaspoon-in-half-cup solution of each salt in distilled water. Label them.
2. Before mixing, use the solubility rules to predict which pairs make a solid.
3. Mix 1 tablespoon of each pair in a clear cup: Epsom + washing soda; calcium chloride + washing soda; calcium chloride + table salt; Epsom + table salt.
4. Record cloudy or clear. Compare with predictions.

Check your work

- ☐ Which pairs formed a precipitate? What is its formula?
- ☐ Write net ionic equations for each precipitate.

LAB NOTEBOOK

1. A prediction and result table for each pair.
2. Net ionic equations for each precipitate.

> **From history.** Doctors in the 1890s first used bismuth salts and then, from the 1910s, barium sulfate to make the digestive tract show on X-rays. Soluble barium salts are poisonous; only the insoluble sulfate is safe.

WORKED EXAMPLES

Example 1. Predict the product and write the net ionic equation for mixing lead(II) nitrate and potassium iodide solutions.

▼ **Show solution**

1. Swap: $\mathrm{PbI_2}$ and $\mathrm{KNO_3}$.
2. $\mathrm{PbI_2}$ is insoluble (rule 3); $\mathrm{KNO_3}$ soluble.
3. Molecular: $\mathrm{Pb(NO_3)_2}(aq) + 2\mathrm{KI}(aq) \rightarrow \mathrm{PbI_2}(s) + 2\mathrm{KNO_3}(aq)$.
4. Cancel $\mathrm{K^+}$ and $\mathrm{NO_3^-}$.

Answer: $\mathrm{Pb^{2+}}(aq) + 2\mathrm{I^-}(aq) \rightarrow \mathrm{PbI_2}(s)$, a bright yellow precipitate

Example 2. Will a precipitate form when sodium nitrate and potassium chloride solutions mix?

▼ **Show solution**

1. Possible products NaCl and $\mathrm{KNO_3}$.
2. Both are soluble.

Answer: No reaction; all ions are spectators

PRACTICE

1. Easy Are all sodium compounds soluble?

   ▼ **Show answer**

   Yes.

2. Easy Is silver chloride soluble?

   ▼ **Show answer**

   No; it is a white precipitate.

3. Medium Is calcium carbonate soluble? Where is it found?

   ▼ **Show answer**

   No; it is chalk, limestone and seashells.

COMMON MISTAKES

- Splitting a solid precipitate into ions.
- Forgetting to balance charges in the net ionic equation.
- Writing a net ionic equation when no precipitate forms.

> **Make it a game · Soluble or Not?.** The parent names an ionic compound. The student answers “soluble” or “precipitate” and names the rule. Ten in a row earns a prize.

> **Scripture.** “Take away the dross from the silver, and there shall come forth a vessel for the finer.” (Proverbs 25:4, KJV)
>
> Taking the dross from silver is a separation. Precipitation lets a chemist take one ion out of a mixture.

UNIT 5 · CHEMICAL REACTIONS · LESSON 5.4 OF 36 · 4 SESSIONS OF 45 MINUTES

5.4 Oxidation and Reduction: Rust, Batteries and Davy’s Electrolysis (1807)

OBJECTIVES

- Assign oxidation numbers.
- Identify oxidation, reduction, oxidizing agent and reducing agent.
- Describe how a battery and electrolysis use redox reactions.

> **Warm-up.** Leave a cut apple slice and a slice dipped in lemon juice on a plate. Predict which browns first. Why might lemon juice help?

WATCH

Introduction to Oxidation Reduction (Redox) Reactions

Tyler DeWitt · 13:05

https://www.youtube.com/watch?v=5rtJdjas-mY

> **WHEN TO WATCH**
> After the oxidation-number rules.

BACKUP VIDEO

Oxidation and Reduction (Redox) Reactions Step-by-Step Example

Melissa Maribel · 3:56

https://www.youtube.com/watch?v=6gtcvqLmPo8

With the battery discussion.

KEY TERMS

oxidation
Loss of electrons; an increase in oxidation number.

reduction
Gain of electrons; a decrease in oxidation number.

oxidation number
The charge an atom would have if all its bonds were ionic.

oxidizing agent
The substance that gains electrons and is reduced, causing oxidation of another.

electrolysis
Using electric current to drive a nonspontaneous reaction.

electrochemical cell
A device that turns the energy of a redox reaction into electric current (a battery).

CORE IDEAS

The story. In 1800 Alessandro Volta stacked zinc and copper discs separated by brine-soaked cardboard and made the first battery. Within weeks chemists used it to split water. In 1807 Humphry Davy at the Royal Institution passed current from a

huge battery through melted potash. Tiny globules of metal burst through the crust and caught fire: potassium. Days later he isolated sodium, and the next year calcium, barium, magnesium and strontium. His assistant, hired in 1813, was a young bookbinder named Michael Faraday.

How they figured it out. The word oxidation first meant combining with oxygen, as in rusting. Chemists realized the key was electrons. When iron rusts, iron atoms lose electrons to oxygen. Loss of electrons is oxidation; gain is reduction. The two always happen together: “LEO the lion says GER” (Lose Electrons Oxidation, Gain Electrons Reduction).

The chemistry. Rules for oxidation numbers: elements alone are 0; a monatomic ion equals its charge; oxygen is usually −2 (−1 in peroxides); hydrogen +1 with nonmetals; group 1 is +1, group 2 is +2; the sum equals the overall charge. In $\mathrm{Zn} + \mathrm{Cu}^{2+} \rightarrow \mathrm{Zn}^{2+} + \mathrm{Cu}$, zinc is oxidized (0 → +2) and is the reducing agent; copper ion is reduced (+2 → 0) and is the oxidizing agent. In a battery, the two half-reactions are separated so electrons travel through a wire. Electrolysis runs the reaction backward with outside current.

Where it is used. Batteries in phones and cars, rusting and its prevention (paint, galvanizing with zinc), bleaching, photography, refining aluminum by the Hall–Héroult process (1886), electroplating, and the respiration in every cell (Biology Unit 4) are redox reactions.

> **Safety first.** Adult handles copper sulfate and the 9 V battery. Do not eat the lemons afterward. Wash hands.

LAB

A Lemon Battery and Copper Plating

You will build a battery from fruit and use current to plate copper onto a key.

You will need

- Lemons (4), galvanized nails or zinc strips, copper coins or wire, alligator clips, an LED or a multimeter
- A 9 V battery, copper wire, a brass key, vinegar with a pinch of salt or a little copper sulfate solution (adult)

Steps

1. Push a zinc nail and a copper coin into a lemon. Measure the voltage with a multimeter.
2. Connect four lemons in series (zinc to copper) and try to light an LED.
3. Electroplating (adult): connect the key to the battery's negative terminal and a copper wire to the positive. Dip both in the copper solution, not touching, for 10 minutes.
4. Observe the key.

Check your work

☐ Which metal was oxidized in the lemon battery?

☐ At which electrode was copper reduced during plating?

LAB NOTEBOOK

1. Voltages for 1, 2, 3 and 4 lemons.
2. Half-reactions for the battery and for plating.

> **From history.** Davy was so delighted by potassium that he danced around the laboratory. He also invented the miner's safety lamp (1815), which saved many lives, and refused to patent it.

Another video. Copper Chloride and Aluminium (reaction only). A short clip with no narration: watch it twice, then describe what happened in your own words.



For parents: Periodic Videos, 0:30. A wordless 30-second clip: aluminium foil reacts with copper chloride solution, fizzing and turning brown. Periodic Videos is an approved channel. Periodic Videos · 0:30

WORKED EXAMPLES

Example 1. Assign oxidation numbers in $KMnO_4$.

▼ **Show solution**

1. K +1, each O −2 (total −8).
2. Sum zero: $+1 + \text{Mn} - 8 = 0$, so Mn = +7.

Answer: K +1, Mn +7, O −2

Example 2. In $2Fe_2O_3 + 3C \rightarrow 4Fe + 3CO_2$, identify what is oxidized and reduced.

▼ **Show solution**

1. Fe: +3 → 0, gained electrons: reduced. Fe_2O_3 is the oxidizing agent.
2. C: 0 → +4, lost electrons: oxidized. C is the reducing agent.

Answer: Carbon oxidized; iron reduced

PRACTICE

1. Easy What is oxidation in terms of electrons?

   ▼ **Show answer**

   Loss of electrons.

2. Easy What is the oxidation number of any element alone, such as O_2?

   ▼ **Show answer**

   0

3. Medium Find the oxidation number of S in H_2SO_4.

   ▼ **Show answer**

   +6

COMMON MISTAKES

- Thinking oxidation always involves oxygen.
- Mixing up the oxidizing agent (which is reduced) and the reducing agent (which is oxidized).
- Forgetting that oxidation and reduction occur together.

> **Make it a game · LEO and GER.** The parent says a change ("Fe to Fe^{3+}"). The student roars (LEO, oxidation) or says "grr" (GER, reduction) and gives the number of electrons.

> **Scripture.** "Lay not up for yourselves treasures upon earth, where moth and rust doth corrupt, and where thieves break through and steal:" (Matthew 6:19, KJV)
>
> Rust is oxidation at work. Jesus warned against treasures that moth and rust corrupt.

UNIT 6

The Mole and Stoichiometry

Chemists cannot count atoms one at a time, so they count them by the mole, a number so large that a mole of grains of sand would bury a continent. This unit introduces Avogadro's number, molar mass, percent composition and empirical formulas, and then the core skill of the course: stoichiometry, the recipe math that

says how many grams of product a given mass of reactant will make. It closes with limiting reactants and percent yield, told through Haber and Bosch's ammonia process, which feeds about half the world.

Expect practice every day. Each worked example shows the factor-label setup; students should copy that format.

PREREQUISITES

- Unit 1: significant figures and dimensional analysis
- Unit 5: balanced equations

MASTERY CHECKLIST

- ☐ I can convert among particles, moles and grams.
- ☐ I can calculate molar mass, percent composition and empirical formulas.
- ☐ I can use mole ratios from a balanced equation to find the mass of a product or reactant.
- ☐ I can identify the limiting reactant and calculate theoretical and percent yield.

LESSONS IN THIS UNIT

6.1	**Avogadro and the Mole**	4 sessions of 45 minutes
6.2	**Molar Mass, Percent Composition and Empirical Formulas**	4 sessions of 45 minutes
6.3	**Stoichiometry: Chemistry's Recipe Math**	4 sessions of 45 minutes
6.4	**Limiting Reactant and Percent Yield: Haber and Bosch (1909–1913)**	4 sessions of 45 minutes

Unit 6 test

Show your work. For every calculation write the equation, substitute with units and give the answer with its unit and sensible significant figures. The test is worth 100 points.

1. **7 pts** **Lesson 6.1**

   What date is Mole Day?

2. **7 pts** **Lesson 6.1**

   Name the representative particle of NaCl, of H_2O and of Fe.

3. **7 pts** **Lesson 6.1**

   How many moles is 5.42×10^{22} atoms of iron?

4. **6 pts** **Lesson 6.2**

   Find the percent oxygen in water.

5. **6 pts** **Lesson 6.2**

   Why do we assume a 100 g sample when finding an empirical formula?

6. **6 pts** **Lesson 6.2**

   What is the percent by mass of H in C_2H_5OH?

7. **6 pts** **Lesson 6.3**

   Who coined the word stoichiometry?

8. **6 pts** **Lesson 6.3**

   How many liters of CO_2 at STP form from 2.0 mol of propane burning?

9. **6 pts** **Lesson 6.3**

   For $4Fe + 3O_2 \rightarrow 2Fe_2O_3$, what mass of Fe_2O_3 forms from 53.7 g of Fe?

10. **6 pts** **Lesson 6.4**

    What catalyst did Bosch's team use?

11. **6 pts** **Lesson 6.4**

What moral lesson does Haber’s life illustrate?

12. **6 pts** **Lesson 6.4**

For $CH_4 + 2O_2 \rightarrow CO_2 + 2H_2O$, 10.5 g of CH_4 reacts with 5.4 g of O_2. (a) Which is the limiting reactant? (b) What is the theoretical yield of CO_2? (c) If 3.1 g is actually collected, what is the percent yield?

LAB AND DATA

13. **10 pts**

You react 2.00 g of baking soda ($NaHCO_3$, 84.01 g/mol) with excess vinegar and evaporate the solution, leaving 1.85 g of sodium acetate ($NaC_2H_3O_2$, 82.03 g/mol). The ratio is 1:1. (a) Find the theoretical yield. (b) Find the percent yield. (c) Give one reason the yield is not 100%.

SHORT ESSAY

14. **15 pts**

Explain why the Haber–Bosch process is called one of the most important inventions of the twentieth century. Include the balanced equation, why nitrogen is hard to use, and what limiting reactant and yield mean for a factory.

ANSWER KEY

1. October 23.
2. Formula unit, molecule, atom.
3. 0.0900 mol

   $5.42 \times 10^{22} / (6.022 \times 10^{23}) = 0.0900$ mol
4. 88.8%
5. Each percent becomes grams directly, simplifying the math; the ratio does not depend on sample size.
6. 13.1%

   6 × 1.008 = 6.05; ÷ 46.07 × 100 = 13.1%
7. Richter, 1792.
8. 6.0 mol × 22.4 = 134 L
9. 76.8 g

   53.7 g ÷ 55.85 g/mol = 0.962 mol of Fe · Mole ratio 2:4 → 0.481 mol Fe_2O_3 · $0.481 \times 159.70 =$ 76.8 g
10. An iron-based catalyst (found by Mittasch).
11. Scientific knowledge can feed people or kill them; scientists are responsible for how they use it.
12. (a) O_2; (b) 3.71 g; (c) 83.5%

    CH_4: 0.655 mol → 0.655 mol CO_2 · O_2: 0.169 mol → 0.0844 mol CO_2 · The smaller amount comes from O_2, so it limits: 0.0844 mol × 44.01 = 3.71 g · Percent yield = 3.1/3.71 × 100 = 83.5%
13. (a) $2.00/84.01 = 0.0238$ mol → $0.0238 \times 82.03 = 1.95$ g. (b) $1.85/1.95 \times 100 = 94.9\%$. (c) Spattering during evaporation, product left on the dish, incomplete reaction or the product still damp (which would raise it).
14. Look for: $N_2 + 3H_2 \rightleftharpoons 2NH_3$; nitrogen's triple bond makes air's nitrogen unavailable to most plants; Haber (1909) found conditions and an iron catalyst, Bosch scaled it up at BASF (1913) with high-pressure steel reactors; ammonia fertilizer greatly increased crop yields; factories add excess of the cheaper reactant and recycle unreacted gas to improve yield; also used for explosives in war (Haber's chlorine gas work is a moral warning).

Printable answer key (PDF, for the parent)

6.1 Avogadro and the Mole

OBJECTIVES

- Define the mole and Avogadro's number.
- Convert between number of particles and moles.
- Explain how Avogadro's hypothesis led to the mole concept.

> **Warm-up.** A dozen is 12, a gross is 144, a ream of paper is 500 sheets. Why do people invent counting units like these? What would be a good unit for counting atoms?

WATCH

Introduction to Moles

Tyler DeWitt · 10:50

https://www.youtube.com/watch?v=wI56mHUDJgQ

> **WHEN TO WATCH**
> After the story.

BACKUP VIDEO

The Most Common Mistake with Moles!

Tyler DeWitt · 4:12

https://www.youtube.com/watch?v=wHwGm9oL-GA

After the first conversions.

KEY TERMS

mole
The amount of substance containing 6.022×10^{23} particles.

Avogadro's number
6.022×10^{23}, the number of particles in one mole.

representative particle

The unit a substance is made of: an atom, a molecule or a formula unit.

Avogadro's hypothesis

Equal volumes of gases at the same temperature and pressure contain equal numbers of particles (1811).

conversion factor

A ratio equal to one used to change units, such as $1\text{ mol}/6.022 \times 10^{23}$.

formula unit

The simplest ratio of ions in an ionic compound.

CORE IDEAS

The story. In 1811 the Italian count Amedeo Avogadro proposed that equal volumes of different gases, at the same temperature and pressure, hold equal numbers of particles. This explained Gay-Lussac's finding that gases combine in simple volume ratios (two volumes of hydrogen with one of oxygen). Almost no one listened. In 1860, four years after Avogadro's death, Stanislao Cannizzaro handed out a pamphlet at the Karlsruhe Congress that showed how the hypothesis gave consistent atomic masses, and chemists were convinced.

How they figured it out. If equal volumes hold equal numbers, then the masses of equal volumes give the relative masses of the particles. Chemists defined a mole as the number of atoms in exactly 12 g of carbon-12, so the atomic mass in amu equals the mass of a mole in grams. The actual number was measured later by several methods; Jean Perrin, studying Brownian motion around 1909, measured it and named it for Avogadro. Since 2019 the mole is defined as exactly $6.02214076 \times 10^{23}$ particles.

The chemistry. Moles are a counting unit like dozens. To convert particles to moles, divide by 6.022×10^{23}; moles to particles, multiply. Be clear what particle you count: one mole of water molecules contains two moles of hydrogen atoms and one mole of oxygen atoms.

Where it is used. Every recipe in chemistry, pharmacy and industry is written in moles because reactions happen particle by particle. A mole of water is only 18 g, about a tablespoon, yet it holds 6×10^{23} molecules.

LAB

Counting by Weighing

You will find the number of rice grains, beans and paper clips in a sample without counting them all, the same trick chemists use with the mole.

You will need

- A kitchen scale (0.1 g), uncooked rice, dried beans, paper clips
- Small cups

Steps

1. Weigh exactly 100 grains of rice. Find the mass of one grain.
2. Weigh a large scoop of rice and calculate how many grains it holds.
3. Count by hand to check (or count a quarter of it and multiply).
4. Repeat with beans and paper clips. Then find how many grams of each make a “mole” of 1000 items.

Check your work

☐ How close was the weighed count to the hand count?

☐ Why do chemists count atoms by weighing?

LAB NOTEBOOK

1. Data table for each item with the calculated and counted values.
2. A paragraph connecting the method to molar mass.

KEY FORMULAS

$$N = n \times N_A \qquad N_A = 6.022 \times 10^{23}\ \text{mol}^{-1}$$

> **From history.** October 23 is celebrated by chemistry teachers as Mole Day, from 6:02 a.m. to 6:02 p.m., for 6.02×10^{23}.

WORKED EXAMPLES

Example 1. How many atoms are in 2.50 mol of copper?

▼ **Show solution**

1. $2.50 \text{ mol} \times 6.022 \times 10^{23} \text{ atoms/mol}$

Answer: 1.51×10^{24} atoms

Example 2. How many moles of hydrogen atoms are in 3.0 mol of methane, CH_4?

▼ **Show solution**

1. Each CH_4 has 4 H atoms.
2. $3.0 \times 4 = 12$

Answer: 12 mol H atoms

PRACTICE

1. Easy What is Avogadro's number?

   ▼ **Show answer**

   6.022×10^{23}

2. Easy How many atoms are in 1 mol of helium?

   ▼ **Show answer**

   6.022×10^{23}

3. Easy How many molecules are in 2.42 mol of water?

   ▼ **Show answer**

   1.46×10^{24}

4. Medium How many moles is 7.64×10^{22} atoms of iron?

   ▼ **Show answer**

   0.127 mol

5. Hard How many molecules (formula units) are in 46.8 g of $CaCO_3$?

▼ **Show answer**

2.82×10^{23}

COMMON MISTAKES

- Dividing when you should multiply by Avogadro's number.
- Not naming the particle being counted (atoms vs molecules).
- Forgetting scientific notation on the calculator (use the EE or $\times 10^x$ key).

> **Make it a game · Mole Estimates.** The parent asks how many moles of something everyday: grains of sand on a beach, stars visible to the eye, people on earth. The student estimates in scientific notation and compares with a mole.

> **Scripture.** "That in blessing I will bless thee, and in multiplying I will multiply thy seed as the stars of the heaven, and as the sand which is upon the sea shore; and thy seed shall possess the gate of his enemies;" (Genesis 22:17, KJV)
>
> God promised Abraham descendants as the stars and the sand. A mole of sand grains would far outnumber both.

UNIT 6 · THE MOLE AND STOICHIOMETRY · LESSON 6.2 OF 36 · 4 SESSIONS OF 45 MINUTES

6.2 Molar Mass, Percent Composition and Empirical Formulas

OBJECTIVES

- Calculate the molar mass of a compound.
- Convert between grams and moles.
- Calculate percent composition and find empirical and molecular formulas.

> **Warm-up.** A bag holds 12 apples (200 g each) and 12 grapes (5 g each). What percent of the bag's mass is apples? How is a compound like this bag?

WATCH

Percent Composition By Mass

Tyler DeWitt · 6:58

https://www.youtube.com/watch?v=lywmGCfIUIA

> **WHEN TO WATCH**
> After the molar mass examples.

BACKUP VIDEO

Empirical Formula and Molecular Formula Introduction

Tyler DeWitt · 8:31

https://www.youtube.com/watch?v=wnRaBWvhYKY

With empirical formulas.

KEY TERMS

molar mass

The mass in grams of one mole of a substance (g/mol).

percent composition

The percent by mass of each element in a compound.

empirical formula

The simplest whole-number ratio of atoms in a compound.

molecular formula

The actual number of atoms of each element in a molecule; a whole-number multiple of the empirical formula.

combustion analysis

Burning a compound and weighing the CO_2 and H_2O produced to find its composition.

hydrate

An ionic compound with water molecules in its crystal, such as $CuSO_4 \cdot 5H_2O$.

CORE IDEAS

The story. How do you find the formula of a new substance? In the 1830s Justus von Liebig in Giessen perfected combustion analysis: burn a weighed sample, trap the

water in calcium chloride and the carbon dioxide in potash bulbs, weigh them, and work backward to the percent of carbon and hydrogen. Students came from all over Europe to his teaching laboratory, the first of its kind.

How they figured it out. Percent composition tells the mass ratio of elements. Dividing each mass by atomic mass turns it into a mole ratio, which is the atom ratio. Assume a 100 g sample so that each percent becomes grams. That gives the empirical formula. A molar mass measurement (from gas density or a mass spectrometer) then gives the molecular formula.

The chemistry. Molar mass is the sum of atomic masses: $\mathrm{H_2O} = 2(1.01) + 16.00 = 18.02$ g/mol. Grams ÷ molar mass = moles; moles × molar mass = grams. Percent composition = (mass of element in one mole ÷ molar mass) × 100. For an empirical formula: grams → moles → divide by the smallest → round (multiply by 2 or 3 if you get .5 or .33). Molecular formula = empirical formula × (molar mass ÷ empirical formula mass); for glucose, $\mathrm{CH_2O}$ (30) × 6 = $\mathrm{C_6H_{12}O_6}$ (180).

Where it is used. Fertilizer bags list percent nitrogen, phosphorus and potassium. Nutrition labels, ore assays and drug purity tests rest on percent composition. Forensic labs identify unknowns with combustion analysis and mass spectrometry.

> **Safety first.** Adult handles the oven. The dish is hot; use mitts and let it cool on a trivet.

LAB

Water in a Hydrate

You will heat Epsom salt to drive off its water of crystallization and find the formula of the hydrate.

You will need

- Epsom salt ($MgSO_4 \cdot xH_2O$), an oven-safe dish, an oven or stovetop, a 0.1 g scale, oven mitts

Steps

1. Weigh the empty dish. Add about 10 g of Epsom salt and weigh again.
2. With an adult, bake at about 200 °C (400 °F) for 45 minutes.
3. Cool in a dry place and weigh. Heat 10 more minutes and reweigh until the mass is constant.
4. Calculate moles of $MgSO_4$ (120.37 g/mol) and of water lost; find x.

Check your work

- ☐ What value of x did you get? The accepted value is 7.
- ☐ Why must you heat to constant mass?

LAB NOTEBOOK

1. Masses: dish, dish + hydrate, dish + anhydrous salt after each heating.
2. Calculation of x and percent error.

> **From history.** Liebig also invented the beef extract cube and studied plant nutrition, showing that crops need nitrogen, phosphorus and potassium, the basis of the fertilizer industry.

WORKED EXAMPLES

Example 1. Find the molar mass of calcium carbonate, $\mathrm{CaCO_3}$, and the moles in 25.0 g.

▼ **Show solution**

1. 40.08 + 12.01 + 3(16.00) = 100.09 g/mol
2. $25.0 / 100.09 = 0.250$ mol

Answer: 100.09 g/mol; 0.250 mol

Example 2. A compound is 40.0% C, 6.7% H and 53.3% O with molar mass 180 g/mol. Find the empirical and molecular formulas.

▼ **Show solution**

1. Moles: C 40.0/12.01 = 3.33; H 6.7/1.01 = 6.6; O 53.3/16.00 = 3.33.
2. Divide by 3.33: C 1, H 2, O 1 → $\mathrm{CH_2O}$ (30.03).
3. 180/30.03 = 6.

Answer: $\mathrm{CH_2O}$; $\mathrm{C_6H_{12}O_6}$

PRACTICE

1. Easy What is the molar mass of $\mathrm{O_2}$?

   ▼ **Show answer**

   32.00 g/mol

2. Easy Find the molar mass of $\mathrm{Fe_2O_3}$.

   ▼ **Show answer**

   159.70 g/mol

3. Medium How many grams are in 0.500 mol of NaCl?

   ▼ **Show answer**

   29.2 g (58.44 × 0.500)

4. Medium What is the percent by mass of O in $Ca(OH)_2$?

▼ **Show answer**

43.2%

5. Hard A compound is 85.6% C, 14.4% H by mass. Find its empirical formula.

▼ **Show answer**

CH_2

COMMON MISTAKES

- Using atomic number instead of atomic mass.
- Forgetting to multiply by subscripts (and parentheses) when finding molar mass.
- Rounding 1.5 to 2 instead of doubling all ratios.

REFERENCE GUIDE

An eight-page, color-coded Organic Chemistry guide: carbon bonding and hybridization, structural formulas, functional groups, naming, isomers and chirality, acids and bases with pKa values, reaction types and mechanisms, spectroscopy (IR and NMR), and the molecules of life. The booklet prints on 11×17 paper and folds to letter size. Prints in color and reads clearly in black and white.

> **Make it a game · Molar Mass Relay.** Write ten formulas on cards. The student calculates molar masses as fast as possible with the periodic table; the parent checks with a calculator.

> **Scripture.** “For precept must be upon precept, precept upon precept; line upon line, line upon line; here a little, and there a little:” (Isaiah 28:10, KJV)
>
> Precept upon precept, line upon line: an empirical formula is built step by step from careful measurements.

6.3 Stoichiometry: Chemistry's Recipe Math

OBJECTIVES

- Use mole ratios from balanced equations.
- Solve mass-to-mass stoichiometry problems with dimensional analysis.
- Solve problems involving particles and moles.

> **Warm-up.** A sandwich takes 2 slices of bread and 1 slice of cheese. How many slices of bread for 15 sandwiches? How many sandwiches from 40 slices of bread and plenty of cheese?

WATCH

Stoichiometry: Converting Grams to Grams

Melissa Maribel · 5:33

https://www.youtube.com/watch?v=OkWV9d7Z6VA

> **WHEN TO WATCH**
> After the mole ratio idea.

BACKUP VIDEO

Stoichiometry grams to grams problem

Melissa Maribel · 3:05

https://www.youtube.com/watch?v=pi9gYPbjcR8

After the mass-to-mass example.

KEY TERMS

stoichiometry
Calculation of quantities of reactants and products from a balanced equation.

mole ratio
A ratio of coefficients from a balanced equation used as a conversion factor.

mass-to-mass problem

A problem converting grams of one substance to grams of another through moles.

dimensional analysis

Setting up unit factors so that unwanted units cancel.

excess

More of a reactant than is needed to react completely.

molar volume

The volume of one mole of gas: 22.4 L at STP.

CORE IDEAS

The story. The word stoichiometry (from Greek *stoicheion*, element, and *metron*, measure) was coined by Jeremias Benjamin Richter in 1792. Richter measured how much of each acid neutralized a given amount of base and found fixed "equivalent" amounts, evidence that elements combine in definite proportions long before Dalton explained why.

How they figured it out. The coefficients of a balanced equation give ratios of particles, and therefore of moles. In $2H_2 + O_2 \rightarrow 2H_2O$, 2 mol of hydrogen make 2 mol of water. The road map: grams A → (÷ molar mass) → moles A → (× mole ratio) → moles B → (× molar mass) → grams B. You cannot go from grams of A to grams of B directly.

The chemistry. Set up the problem as one chain of factors so the units cancel: $\text{g A} \times \frac{1\ \text{mol A}}{\text{g A}} \times \frac{\text{mol B}}{\text{mol A}} \times \frac{\text{g B}}{1\ \text{mol B}}$. Round only at the end to the correct significant figures.

Where it is used. Rocket engineers calculate the oxygen needed to burn fuel; pharmacists the reagents for a batch of medicine; bakers (unknowingly) the baking soda to neutralize buttermilk; and the space shuttle's designers the lithium hydroxide needed to remove the astronauts' CO_2.

LAB

How Much Gas from an Antacid?

You will predict and then measure the mass of carbon dioxide released by baking soda and vinegar.

You will need

- Baking soda, vinegar (excess), a cup, a 0.1 g scale

Steps

1. Weigh 2.00 g of baking soda. Predict the mass of CO_2 from $NaHCO_3 + HC_2H_3O_2 \rightarrow NaC_2H_3O_2 + H_2O + CO_2$.
2. Weigh a cup with 50 mL of vinegar.
3. Add the baking soda slowly to avoid splashing. Stir until fizzing stops.
4. Weigh again. The loss is the CO_2 that escaped.

Check your work

- ☐ Compare the measured loss with your prediction.
- ☐ Why might the measured value be low? (CO_2 stays dissolved.)

LAB NOTEBOOK

1. Your stoichiometry setup for the prediction.
2. Measured mass loss and percent difference.

> **From history.** Lithium hydroxide canisters scrubbed CO_2 aboard Apollo spacecraft. On Apollo 13 in 1970 the square command-module canisters did not fit the lunar module's round sockets, and engineers on the ground designed an adapter from plastic bags, cardboard and tape that saved the crew.

WORKED EXAMPLES

Example 1. How many grams of water form when 8.00 g of hydrogen burns?
$2H_2 + O_2 \rightarrow 2H_2O$

▼ **Show solution**

1. $8.00\ \text{g}\ H_2 \div 2.02 = 3.96\ \text{mol}$
2. Ratio 2:2 → $3.96\ \text{mol}\ H_2O$
3. $3.96 \times 18.02 = 71.4\ \text{g}$

Answer: 71.4 g

Example 2. How many grams of oxygen are needed to burn 100. g of propane?
$C_3H_8 + 5O_2 \rightarrow 3CO_2 + 4H_2O$

▼ **Show solution**

1. $100./44.11 = 2.267\ \text{mol}$ propane
2. × 5 = 11.34 mol O_2
3. × 32.00 = 363 g

Answer: 363 g

PRACTICE

1. Easy In $N_2 + 3H_2 \rightarrow 2NH_3$, what is the mole ratio of H_2 to NH_3?

   ▼ **Show answer**

   3:2

2. Medium Why can't you convert grams of A directly to grams of B?

   ▼ **Show answer**

   Coefficients give mole ratios, not mass ratios; you must go through moles.

3. Medium For $N_2 + 3H_2 \rightarrow 2NH_3$, what mass of NH_3 forms from 7.4 g of N_2?

   ▼ **Show answer**

   9.00 g

4. **Medium** For $CH_4 + 2O_2 \rightarrow CO_2 + 2H_2O$, what mass of CO_2 forms from 45.9 g of CH_4?

   ▼ **Show answer**

   126 g

5. **Medium** For $4Fe + 3O_2 \rightarrow 2Fe_2O_3$, what mass of Fe_2O_3 forms from 56.8 g of Fe?

   ▼ **Show answer**

   81.2 g

COMMON MISTAKES

- Using the mole ratio with grams instead of moles.
- Flipping the mole ratio (wanted substance goes on top).
- Using an unbalanced equation.

> **Make it a game · Road Map Race.** Draw the grams → moles → moles → grams road map on paper. The parent calls out a starting point and destination; the student names each factor needed.

> **Scripture.** "For which of you, intending to build a tower, sitteth not down first, and counteth the cost, whether he have sufficient to finish it?" (Luke 14:28, KJV)
>
> Jesus spoke of counting the cost before building. Stoichiometry counts the cost in reactants before the reaction.

UNIT 6 · THE MOLE AND STOICHIOMETRY · LESSON 6.4 OF 36 · 4 SESSIONS OF 45 MINUTES

6.4 Limiting Reactant and Percent Yield: Haber and Bosch (1909–1913)

OBJECTIVES

- Identify the limiting and excess reactants.
- Calculate theoretical yield, actual yield and percent yield.
- Explain the importance of the Haber–Bosch process.

> **Warm-up.** You have 10 hot-dog buns and 8 hot dogs. How many complete hot dogs can you make? What runs out first? What is left over?

WATCH

How to Find Limiting Reactants

Melissa Maribel · 8:52

https://www.youtube.com/watch?v=ymCZ2ShhBAw

> **WHEN TO WATCH**
> After the story.

BACKUP VIDEO

GCSE Chemistry - The Haber Process Explained

Cognito · 5:37

https://www.youtube.com/watch?v=1_HoWz5Kxfk

After the worked example.

KEY TERMS

limiting reactant
The reactant used up first, which determines the amount of product.

excess reactant
The reactant left over when the reaction stops.

theoretical yield
The maximum amount of product calculated from the limiting reactant.

actual yield
The amount of product actually obtained.

percent yield
(Actual yield ÷ theoretical yield) × 100.

catalyst
A substance that speeds a reaction without being used up.

CORE IDEAS

The story. By 1898 the British scientist William Crookes warned that the world would run out of fixed nitrogen for fertilizer; guano and Chilean saltpeter were being

mined out. Air is 78% nitrogen, but its triple bond makes it useless to most plants. In 1909 Fritz Haber at Karlsruhe made ammonia from nitrogen and hydrogen over an osmium catalyst at high pressure. Carl Bosch at the BASF company spent four years building steel reactors that could survive 200 atmospheres and 500 °C, found a cheap iron catalyst with Alwin Mittasch, and opened the first plant at Oppau in 1913.

How they figured it out. The reaction $\mathrm{N_2 + 3H_2 \rightleftharpoons 2NH_3}$ never goes to completion; only about 15% converts in each pass. Bosch's engineers cooled the gas to condense the ammonia and sent the unreacted gases back through, raising the overall yield. They chose ratios carefully so the expensive hydrogen would not be wasted.

The chemistry. To find the limiting reactant, calculate how much product each reactant could make; the one that makes less is limiting, and that amount is the theoretical yield. Percent yield = actual ÷ theoretical × 100. Yields fall short because of side reactions, incomplete reactions, losses in transfer and impure reactants.

Where it is used. Haber–Bosch ammonia now makes fertilizer that grows food for roughly half the people on earth. Its history also carries a moral lesson: Haber led Germany's poison gas program in World War I, and his wife, the chemist Clara Immerwahr, took her own life soon after its first use. Bosch, by contrast, later opposed the Nazi regime. Knowledge is powerful; how it is used is a moral choice.

LAB

Limiting Reactant with Balloons

You will vary the amount of baking soda with a fixed amount of vinegar and see where the balloon stops growing.

You will need

- Five identical balloons, five bottles, a funnel, vinegar, baking soda, a tape measure

Steps

1. Put 50 mL of vinegar in each bottle.
2. Using the funnel, put 1, 2, 3, 4 and 5 teaspoons of baking soda in the five balloons.
3. Stretch each balloon over a bottle neck and lift to dump the soda in. Swirl.
4. When fizzing stops, measure each balloon's circumference.

Check your work

- ☐ At what amount did the balloons stop getting bigger?
- ☐ Which reactant was limiting in the first bottles? In the last?

LAB NOTEBOOK

1. A graph of balloon size versus teaspoons of baking soda.
2. An explanation of the point where the graph levels off.

> **From history.** Haber received the 1918 Nobel Prize for ammonia synthesis; Bosch shared the 1931 prize for high-pressure chemistry. Haber, of Jewish descent, was forced out of Germany in 1933 and died in exile the next year.

WORKED EXAMPLES

Example 1. 6.0 g of H_2 and 28.0 g of N_2 react: $N_2 + 3H_2 \rightarrow 2NH_3$. Find the limiting reactant and theoretical yield.

▼ **Show solution**

1. N_2: 28.0/28.02 = 1.00 mol → 2.00 mol NH_3.
2. H_2: 6.0/2.02 = 2.97 mol → 1.98 mol NH_3.
3. Hydrogen makes less, so it limits: 1.98 × 17.03 = 33.7 g.

Answer: H_2 limits; 33.7 g NH_3

Example 2. If 25.0 g of ammonia is collected, what is the percent yield?

▼ **Show solution**

1. $25.0/33.7 \times 100$

Answer: 74.2%

PRACTICE

1. Easy What is the limiting reactant?

   ▼ **Show answer**

   The reactant used up first, which limits the amount of product.

2. Easy Write the formula for percent yield.

   ▼ **Show answer**

   Actual ÷ theoretical × 100

3. Hard For 2Na + Cl_2 → 2NaCl, 11.4 g of Na reacts with 9.2 g of Cl_2. (a) Which is the limiting reactant? (b) What is the theoretical yield of NaCl? (c) If 12.7 g is actually collected, what is the percent yield?

   ▼ **Show answer**

   (a) Cl_2; (b) 15.2 g; (c) 83.7%

4. Hard For $2Na + Cl_2 \rightarrow 2NaCl$, 17.8 g of Na reacts with 53.4 g of Cl_2. (a) Which is the limiting reactant? (b) What is the theoretical yield of NaCl? (c) If 35.7 g is actually collected, what is the percent yield?

   ▼ **Show answer**

   (a) Na; (b) 45.2 g; (c) 78.9%

5. Hard For $NaHCO_3 + HCl \rightarrow NaCl + H_2O + CO_2$, 23.6 g of $NaHCO_3$ reacts with 43.5 g of HCl. (a) Which is the limiting reactant? (b) What is the theoretical yield of CO_2? (c) If 10.0 g is actually collected, what is the percent yield?

   ▼ **Show answer**

   (a) $NaHCO_3$; (b) 12.4 g; (c) 80.9%

COMMON MISTAKES

- Choosing the limiting reactant by smaller mass instead of by product formed.
- Using the excess reactant to calculate yield.
- Reporting percent yield over 100% without questioning it (wet product or impurities).

> **Make it a game · Sandwich Shop.** Give the student piles of "bread" and "cheese" cards with a recipe (2 bread + 1 cheese). Time how fast they name the limiting ingredient and the number of sandwiches for five setups.

> **Scripture.** "Thou visitest the earth, and waterest it: thou greatly enrichest it with the river of God, which is full of water: thou preparest them corn, when thou hast so provided for it." (Psalm 65:9, KJV)
>
> God enriches the earth and prepares corn. Fertilizer from the air is one way He has let people share in that care.

UNIT 7

Gases

Gases are the easiest state of matter to describe with simple laws, and the history of those laws is the history of the experimental method itself. Students follow

Torricelli's mercury barometer, Boyle's J-tube, the balloonists Charles and Gay-Lussac, and finally the ideal gas law that ties pressure, volume, temperature and moles together. The unit ends with the kinetic-molecular theory, which explains every gas law as the motion of tiny particles, and with Dalton's partial pressures and Graham's diffusion.

The Physics course (Unit 8) cross-links to this unit for the ideal gas law.

PREREQUISITES

- Unit 6: moles and molar mass
- Algebra: solving proportions

MASTERY CHECKLIST

- ☐ I can convert among pressure units and explain how a barometer works.
- ☐ I can use Boyle's, Charles's and Gay-Lussac's laws, always in kelvins.
- ☐ I can use the ideal gas law and molar volume at STP.
- ☐ I can explain the gas laws with kinetic-molecular theory and use Dalton's and Graham's laws.

LESSONS IN THIS UNIT

7.1 Pressure: Torricelli and Boyle's Law (1662)	4 sessions of 45 minutes
7.2 Charles and Gay-Lussac: Balloons and Temperature (1783–1804)	4 sessions of 45 minutes
7.3 The Ideal Gas Law and Molar Volume	4 sessions of 45 minutes
7.4 Kinetic-Molecular Theory, Dalton's Partial Pressures and Graham's Diffusion	4 sessions of 45 minutes

Unit 7 test

Show your work. For every calculation write the equation, substitute with units and give the answer with its unit and sensible significant figures. The test is worth 100 points.

1. **7 pts** **Lesson 7.1**

   What happened to the water in the warm-up glass, and why?

2. **7 pts** **Lesson 7.1**

   What is the SI unit of pressure?

3. **7 pts** **Lesson 7.1**

   Convert 2.32 atm to kPa and to mmHg (1 atm = 101.3 kPa = 760 mmHg).

4. **6 pts** **Lesson 7.2**

   Write the combined gas law.

5. **6 pts** **Lesson 7.2**

   What happened to the balloon on the bottle in hot water?

6. **6 pts** **Lesson 7.2**

   A sealed can of gas is at 188 kPa and 25 °C. What is its pressure at 80 °C?

7. **6 pts** **Lesson 7.3**

   Why must the lighter be dried before reweighing?

8. **6 pts** **Lesson 7.3**

   What is the value of R in L·atm/(mol·K)?

9. **6 pts** **Lesson 7.3**

   What volume does 2.0 mol of gas occupy at 37 °C and 2.95 atm? ($R = 0.08206$ L·atm/(mol·K))

10. **6 pts** **Lesson 7.4**

    What words did Maxwell carve over the Cavendish Laboratory door?

11. **6 pts** **Lesson 7.4**

Did the food coloring spread faster in hot or cold water?

12. **6 pts** **Lesson 7.4**

A mixture contains nitrogen at 261 mmHg, oxygen at 107 mmHg and water vapor at 8 mmHg. What is the total pressure?

LAB AND DATA

13. **10 pts**

A student traps air in a syringe: 20.0 mL at 1.00 atm. Pressing the plunger, she reads these volumes at 1.25, 1.67 and 2.50 atm: 16.0, 12.0 and 8.0 mL. (a) Show that the data follow Boyle's law. (b) Predict the volume at 2.00 atm. (c) What two quantities must stay constant?

SHORT ESSAY

14. **15 pts**

Explain, using the kinetic-molecular theory, why a sealed bag of chips puffs up when driven from the coast up to a mountain pass, and why a car's tires look low on a cold morning.

ANSWER KEY

1. It stayed in; air pressure pushing up on the card exceeded the water's weight.
2. The pascal (N/m^2).
3. 235 kPa, 1.76×10^3 mmHg
   $2.32 \times 101.3 = 235 \cdot 2.32 \times 760 = 1.76 \times 10^3$
4. $P_1V_1/T_1 = P_2V_2/T_2$
5. It inflated because the air inside expanded.
6. 223 kPa
   Kelvin: 298 K and 353 K · $P_2 = P_1T_2/T_1 = (188)(353)/(298) = 223$ kPa
7. Water on it would add mass and make the butane mass too small.
8. 0.08206
9. 17.2 L
   $V = nRT/P = (2.0)(0.08206)(310)/(2.95) = 17.2$ L
10. Psalm 111:2: "The works of the LORD are great, sought out of all them that have pleasure therein."
11. Hot water.
12. 376 mmHg
    Dalton: total pressure = sum of partial pressures.
13. (a) PV = 20.0, 20.0, 20.0, 20.0 atm·mL: constant. (b) 20.0/2.00 = 10.0 mL. (c) Temperature and amount of gas.
14. Look for: gas particles in constant random motion; pressure from collisions with walls; on the mountain the outside air pressure is lower, so the gas inside expands until pressures balance (Boyle); in cold, particles move slower, hitting the walls less often and less hard, so pressure (and tire volume) drops (Gay-Lussac/Charles); temperature in kelvins proportional to average kinetic energy.

Printable answer key (PDF, for the parent)

UNIT 7 · GASES · LESSON 7.1 OF 36 · 4 SESSIONS OF 45 MINUTES

7.1 Pressure: Torricelli and Boyle's Law (1662)

OBJECTIVES

- Define pressure and convert among atm, kPa and mmHg.

- Explain how Torricelli's barometer measures air pressure.
- Use Boyle's law, $P_1V_1 = P_2V_2$.

> **Warm-up.** Fill a glass to the brim with water, cover it with an index card, and turn it upside down over a sink. Let go of the card. What holds the water in?

WATCH

Boyle's Law

Tyler DeWitt · 5:42

https://www.youtube.com/watch?v=ZoGtVVu3ymQ

> **WHEN TO WATCH**
> After the warm-up.

BACKUP VIDEO

Atmospheric Pressure

Tyler DeWitt · 4:23

https://www.youtube.com/watch?v=YU4ZLSH5m1Q

After Boyle's law.

KEY TERMS

pressure

Force per unit area; for gases, from particles striking surfaces.

barometer

An instrument that measures atmospheric pressure.

atmosphere (atm)

A pressure unit equal to average air pressure at sea level: 101.3 kPa or 760 mmHg.

pascal

The SI unit of pressure, one newton per square meter.

Boyle's law

At constant temperature, the pressure and volume of a gas are inversely proportional.

vacuum

A space with almost no matter in it.

CORE IDEAS

The story. Well-diggers in Florence found that suction pumps could not lift water more than about 10 meters. In 1643 Evangelista Torricelli, Galileo's student, filled a glass tube with mercury, which is 13.6 times denser than water, and turned it upside down in a dish. The mercury fell to about 76 cm and stopped, leaving empty space above. Torricelli reasoned that the weight of the air pressing on the dish held the column up: "We live submerged at the bottom of an ocean of air." Blaise Pascal had his brother-in-law carry a barometer up the Puy de Dôme in 1648; the mercury fell as he climbed.

How they figured it out. Robert Boyle and his assistant Robert Hooke used a J-shaped glass tube closed at the short end. Pouring mercury into the long arm trapped air in the short arm. Doubling the height of mercury (the pressure) halved the volume of air. Boyle published the table in 1662: pressure times volume stayed constant. Edme Mariotte found the same law in France in 1676.

The chemistry. $P_1V_1 = P_2V_2$ at constant temperature and amount of gas. Units: 1 atm = 101.3 kPa = 760 mmHg = 760 torr = 14.7 psi. Squeezing a gas into half the space means particles strike the walls twice as often, doubling the pressure.

Where it is used. Breathing is Boyle's law: the diaphragm enlarges the chest, lowering pressure in the lungs so air flows in. Syringes, scuba tanks, weather barometers and the bends in divers (gas expanding as pressure drops) all follow it.

LAB

Boyle’s Law with a Syringe and Books

You will add books on top of a sealed syringe and measure the volume of trapped air.

You will need

- A 60 mL plastic syringe with a cap (or hot glue the tip shut, adult)
- Several identical hardcover books, a board, a stand to hold the syringe upright

Steps

1. Pull the plunger to 50 mL and seal the tip.
2. Stand the syringe upright, plunger up, with a board balanced on top.
3. Add books one at a time, recording the volume each time.
4. Plot volume versus number of books, and 1/volume versus books.

Check your work

- ☐ Which graph is a straight line?
- ☐ What does this show about pressure and volume?

LAB NOTEBOOK

1. Data table and both graphs.
2. A sentence stating Boyle’s law in your own words.

KEY FORMULAS

$$P = \frac{F}{A} \qquad P_1 V_1 = P_2 V_2$$

$$1\ \text{atm} = 101.3\ \text{kPa} = 760\ \text{mmHg}$$

> **From history.** Boyle called his vacuum pump the “pneumatical engine.” Hooke built it; the two used it to show that sound does not travel in a vacuum and that a candle and a mouse both need air.

WORKED EXAMPLES

Example 1. A diver's lungs hold 6.0 L of air at 3.0 atm at depth. If she rose to the surface (1.0 atm) holding her breath, what volume would the air try to occupy?

▼ **Show solution**

1. $V_2 = P_1V_1/P_2 = (3.0)(6.0)/1.0$

Answer: 18 L, three times her lung volume, which is why divers must exhale while rising

Example 2. Convert 745 mmHg to atm and kPa.

▼ **Show solution**

1. $745/760 = 0.980$ atm
2. $0.980 \times 101.3 = 99.3$ kPa

Answer: 0.980 atm; 99.3 kPa

PRACTICE

1. Easy What holds the mercury up in a barometer?

   ▼ **Show answer**

   The pressure of the atmosphere on the mercury in the dish.

2. Easy What is standard atmospheric pressure in mmHg?

   ▼ **Show answer**

   760 mmHg

3. Easy Convert 0.89 atm to kPa and to mmHg (1 atm = 101.3 kPa = 760 mmHg).

   ▼ **Show answer**

   $90.2\ \text{kPa},\ 676\ \text{mmHg}$

4. Easy Convert 3.21 atm to kPa and to mmHg (1 atm = 101.3 kPa = 760 mmHg).

   ▼ **Show answer**

   $325\ \text{kPa},\ 2.44 \times 10^3\ \text{mmHg}$

5. **Medium** A gas occupies 2.8 L at 760 mmHg. At constant temperature it is squeezed to 2.0 L. What is the new pressure?

▼ **Show answer**

1.06×10^{3} mmHg

COMMON MISTAKES

- Thinking P and V both increase together; they are inversely related.
- Mixing pressure units in one equation.
- Forgetting the law applies only at constant temperature.

> **Make it a game · Squeeze or Stretch.** The parent describes a change (deeper dive, higher mountain, pushing a plunger). The student says whether the gas volume gets bigger or smaller and by what factor.

> **Scripture.** "To make the weight for the winds; and he weigheth the waters by measure." (Job 28:25, KJV)
>
> God makes the weight for the winds. Torricelli discovered that air has weight long after Job's words were written.

UNIT 7 · GASES · LESSON 7.2 OF 36 · 4 SESSIONS OF 45 MINUTES

7.2 Charles and Gay-Lussac: Balloons and Temperature (1783–1804)

OBJECTIVES

- Use the Kelvin scale and explain absolute zero.
- Use Charles's law and Gay-Lussac's law.
- Combine the gas laws into the combined gas law.

> **Warm-up.** Stretch a balloon over an empty bottle. Set the bottle in a bowl of hot tap water, then in ice water. Describe what happens.

WATCH

Charles' Law

Tyler DeWitt · 6:42

https://www.youtube.com/watch?v=oIfFoiwRCVE

> **WHEN TO WATCH**
> After the balloon story.

BACKUP VIDEO

Rearranging Charles', Gay-Lussac's, and Avogadro's Laws

Tyler DeWitt · 5:45

https://www.youtube.com/watch?v=tV1OzhEaQkY

With the combined gas law.

KEY TERMS

Kelvin scale

A temperature scale starting at absolute zero; K = °C + 273.

absolute zero

0 K (−273.15 °C), the lowest possible temperature.

Charles’s law

At constant pressure, gas volume is directly proportional to kelvin temperature.

Gay-Lussac’s law

At constant volume, gas pressure is directly proportional to kelvin temperature.

combined gas law

$P_1V_1/T_1 = P_2V_2/T_2$.

direct proportion

A relationship in which one quantity doubles when the other doubles.

CORE IDEAS

The story. On November 21, 1783, two men rose over Paris in the Montgolfier brothers’ hot-air balloon. Ten days later the physicist Jacques Charles flew a hydrogen balloon over 3,000 meters. Charles found, but never published, that gases expand evenly as they warm. In 1802 Joseph Louis Gay-Lussac measured it carefully

and gave Charles credit. In 1804 Gay-Lussac himself rode a balloon to about 7,000 meters to sample the air and measure the magnetic field, a height record that stood for 50 years.

How they figured it out. Every gas expanded by the same fraction, about 1/273 of its volume at 0 °C for each degree of warming. Extending the line backward, the volume would reach zero at −273 °C. In 1848 William Thomson (Lord Kelvin) proposed an absolute scale starting there. On the Kelvin scale volume and temperature are directly proportional.

The chemistry. Charles: $V_1/T_1 = V_2/T_2$ (constant P). Gay-Lussac: $P_1/T_1 = P_2/T_2$ (constant V). Combined: $P_1V_1/T_1 = P_2V_2/T_2$. *Always* use kelvins: K = °C + 273. Doubling the temperature from 10 °C to 20 °C does not double the volume, because 283 K to 293 K is only a 3.5% increase.

Where it is used. Hot-air balloons rise because heated air expands and becomes less dense. Tire pressure drops in cold weather. Aerosol cans warn against heating because pressure rises with temperature. Turkey pop-up timers and the whistle of a pressure cooker follow the same laws.

> **Safety first.** Adult handles hot water. Keep water below boiling; use tongs.

LAB

Charles’s Law in a Pot of Water

You will measure the volume of trapped air at several temperatures and extrapolate to absolute zero.

You will need

- A 10 mL or 60 mL syringe sealed at the tip, a pot, a kitchen thermometer, ice
- Stove (adult)

Steps

1. Pull the plunger to about half and seal. Note the volume at room temperature.
2. Submerge the syringe in ice water, then cool water, warm water and hot (not boiling) water, holding each for 3 minutes. Tap the plunger gently and record volume and temperature.
3. Graph volume versus °C and draw the best line.
4. Extend the line to zero volume and read the temperature.

Check your work

- ☐ How close was your intercept to −273 °C?
- ☐ Why does friction in the syringe cause error?

LAB NOTEBOOK

1. Data table and graph with extrapolation.
2. Percent error from −273 °C.

KEY FORMULAS

$$\frac{V_1}{T_1} = \frac{V_2}{T_2} \qquad \frac{P_1}{T_1} = \frac{P_2}{T_2} \qquad \frac{P_1 V_1}{T_1} = \frac{P_2 V_2}{T_2}$$

> **From history.** Gay-Lussac also found that gases combine in simple volume ratios (1808), the clue that led Avogadro to his hypothesis. He discovered boron and named iodine.

WORKED EXAMPLES

Example 1. A balloon holds 2.00 L at 25 °C. It is put in a freezer at −18 °C. Find its new volume.

▼ **Show solution**

1. $T_1 = 298$ K, $T_2 = 255$ K
2. $V_2 = 2.00 \times 255/298$

Answer: 1.71 L

Example 2. A gas at 1.00 atm, 3.00 L and 27 °C is compressed to 1.50 L and heated to 127 °C. Find the new pressure.

▼ **Show solution**

1. $T_1 = 300$ K, $T_2 = 400$ K
2. $P_2 = P_1V_1T_2/(V_2T_1) = (1.00)(3.00)(400)/((1.50)(300))$

Answer: 2.67 atm

PRACTICE

1. Easy Convert 25 °C to kelvins.

   ▼ **Show answer**

   298 K

2. Easy What is absolute zero in °C?

   ▼ **Show answer**

   −273 °C (−273.15)

3. Medium A sealed can of gas is at 177 kPa and 0 °C. What is its pressure at 100 °C?

   ▼ **Show answer**

   242 kPa

4. **Medium** A balloon holds 3.5 L of air at 20 °C. At constant pressure, what is its volume at 100 °C?

   ▼ **Show answer**

   4.46 L

5. **Medium** A sealed can of gas is at 169 kPa and 37 °C. What is its pressure at 60 °C?

   ▼ **Show answer**

   182 kPa

COMMON MISTAKES

- Using Celsius in the gas laws.
- Thinking volume drops to zero at 0 °C.
- Mixing up which variable is constant in each law.

> **Make it a game · Kelvin Quick-Draw.** The parent calls out Celsius temperatures (body, boiling, freezing, a summer day, liquid nitrogen at −196 °C). The student converts to kelvins as fast as possible.

> **Scripture.** “He giveth snow like wool: he scattereth the hoarfrost like ashes. He casteth forth his ice like morsels: who can stand before his cold?” (Psalm 147:16–17, KJV)
>
> Snow, frost and ice obey the cold God sends. Gas volumes shrink in the cold by a law He set.

UNIT 7 · GASES · LESSON 7.3 OF 36 · 4 SESSIONS OF 45 MINUTES

7.3 The Ideal Gas Law and Molar Volume

OBJECTIVES

- Use the ideal gas law $PV = nRT$.
- Use molar volume at STP (22.4 L/mol).
- Find the molar mass or density of a gas.

> **Warm-up.** A party balloon and a basketball both contain air. What four things would you need to know to say exactly how much air each holds?

WATCH

Ideal Gas Law Introduction

Tyler DeWitt · 6:18

https://www.youtube.com/watch?v=WhP6zJbSxec

> **WHEN TO WATCH**
> After the derivation of PV = nRT.

BACKUP VIDEO

Ideal Gas Law Practice Problems

Tyler DeWitt · 10:53

https://www.youtube.com/watch?v=TqLlfHBFY08

With gas stoichiometry.

KEY TERMS

ideal gas law

$PV = nRT$, relating pressure, volume, moles and temperature.

gas constant (R)

0.08206 L·atm/(mol·K), or 8.314 J/(mol·K).

STP

Standard temperature and pressure: 0 °C (273 K) and 1 atm.

molar volume

The volume of one mole of an ideal gas: 22.4 L at STP.

ideal gas

A model gas whose particles have no volume and no attractions.

real gas

An actual gas, which deviates from ideal behavior at high pressure and low temperature.

CORE IDEAS

The story. In 1834 the French engineer Émile Clapeyron combined Boyle's, Charles's and Avogadro's laws into a single equation. Avogadro's law says volume is proportional to the number of moles. Put together: PV is proportional to nT, and the constant of proportionality is R, the same for every gas.

How they figured it out. Measurements of many gases at low pressure showed the same R. At high pressure or near condensation, real gases deviate because their particles take up space and attract one another; Johannes van der Waals corrected the equation for this in 1873 and won the Nobel Prize in 1910.

The chemistry. $PV = nRT$ with $R = 0.08206$ L·atm/(mol·K): use L, atm, mol and K. At STP one mole of any ideal gas occupies 22.4 L, about the volume of three basketballs. Since $n = m/M$, the law can find molar mass ($M = mRT/PV$) or density ($d = PM/RT$). Gas stoichiometry uses moles from $PV = nRT$ in the usual road map.

Where it is used. Airbag designers calculate the moles of nitrogen needed to fill a bag; weather balloons are filled to the right volume to reach altitude; anesthesiologists dose gases by partial pressure; engineers size gas tanks.

> **Safety first.** Outdoors or very well ventilated. No flames anywhere nearby; butane is highly flammable. Adult only presses the lighter. Never strike the spark wheel during the lab.

LAB

Molar Mass of Butane from a Lighter

You will collect butane gas from a lighter over water and calculate its molar mass. (Adult does all handling of the lighter.)

You will need

- A disposable butane lighter, a large bowl of water, a 250 mL graduated cylinder or measuring cup
- A 0.01 g scale, a thermometer, a barometer reading (weather website)

Steps

1. Weigh the lighter dry.
2. Fill the cylinder with water and invert it in the bowl.
3. Hold the lighter underwater beneath the cylinder and press the release (no spark) until about 200 mL of gas collects.
4. Dry the lighter thoroughly and reweigh. Record the gas volume, water temperature and pressure.
5. Calculate moles with $PV = nRT$ (subtract water vapor pressure, about 20 mmHg at 22 °C) and divide mass by moles.

Check your work

☐ Compare your molar mass with butane, C_4H_{10}, 58.1 g/mol.

☐ Why must the lighter be completely dry before reweighing?

LAB NOTEBOOK

1. All measurements and the calculation.
2. Percent error and two sources of error.

KEY FORMULAS

$$PV = nRT \qquad R = 0.08206\ \frac{\text{L}\cdot\text{atm}}{\text{mol}\cdot\text{K}}$$

$$M = \frac{mRT}{PV} \qquad d = \frac{PM}{RT}$$

> **From history.** Van der Waals was a schoolteacher who studied for his doctorate in his spare time. His 1873 thesis on real gases so impressed James Clerk Maxwell that Maxwell learned Dutch to read it.

WORKED EXAMPLES

Example 1. How many moles of air are in a 2.50 L balloon at 1.00 atm and 22 °C?

▼ **Show solution**

1. $T = 295\ \text{K}$
2. $n = PV/RT = (1.00)(2.50)/((0.08206)(295))$

Answer: 0.103 mol

Example 2. A 1.00 L flask holds 1.25 g of an unknown gas at STP. Find its molar mass and identify it.

▼ **Show solution**

1. $n = 1.00/22.4 = 0.0446\ \text{mol}$
2. $M = 1.25/0.0446 = 28.0\ \text{g/mol}$

Answer: 28.0 g/mol, likely nitrogen (N_2)

PRACTICE

1. Easy What does each letter in $PV = nRT$ stand for?

   ▼ **Show answer**

   Pressure, volume, moles, gas constant, kelvin temperature.

2. Easy What is the volume of 1 mol of gas at STP?

▼ **Show answer**

22.4 L

3. Easy What volume does 2.75 mol of gas occupy at STP (0 °C, 1 atm; 22.4 L/mol)?

▼ **Show answer**

61.6 L

4. Medium What volume does 1.81 mol of gas occupy at 37 °C and 2.49 atm? ($R = 0.08206$ L·atm/(mol·K))

▼ **Show answer**

18.5 L

5. Hard An antacid tablet makes 1.95 L of carbon dioxide at 1.015 atm and 25 °C. How many moles and grams of CO_2 is that?

▼ **Show answer**

0.0809 mol, 3.56 g

COMMON MISTAKES

- Using °C or mL with R = 0.08206.
- Using 22.4 L/mol when not at STP.
- Forgetting to convert grams to moles.

> **Make it a game · Unit Police.** The parent writes a gas problem with mixed units (mL, °C, mmHg, kPa). The student converts every quantity to L, K, atm and mol before solving. Points for each correct conversion.

> **Scripture.** “The wind bloweth where it listeth, and thou hearest the sound thereof, but canst not tell whence it cometh, and whither it goeth: so is every one that is born of the Spirit.” (John 3:8, KJV)
>
> The wind blows where it lists, and Jesus compared it to the work of the Spirit. Even invisible gases follow laws.

7.4 Kinetic-Molecular Theory, Dalton's Partial Pressures and Graham's Diffusion

OBJECTIVES

- State the assumptions of the kinetic-molecular theory.
- Explain pressure and temperature in terms of particle motion.
- Use Dalton's law of partial pressures and Graham's law of effusion.

> **Warm-up.** Open a bottle of perfume or vanilla in one corner of a room. Time how long until someone across the room smells it. Why does it take so long if molecules move hundreds of meters per second?

WATCH

The Kinetic Molecular Theory of Gas (part 1)

Tyler DeWitt · 9:59

https://www.youtube.com/watch?v=fIMdIMACyN4

> **WHEN TO WATCH**
> After the assumptions.

BACKUP VIDEO

The Kinetic Molecular Theory of Gas (part 2)

Tyler DeWitt · 5:32

https://www.youtube.com/watch?v=apOSDqZd6Fg

With Graham's law.

KEY TERMS

kinetic-molecular theory
The model that gases are tiny particles in constant random motion with elastic collisions.

kinetic energy
Energy of motion, $\frac{1}{2}mv^2$.

partial pressure

The pressure one gas in a mixture would exert alone.

Dalton's law

The total pressure of a gas mixture equals the sum of the partial pressures.

diffusion

The spreading of particles from high to low concentration.

effusion

The escape of gas through a tiny hole; Graham's law: rate $\propto 1/\sqrt{M}$.

CORE IDEAS

The story. Daniel Bernoulli suggested in 1738 that gas pressure comes from countless tiny particles striking the walls, but the idea waited a century. James Clerk Maxwell (1860) and Ludwig Boltzmann worked out the mathematics, showing how particle speeds spread around an average. Maxwell, a devout Christian, wrote that the molecules bear the stamp of a manufactured article. Meanwhile, John Dalton (1801) found that each gas in a mixture behaves as if alone, and Thomas Graham (1833) measured that light gases escape through tiny holes faster than heavy ones.

How they figured it out. The kinetic-molecular theory assumes: (1) gas particles are tiny compared with the distances between them; (2) they move constantly and randomly in straight lines; (3) collisions are perfectly elastic; (4) there are no attractions between them; (5) the average kinetic energy is proportional to kelvin temperature. From these, all the gas laws follow: squeeze the gas and particles hit walls more often (Boyle); heat it and they hit harder and more often (Charles, Gay-Lussac).

The chemistry. At the same temperature, all gases have the same average kinetic energy, so lighter particles move faster: $\text{rate}_1/\text{rate}_2 = \sqrt{M_2/M_1}$. Hydrogen molecules average about 1,900 m/s at room temperature. Perfume spreads slowly because each molecule collides billions of times per second and wanders. Dalton: $P_{\text{total}} = P_1 + P_2 + \ldots$. Gas collected over water includes water vapor, which must be subtracted.

Where it is used. Divers breathe mixtures with reduced nitrogen partial pressure to avoid narcosis. Helium balloons shrink overnight because helium effuses through the rubber faster than air. Graham's law was used in the 1940s to separate uranium-235 from uranium-238 as uranium hexafluoride gas.

LAB

Diffusion Races

You will compare diffusion rates with food coloring in hot and cold water and with balloons of different gases.

You will need

- Two clear glasses, hot and ice water, food coloring, a timer
- Two balloons: one filled with helium (party store), one with air of the same size; a tape measure

Steps

1. Fill one glass with hot water and one with ice water. Let them settle. Add one drop of food coloring to each at the same moment. Time until each is evenly colored.
2. Measure both balloons' circumferences. Measure again every 6 hours for a day.
3. Graph circumference versus time.

Check your work

- ☐ Which glass mixed faster? Explain with kinetic energy.
- ☐ Which balloon shrank faster? Explain with Graham's law.

LAB NOTEBOOK

1. Times for both glasses.
2. Balloon graph with a Graham's-law explanation.

KEY FORMULAS

$$P_{\text{total}} = P_1 + P_2 + P_3 + \cdots \qquad \frac{r_1}{r_2} = \sqrt{\frac{M_2}{M_1}}$$

> **From history.** Maxwell is honored on the cover of every physics textbook through his equations of electromagnetism. At Cambridge he had Psalm 111:2 carved over the door of the Cavendish Laboratory.

WORKED EXAMPLES

Example 1. Air at 1.00 atm is 78% nitrogen, 21% oxygen and 1% argon by moles. Find the partial pressure of oxygen in mmHg.

▼ **Show solution**

1. $P_{O_2} = 0.21 \times 760$

Answer: 160 mmHg

Example 2. Ammonia (17.0 g/mol) and hydrogen chloride (36.5 g/mol) are released at opposite ends of a 1.00 m tube. Where do they meet and form a white ring?

▼ **Show solution**

1. Rate ratio $\sqrt{36.5/17.0} = 1.47$
2. Ammonia travels 1.47 times as far: $1.47/(1.47+1) = 0.594$

Answer: About 59 cm from the ammonia end

PRACTICE

1. Easy What is the average kinetic energy of gas particles proportional to?

   ▼ **Show answer**

   Kelvin temperature.

2. Easy State Dalton’s law of partial pressures.

   ▼ **Show answer**

   Total pressure equals the sum of the partial pressures.

3. Easy A mixture contains nitrogen at 180 mmHg, oxygen at 81 mmHg and water vapor at 9 mmHg. What is the total pressure?

   ▼ **Show answer**

   270 mmHg

4. Easy A mixture contains nitrogen at 165 mmHg, oxygen at 121 mmHg and water vapor at 38 mmHg. What is the total pressure?

   ▼ **Show answer**

   324 mmHg

5. Hard By Graham's law, how many times faster does He effuse than N_2?

   ▼ **Show answer**

   2.65 times

COMMON MISTAKES

- Thinking heavier gases have more kinetic energy at the same temperature (all have the same average).
- Forgetting to subtract water vapor pressure.
- Inverting the molar masses in Graham's law.

> **Make it a game · Molecule Bumper Cars.** In a hallway, family members walk randomly as gas particles in a box marked with tape. Shrink the box (Boyle) or speed up (heating) and count wall bumps per minute.

> **Scripture.** "The works of the LORD are great, sought out of all them that have pleasure therein." (Psalm 111:2, KJV)
>
> Maxwell placed these words over the Cavendish Laboratory door; they are carved over the new laboratory too.

UNIT 8

Solutions, Acids and Bases

Most of the chemistry in living things, oceans and kitchens happens in water. This unit studies how substances dissolve, how to measure concentration in molarity,

and why salt melts ice and makes homemade ice cream possible. It then turns to acids and bases: Arrhenius's ions, the broader Brønsted–Lowry definition, Sørensen's pH scale invented in a brewery laboratory, indicators from red cabbage, and the titration that measures the acid in vinegar to a fraction of a percent.

The red-cabbage indicator and the vinegar titration are two of the best home labs of the year and need no special equipment beyond a medicine syringe.

PREREQUISITES

- Unit 4: polarity and intermolecular forces
- Unit 6: moles and molar mass

MASTERY CHECKLIST

- ☐ I can explain dissolving with polarity and calculate molarity and dilutions.
- ☐ I can explain and calculate freezing-point depression and boiling-point elevation.
- ☐ I can identify acids and bases by the Arrhenius and Brønsted–Lowry definitions, including conjugate pairs.
- ☐ I can calculate pH and use a titration to find an unknown concentration.

LESSONS IN THIS UNIT

8.1	**Solutions, Solubility and Molarity**	4 sessions of 45 minutes
8.2	**Colligative Properties: Salt on Ice and Homemade Ice Cream**	4 sessions of 45 minutes
8.3	**Acids and Bases: Arrhenius (1884) to Brønsted and Lowry (1923)**	4 sessions of 45 minutes
8.4	**pH, Indicators and Titration: Sørensen at the Carlsberg Laboratory (1909)**	4 sessions of 45 minutes

Unit 8 test

Show your work. For every calculation write the equation, substitute with units and give the answer with its unit and sensible significant figures. The test is worth 100 points.

1. **7 pts** **Lesson 8.1**

   How does a sodium acetate hand warmer work?

2. **7 pts** **Lesson 8.1**

   Does sugar dissolve better in hot or cold water?

3. **7 pts** **Lesson 8.1**

   What mass of NaCl is needed for 431 mL of 1.53 M solution?

4. **7 pts** **Lesson 8.2**

   How does antifreeze protect a car in summer?

5. **7 pts** **Lesson 8.2**

   Does salt water boil at a higher or lower temperature than pure water?

6. **7 pts** **Lesson 8.2**

   39 g of $\mathrm{CaCl_2}$ is dissolved in 0.74 kg of water. Estimate the new freezing point. ($K_f = 1.86\ ^\circ\mathrm{C{\cdot}kg/mol}$; $i = 3$)

7. **7 pts** **Lesson 8.3**

   Write the conjugate base of $\mathrm{H_2PO_4^-}$ and its conjugate acid.

8. **7 pts** **Lesson 8.3**

   What did Davy show about Lavoisier’s idea that all acids contain oxygen?

9. **7 pts** **Lesson 8.4**

   What color is phenolphthalein in acid and in base?

10. **6 pts** **Lesson 8.4**

    What is a buffer?

11. **6 pts** **Lesson 8.4**

Find the pH of a solution with $[\mathrm{H}^+] = 10^{-1}$ M. Is it acidic, basic or neutral?

LAB AND DATA

12. **10 pts**

A student titrates 10.0 mL of vinegar with 0.500 M NaOH using phenolphthalein. The pink color first lasts at 16.8 mL. (a) Find the molarity of acetic acid. (b) Vinegar's density is about 1.00 g/mL; find the percent acetic acid by mass (60.05 g/mol). (c) Why add the base drop by drop near the end?

SHORT ESSAY

13. **15 pts**

Compare the Arrhenius and Brønsted–Lowry definitions of acids and bases. Use ammonia as an example of why the newer definition was needed, and explain what pH measures.

ANSWER KEY

1. A supersaturated solution crystallizes suddenly when triggered, releasing heat.
2. Hot water.
3. 38.5 g
   $n = MV = (1.53)(0.431) = 0.659$ mol · $0.659 \times 58.44 = 38.5$ g
4. It raises the boiling point of the coolant.
5. Higher.
6. -2.65 °C
   $m = (39/110.98)/0.74 = 0.475$ mol/kg · $\Delta T_f = iK_f m = (3)(1.86)(0.475) = 2.65$ °C · New freezing point = 0 − 2.65 °C
7. Conjugate base HPO_4^{2-}; conjugate acid H_3PO_4.
8. Hydrochloric acid contains no oxygen.
9. Colorless in acid; pink in base.
10. A solution that resists changes in pH.
11. pH = 1.00, acidic
    $\text{pH} = -\log(0.10) = 1.00$
12. (a) $0.500 \times 0.0168 = 0.00840$ mol; ÷ 0.0100 L = 0.840 M. (b) 0.00840 × 60.05 = 0.504 g in 10.0 g = 5.04%. (c) The endpoint is reached with one drop; adding quickly overshoots it.
13. Look for: Arrhenius (1884): acids produce H^+ in water, bases OH^-; Brønsted–Lowry (1923): acids donate protons, bases accept them, not limited to water or hydroxides; ammonia has no OH but accepts a proton from water to form NH_4^+ and OH^-; conjugate pairs; pH = −log[H^+], each unit a factor of 10; 7 neutral at 25 °C.

Printable answer key (PDF, for the parent)

UNIT 8 · SOLUTIONS, ACIDS AND BASES · LESSON 8.1 OF 36 · 4 SESSIONS OF 45 MINUTES

8.1 Solutions, Solubility and Molarity

OBJECTIVES

- Describe solute, solvent and the dissolving process.
- Use solubility curves and the terms saturated, unsaturated and supersaturated.

- Calculate molarity and dilutions.

> **Warm-up.** Stir sugar into iced tea and hot tea. Which dissolves more? Now shake a can of soda and a warm can of soda (over the sink, adult). Which fizzes more? What do temperature and dissolving have to do with each other?

WATCH

Molarity: A Deeper Understanding

Tyler DeWitt · 11:51

https://www.youtube.com/watch?v=mqRPLaOAoEU

> **WHEN TO WATCH**
> After the dissolving discussion.

BACKUP VIDEO

Temperature and Solubility: Solids and Gases

Tyler DeWitt · 10:56

https://www.youtube.com/watch?v=awMdCCidgOc

With molarity problems.

KEY TERMS

solution
A uniform mixture of a solute dissolved in a solvent.

solute
The substance dissolved.

solvent
The substance that does the dissolving, usually present in greater amount.

molarity
Moles of solute per liter of solution (M).

saturated
Holding the maximum solute that will dissolve at that temperature.

supersaturated
Holding more dissolved solute than is normally possible; unstable.

CORE IDEAS

The story. Water was called the universal solvent by the alchemists, and it nearly is. In the 1880s Svante Arrhenius, Jacobus van ’t Hoff and Wilhelm Ostwald founded physical chemistry by studying how dissolved substances behave. They found solutions follow laws much like gases.

How they figured it out. When salt dissolves, polar water molecules surround each ion, the negative oxygen ends toward $\mathrm{Na^+}$ and the positive hydrogen ends toward $\mathrm{Cl^-}$, pulling the crystal apart. Sugar dissolves because its OH groups hydrogen-bond with water. Oil does not because it is nonpolar: like dissolves like. Most solids dissolve better in hot water, but gases dissolve better in cold water, which is why warm soda goes flat and why fish suffer in warm ponds with little oxygen.

The chemistry. Molarity $M = n/V$ (moles per liter of *solution*). To make a 1.00 M NaCl solution, dissolve 58.44 g of salt and add water until the total volume is 1.00 L. Dilution keeps moles constant: $M_1V_1 = M_2V_2$. A solubility curve shows how many grams dissolve in 100 g of water at each temperature; points on the curve are saturated, below unsaturated, above supersaturated.

Where it is used. IV saline is 0.9% NaCl (0.154 M). Pharmacists dilute stock solutions; farmers mix sprays; rock candy grows from a supersaturated solution; hand warmers flash-crystallize supersaturated sodium acetate.

> **Safety first.** Adult handles boiling sugar syrup: it causes severe burns.

LAB

Rock Candy and a Solubility Curve

You will measure how much sugar dissolves at several temperatures and grow rock candy from the saturated solution.

You will need

- Sugar, a measuring cup, a kitchen scale, a saucepan, a thermometer
- Jars, wooden skewers, clothespins

Steps

1. With an adult, heat 100 g of water to 20, 40, 60 and 80 °C in turn, adding weighed sugar until no more dissolves. Record grams dissolved.
2. Graph grams per 100 g water versus temperature.
3. Make a hot saturated solution, pour into a jar, and hang a sugar-rolled skewer from a clothespin.
4. Observe crystal growth for a week.

Check your work

- ☐ How does sugar’s solubility change with temperature?
- ☐ Why does crystal growth happen as the solution cools?

LAB NOTEBOOK

1. Solubility data and curve.
2. Rock candy sketches over a week.

KEY FORMULAS

$$M = \frac{n}{V} \qquad M_1V_1 = M_2V_2$$

> **From history.** Arrhenius’s 1884 thesis on ions in solution was given the lowest passing grade at Uppsala; in 1903 it won him the Nobel Prize.

Another video. Crystallization - Chem Definition. A short extra on the same idea, after the core reading.



For parents: Periodic Videos, 4:46. A short definition of crystallization from the University of Nottingham chemists. Periodic Videos is an approved channel. Periodic Videos · 4:46

WORKED EXAMPLES

Example 1. Find the molarity of 11.7 g NaCl in 500. mL of solution.

▼ **Show solution**

1. $11.7/58.44 = 0.200$ mol
2. $0.200/0.500$

Answer: 0.400 M

Example 2. How would you make 250. mL of 0.100 M solution from 2.00 M stock?

▼ **Show solution**

1. $V_1 = (0.100)(250.)/2.00 = 12.5$ mL

Answer: Measure 12.5 mL of stock and add water to 250. mL

PRACTICE

1. Easy In salt water, which is the solute and which the solvent?

   ▼ **Show answer**

   Salt is the solute; water is the solvent.

2. Easy What is molarity?

   ▼ **Show answer**

   Moles of solute per liter of solution.

3. Medium 24.0 g of $CaCl_2$ is dissolved to make 419 mL of solution. Find the molarity.

   ▼ **Show answer**

   0.516 M

4. Medium What mass of $NaHCO_3$ is needed for 475 mL of 0.28 M solution?

   ▼ **Show answer**

   11.2 g

5. Hard How many milliliters of 1.1 M stock solution are needed to make 250 mL of 0.15 M solution?

   ▼ **Show answer**

   34.1 mL

COMMON MISTAKES

- Dividing by liters of solvent instead of solution.
- Forgetting to convert mL to L.
- Thinking gases become more soluble when warmed.

REFERENCE GUIDE: NURSING ENTRANCE EXAM REVIEW

An eight-page, color-coded review for the two nursing entrance exams, TEAS 7 and HESI A2: reading, math (numbers, algebra and ratios, measurement and conversions, geometry and data), science (anatomy and physiology, biology, chemistry, physics, scientific reasoning) and English (conventions, grammar, vocabulary), with a side-by-side look at both exams and practice questions. Not affiliated with or endorsed by ATI or Elsevier; TEAS and HESI are their owners'

trademarks. The booklet prints on 11×17 paper and folds to letter size. Prints in color and reads clearly in black and white.

> **Make it a game · Saturated or Not?.** The parent names an amount and temperature (“50 g sugar in 100 g water at 20 °C”). Using the class solubility curve, the student calls saturated, unsaturated or supersaturated.

> **Scripture.** “For as the rain cometh down, and the snow from heaven, and returneth not thither, but watereth the earth, and maketh it bring forth and bud, that it may give seed to the sower, and bread to the eater:” (Isaiah 55:10, KJV)
>
> Rain and snow water the earth, carrying dissolved minerals to plant roots.

UNIT 8 · SOLUTIONS, ACIDS AND BASES · LESSON 8.2 OF 36 · 4 SESSIONS OF 45 MINUTES

8.2 Colligative Properties: Salt on Ice and Homemade Ice Cream

OBJECTIVES

- Define colligative properties.
- Explain freezing-point depression and boiling-point elevation.
- Calculate molality and freezing-point changes, including the van ’t Hoff factor.

> **Warm-up.** Put an ice cube on a plate and sprinkle salt on half of it. After 5 minutes, which half is melting faster? Why do road crews salt icy roads?

WATCH

Boiling point elevation and freezing point depression

Khan Academy · 14:00

https://www.youtube.com/watch?v=z9LxdqYntlU

> **WHEN TO WATCH**
> After the warm-up.

BACKUP VIDEO

How Science Affects Your Ice Cream

Reactions · 4:25

https://www.youtube.com/watch?v=-rlapUkWCSM

With the calculations.

KEY TERMS

colligative property
A property depending on the number of dissolved particles, not their identity.

boiling-point elevation
The raising of a solvent's boiling point by a dissolved solute.

van 't Hoff factor
The number of particles a solute forms in solution (i; 2 for NaCl, 3 for $CaCl_2$).

freezing-point depression
The lowering of a solvent's freezing point by a dissolved solute.

molality
Moles of solute per kilogram of solvent (m).

osmotic pressure
The pressure needed to stop solvent from flowing across a membrane into a more concentrated solution.

CORE IDEAS

The story. François-Marie Raoult in Grenoble measured in the 1880s how hundreds of solutes lowered freezing points. He found the drop depends only on how many particles are dissolved. Van 't Hoff explained it, and salts that lowered the freezing

point twice as much as expected were part of Arrhenius's evidence that salts break into ions.

How they figured it out. Dissolved particles get in the way of solvent molecules locking into a crystal, so the solution must be colder to freeze. They also take places at the surface, slowing evaporation, so the solution must be hotter to boil. One mole of sugar gives one mole of particles; one mole of NaCl gives two; one mole of $\mathrm{CaCl_2}$ gives three.

The chemistry. $\Delta T_f = iK_f m$ and $\Delta T_b = iK_b m$, where m is molality (mol solute per kg solvent). For water $K_f = 1.86$ °C·kg/mol and $K_b = 0.512$ °C·kg/mol. Molality is used because it does not change with temperature. Osmosis, the flow of water across a membrane toward the more concentrated side, is another colligative effect (Biology Unit 3).

Where it is used. Salting roads; rock salt and ice in an ice cream maker reaching −10 °C; antifreeze (ethylene glycol) keeping radiators from freezing and boiling; fish in polar seas with antifreeze proteins; salt water taking slightly longer to boil.

LAB

Ice Cream in a Bag

You will make ice cream using freezing-point depression and measure how cold salted ice gets.

You will need

- 1 cup half-and-half, 2 tbsp sugar, 1/2 tsp vanilla
- A quart and a gallon zip-top bag, 4 cups ice, 1/2 cup rock salt, a thermometer, gloves or a towel

Steps

1. Mix the half-and-half, sugar and vanilla in the quart bag and seal it well.
2. Fill the gallon bag with ice and measure its temperature. Add the salt and measure again after 1 minute.
3. Put the small bag inside, seal, and shake (wrapped in a towel) for 5–8 minutes.
4. Record the lowest temperature of the ice bath. Enjoy the ice cream.

Check your work

- ☐ How much colder did the salted ice get?
- ☐ Why did the ice melt but get colder?

LAB NOTEBOOK

1. Temperatures of plain and salted ice.
2. An explanation using freezing-point depression.

KEY FORMULAS

$$m = \frac{\text{mol solute}}{\text{kg solvent}} \qquad \Delta T_f = iK_f m \qquad \Delta T_b = iK_b m$$

> **From history.** Antarctic icefish have antifreeze glycoproteins that bind to tiny ice crystals and keep their blood from freezing at −1.9 °C, colder than ordinary fish blood can survive.

WORKED EXAMPLES

Example 1. Find the freezing point of a solution of 58.5 g NaCl in 1.00 kg water.

▼ **Show solution**

1. $58.5/58.44 = 1.00$ mol; $m = 1.00$
2. $\Delta T_f = (2)(1.86)(1.00) = 3.72$ °C

Answer: −3.72 °C

Example 2. Which lowers the freezing point more per mole: NaCl or $\mathrm{CaCl_2}$?

▼ **Show solution**

1. NaCl gives 2 particles; $\mathrm{CaCl_2}$ gives 3.

Answer: $\mathrm{CaCl_2}$

PRACTICE

1. Easy What is a colligative property?

   ▼ **Show answer**

   One that depends on the number of dissolved particles, not their identity.

2. Easy What is the van ’t Hoff factor for NaCl?

   ▼ **Show answer**

   2

3. Hard 53 g of NaCl is dissolved in 0.86 kg of water. Estimate the new freezing point. ($K_f = 1.86$ °C·kg/mol; $i = 2$)

   ▼ **Show answer**

   -3.92 °C

4. **Hard** 144 g of KCl is dissolved in 0.34 kg of water. Estimate the new freezing point. ($K_f = 1.86$ °C·kg/mol; $i = 2$)

   ▼ **Show answer**

   $-21.1\ ^\circ\text{C}$

5. **Hard** 115 g of NaCl is dissolved in 0.69 kg of water. Estimate the new freezing point. ($K_f = 1.86$ °C·kg/mol; $i = 2$)

   ▼ **Show answer**

   $-10.6\ ^\circ\text{C}$

COMMON MISTAKES

- Using molarity instead of molality.
- Forgetting the van ’t Hoff factor for ionic solutes.
- Reporting the change instead of the new freezing point (subtract from 0 °C).

> **Make it a game · Particle Counting.** The parent names a solute (sugar, NaCl, $CaCl_2$, $Al(NO_3)_3$, glucose, KBr). The student gives the van ’t Hoff factor and ranks the effect on freezing point.

> **Scripture.** “By the breath of God frost is given: and the breadth of the waters is straitened.” (Job 37:10, KJV)
>
> By the breath of God frost is given. Salt and sugar change where that frost begins.

UNIT 8 · SOLUTIONS, ACIDS AND BASES · LESSON 8.3 OF 36 · 4 SESSIONS OF 45 MINUTES

8.3 Acids and Bases: Arrhenius (1884) to Brønsted and Lowry (1923)

OBJECTIVES

- List properties of acids and bases.
- Define acids and bases by the Arrhenius and Brønsted–Lowry theories.
- Identify conjugate acid–base pairs and distinguish strong from weak acids.

Warm-up. Make red cabbage juice: chop a few leaves, cover with hot water for 10 minutes, and strain. Add a spoon of the purple juice to small cups of vinegar, lemon juice, baking soda water and soapy water. Record the colors.

WATCH

Intro to acids and bases

Khan Academy · 12:13

https://www.youtube.com/watch?v=ja7p_tzTTEA

> **WHEN TO WATCH**
> After the warm-up.

BACKUP VIDEO

Properties of acids and bases

Khan Academy · 7:02

https://www.youtube.com/watch?v=028VyNbK2bY

With conjugate pairs.

KEY TERMS

acid

Arrhenius: produces H^+ in water. Brønsted–Lowry: a proton donor.

base

Arrhenius: produces OH^- in water. Brønsted–Lowry: a proton acceptor.

hydronium ion

H_3O^+, the form H^+ takes in water.

conjugate pair

Two species differing by one proton, such as NH_4^+ and NH_3.

strong acid

An acid that ionizes completely in water, such as HCl.

neutralization

The reaction of an acid with a base to form water and a salt.

CORE IDEAS

The story. Acids were known by taste (sour, Latin *acidus*) and by turning litmus red; bases (alkalis, from Arabic *al-qali*, plant ashes) were bitter and slippery. Lavoisier thought all acids contained oxygen (the name means "acid-former"), but Davy showed hydrochloric acid has none. In 1884 Arrhenius defined acids as substances that release H^+ ions in water and bases as those that release OH^-. In 1923 Johannes Brønsted in Copenhagen and Thomas Lowry in Cambridge independently broadened the idea.

How they figured it out. Arrhenius could not explain ammonia: it contains no OH yet turns litmus blue. Brønsted and Lowry said an acid is any proton (H^+) donor and a base any proton acceptor. Ammonia accepts a proton from water: $NH_3 + H_2O \rightleftharpoons NH_4^+ + OH^-$. Water acts as an acid here and as a base with HCl: $HCl + H_2O \rightarrow H_3O^+ + Cl^-$. Substances that can do both are amphoteric.

The chemistry. After an acid gives up a proton, what remains is its conjugate base; a base that gains a proton becomes its conjugate acid. Strong acids (HCl, HBr, HI, HNO_3, H_2SO_4, $HClO_4$) ionize completely; weak acids such as acetic acid ionize only a few percent. Strong does not mean concentrated: dilute HCl is still strong; concentrated vinegar is still weak. Neutralization: acid + base → salt + water.

Where it is used. Stomach acid (HCl) digests protein; antacids (bases) neutralize it. Baking soda (a base) reacts with buttermilk or cream of tartar (acids) to raise bread. Soil acidity decides what crops grow; farmers add lime. Car batteries hold sulfuric acid.

> **Safety first.** Goggles on. Adult handles ammonia cleaner and washing soda (both caustic). Never mix ammonia with bleach. Do not taste anything.

LAB

Red Cabbage Indicator Rainbow

You will build a color scale with red cabbage indicator and test household substances.

You will need

- Red cabbage juice (from the warm-up), clear cups or an ice-cube tray
- Lemon juice, vinegar, soda water, milk, distilled water, baking soda solution, borax solution, washing soda solution, diluted ammonia cleaner (adult)

Steps

1. Put 2 tablespoons of indicator in each cup.
2. Add a teaspoon of each substance and record the color.
3. Arrange the cups from red (most acidic) through purple to green and yellow (most basic).
4. Neutralize: add baking soda solution a little at a time to a vinegar cup and watch the color change.

Check your work

- ☐ Rank the substances from most acidic to most basic.
- ☐ What happened as you neutralized the vinegar?

LAB NOTEBOOK

1. A color chart of all substances.
2. A ranked list and the neutralization observations.

> **From history.** Lowry and Brønsted published within weeks of each other in 1923. The same year G. N. Lewis proposed an even broader definition based on electron pairs, used in advanced chemistry.

Another video. Cans in Acid and Base (reaction only). A short clip with no narration: watch it twice, then describe what happened in your own words.



For parents: Periodic Videos, 0:39. A wordless 39-second clip: soft-drink cans dissolving in strong acid and strong base. Periodic Videos is an approved channel. Periodic Videos · 0:39

WORKED EXAMPLES

Example 1. Identify the acid, base and conjugate pairs: $HNO_3 + H_2O \rightarrow H_3O^+ + NO_3^-$.

▼ **Show solution**

1. HNO_3 donates a proton: acid. Water accepts: base.
2. Pairs: HNO_3/NO_3^- and H_3O^+/H_2O.

Answer: Acid HNO_3, base H_2O; conjugate base NO_3^-, conjugate acid H_3O^+

Example 2. Write the neutralization of sulfuric acid by sodium hydroxide.

▼ **Show solution**

1. Acid + base → salt + water
2. Balance: two OH^- for two H^+.

Answer: $H_2SO_4 + 2NaOH \rightarrow Na_2SO_4 + 2H_2O$

PRACTICE

1. **Easy** Give two properties of acids.

   ▼ **Show answer**

   Sour taste, turn litmus red, react with metals to give hydrogen, react with carbonates to give CO_2.

2. **Easy** Give two properties of bases.

   ▼ **Show answer**

   Bitter taste, slippery feel, turn litmus blue.

3. **Medium** State the Arrhenius definitions of acid and base.

   ▼ **Show answer**

   Acids produce H^+ in water; bases produce OH^-.

COMMON MISTAKES

- Confusing strong with concentrated.
- Forgetting that H^+ in water is really H_3O^+.
- Writing the conjugate base with an extra proton instead of one fewer.

REFERENCE GUIDE

An eight-page, color-coded Organic Chemistry guide: carbon bonding and hybridization, structural formulas, functional groups, naming, isomers and chirality, acids and bases with pKa values, reaction types and mechanisms, spectroscopy (IR and NMR), and the molecules of life. The booklet prints on 11×17 paper and folds to letter size. Prints in color and reads clearly in black and white.

> **Make it a game · Acid, Base or Neither?.** The parent names everyday items (lemon, soap, water, coffee, bleach, milk of magnesia, cola, salt water). The student calls acid, base or neutral and predicts the cabbage color.

> **Scripture.** “As he that taketh away a garment in cold weather, and as vinegar upon nitre, so is he that singeth songs to an heavy heart.” (Proverbs 25:20, KJV)
>
> Vinegar upon nitre fizzes, a picture of a song sung to a heavy heart. Solomon knew an acid–base reaction when he saw one.

8.4 pH, Indicators and Titration: Sørensen at the Carlsberg Laboratory (1909)

OBJECTIVES

- Calculate pH and pOH from concentrations and the reverse.
- Explain how indicators work.
- Perform a titration and calculate an unknown concentration.

> **Warm-up.** One drop of lemon juice in a glass of water: is it still acidic? How could you put a number on how acidic something is?

WATCH

Introduction to pH

Khan Academy · 5:22

https://www.youtube.com/watch?v=oda2K4IFBaE

> **WHEN TO WATCH**
> After the pH scale.

BACKUP VIDEO

Acid Base Titration Problems, Basic Introduction, Calculations, Examples, Solution Stoichiometry

The Organic Chemistry Tutor · 18:35

https://www.youtube.com/watch?v=HJvALCcKYAc

Before the titration lab.

KEY TERMS

pH

$-\log[H^+]$; below 7 acidic, 7 neutral, above 7 basic at 25 °C.

pOH

$-\log[OH^-]$; pH + pOH = 14.

indicator

A dye that changes color over a certain pH range.

titration

Adding a solution of known concentration to a measured sample until the reaction is complete.

equivalence point

The point where moles of acid and base are equal.

buffer

A solution that resists changes in pH.

CORE IDEAS

The story. Søren Sørensen directed the chemistry department of the Carlsberg Laboratory in Copenhagen, founded by the brewer J. C. Jacobsen to put brewing on a scientific footing. Studying how enzymes in malt worked, Sørensen found that acidity mattered enormously and needed a convenient number. In 1909 he proposed pH: the negative logarithm of the hydrogen ion concentration. A concentration of 0.0000001 M became simply pH 7.

How they figured it out. Water itself ionizes slightly: $[\mathrm{H^+}][\mathrm{OH^-}] = 1.0 \times 10^{-14}$ at 25 °C. In pure water both are 10^{-7} M, pH 7. Each pH unit is a factor of 10: pH 3 is ten times more acidic than pH 4 and a thousand times more than pH 6. Indicators are weak acids whose acid and base forms have different colors; phenolphthalein is colorless below about 8.2 and pink above.

The chemistry. $\mathrm{pH} = -\log[\mathrm{H^+}]$, $[\mathrm{H^+}] = 10^{-\mathrm{pH}}$, pH + pOH = 14. In a titration, base from a burette (or medicine syringe) is added to an acid with indicator until the color just changes. At the equivalence point moles acid = moles base (adjusted for the equation), so $M_a V_a = M_b V_b$ for 1:1 reactions.

Where it is used. Blood is held at pH 7.35–7.45 by buffers; outside that range life is in danger. Pools are kept near 7.4. Gardeners test soil pH (blueberries like 4.5–5.5). Food scientists titrate acidity for canning safety; the brewing and dairy industries live by pH.

> **Safety first.** Goggles and gloves on. Bases are caustic; an adult prepares them. Rinse spills with plenty of water.

LAB

How Strong Is Your Vinegar?

You will titrate vinegar with a sodium hydroxide solution of known strength and calculate its acetic acid percent.

You will need

- White vinegar, a 10 mL medicine syringe (no needle) as a burette, a 5 mL syringe for the vinegar
- 0.50 M sodium hydroxide solution from a science supplier (about $10 for 500 mL); adult handles it
- Phenolphthalein drops or red cabbage indicator, a white saucer, a clear glass

Steps

1. Measure 2.0 mL vinegar into the glass with 20 mL water and 3 drops of indicator.
2. Fill the 10 mL syringe with base. Add it slowly while swirling until the color change lasts 30 seconds.
3. Record the volume used. Repeat twice and average.
4. Calculate the molarity and percent of acetic acid.

Check your work

- ☐ How close are your three trials?
- ☐ Is the label (usually 5%) accurate?

LAB NOTEBOOK

1. Three titration volumes and the average.
2. The full calculation of molarity and percent.

KEY FORMULAS

$$\mathrm{pH} = -\log[\mathrm{H^+}] \qquad [\mathrm{H^+}] = 10^{-\mathrm{pH}} \qquad \mathrm{pH} + \mathrm{pOH} = 14$$

$$M_a V_a = M_b V_b \;\; (1:1)$$

From history. Carlsberg's founder made the laboratory independent so its results would be published freely; Sørensen's colleague Emil Christian Hansen isolated pure brewing yeast in 1883 and gave the method away to rival breweries.

Try the lab. WeduLab AnalytLab: a free virtual titration bench. Turn the stopcock drop by drop, find the phenolphthalein endpoint against the light box, read the burette at eye level, then see the pH titration curve. Free, with no account. Opens in a new tab.

WORKED EXAMPLES

Example 1. Find the pH of 0.0010 M HCl.

▼ **Show solution**

1. HCl is strong: $[\mathrm{H^+}] = 1.0 \times 10^{-3}$
2. $\mathrm{pH} = -\log(10^{-3}) = 3.00$

Answer: pH 3.00

Example 2. Find the pH of 0.010 M NaOH.

▼ **Show solution**

1. $[\mathrm{OH^-}] = 10^{-2}$, pOH = 2.00
2. pH = 14 − 2.00

Answer: pH 12.00

PRACTICE

1. Easy What is the pH of pure water at 25 °C?

   ▼ **Show answer**

   7

2. Easy Is pH 9 acidic or basic?

   ▼ **Show answer**

   Basic.

3. Medium Find the pH of a solution with $[\mathrm{H}^+] = 10^{-5}$ M. Is it acidic, basic or neutral?

   ▼ **Show answer**

   pH = 5.00, acidic

4. Medium A solution has pH 3.3. Find $[\mathrm{H}^+]$.

   ▼ **Show answer**

   5.0×10^{-4} M

5. Hard Titration: 25.0 mL of vinegar (acetic acid) is neutralized by 23.0 mL of 0.97 M sodium hydroxide (1:1). Find the molarity of the acid.

   ▼ **Show answer**

   0.892 M

COMMON MISTAKES

- Forgetting the negative sign in the pH formula.
- Thinking pH 4 is twice as acidic as pH 8 (it is 10,000 times).
- Using the 1:1 shortcut when the ratio is not 1:1.

> **Make it a game · Powers of Ten.** The parent names two pH values; the student says how many times more acidic one is. Then reverse with concentrations.

> **Scripture.** “Divers weights are an abomination unto the LORD; and a false balance is not good.” (Proverbs 20:23, KJV)
>
> Divers weights are abomination; honest measurement, including a titration done carefully, honors God.

UNIT 9

Energy, Rates, Equilibrium and the Nucleus

The last unit asks three questions about every reaction: how much energy does it release or absorb, how fast does it go, and how far does it go? Students measure heat with a coffee-cup calorimeter in the tradition of Lavoisier and Laplace’s ice calorimeter and use Hess’s law; they speed up reactions with temperature, concentration and catalysts; and they push equilibria with Le Chatelier’s principle. The course ends where chemistry meets physics: Becquerel’s fogged plates, the Curies’ polonium and radium, and half-life.

The Physics course’s Unit 11 goes deeper into nuclear energy and relativity for students who take both.

PREREQUISITES

- Units 5–8, especially balanced equations, moles and molarity

MASTERY CHECKLIST

- ☐ I can calculate heat with $q = mc\Delta T$, classify reactions as exothermic or endothermic and use Hess’s law.
- ☐ I can explain reaction rates with collision theory and describe how catalysts work.
- ☐ I can describe dynamic equilibrium and predict shifts with Le Chatelier’s principle.
- ☐ I can write nuclear equations for alpha, beta and gamma decay and solve half-life problems.

LESSONS IN THIS UNIT

9.1 Thermochemistry: Lavoisier and Laplace's Ice Calorimeter and Hess's Law (1840)

4 sessions of 45 minutes

9.2 Reaction Rates and Catalysts — 4 sessions of 45 minutes

9.3 Equilibrium and Le Chatelier's Principle (1884) — 4 sessions of 45 minutes

9.4 Nuclear Chemistry: Becquerel (1896), the Curies (1898) and Half-Life

4 sessions of 45 minutes

Unit 9 test

Show your work. For every calculation write the equation, substitute with units and give the answer with its unit and sensible significant figures. The test is worth 100 points.

1. **8 pts** **Lesson 9.1**
   If ΔH for A → B is −50 kJ, what is ΔH for B → A?

2. **8 pts** **Lesson 9.1**
   What does a calorimeter measure?

3. **8 pts** **Lesson 9.1**
   1562 J of heat is added to 99 g of iron ($c = 0.449$ J/(g·°C)). By how much does its temperature rise?

4. **8 pts** **Lesson 9.2**
   On an energy diagram, how can you tell an endothermic reaction?

5. **8 pts** **Lesson 9.2**
   Two conditions must be met for a collision to react. What are they?

6. **7 pts** **Lesson 9.3**

   How does hemoglobin use Le Chatelier’s principle?

7. **7 pts** **Lesson 9.3**

   Does equilibrium mean equal amounts of reactants and products?

8. **7 pts** **Lesson 9.4**

   How do time, distance and shielding protect against radiation?

9. **7 pts** **Lesson 9.4**

   What is half-life?

10. **7 pts** **Lesson 9.4**

    radon-222 has a half-life of 3.82 days. How much of a 400 mg sample remains after 7.64 days?

LAB AND DATA

11. **10 pts**

    In a coffee-cup calorimeter, 50.0 mL of 1.0 M HCl and 50.0 mL of 1.0 M NaOH, both at 21.0 °C, are mixed. The temperature rises to 27.6 °C. Assume 100. g of solution with $c = 4.184$ J/(g·°C). (a) Find q for the solution. (b) Find the heat released per mole of water formed. (c) Is the reaction exothermic or endothermic?

SHORT ESSAY

12. **15 pts**

    A factory makes ammonia: $\mathrm{N_2 + 3H_2 \rightleftharpoons 2NH_3 + heat}$. Use Le Chatelier’s principle and collision theory to explain the choice of high pressure, a moderate temperature of about 450 °C, an iron catalyst and removal of ammonia as it forms.

ANSWER KEY

1. +50 kJ
2. Heat.
3. 35.1 °C

   $\Delta T = q/(mc) = 1562/((99)(0.449)) = 35.1\ ^\circ\text{C}$
4. The products are higher in energy than the reactants.
5. Enough energy and the right orientation.
6. In the lungs high O_2 shifts it to bind oxygen; in muscles low O_2 shifts it to release oxygen.
7. No; it means constant amounts and equal rates.
8. Less time exposed, more distance from the source and barriers all reduce the dose.
9. The time for half of a radioactive sample to decay.
10. 100 mg

    7.64/3.82 = 2 half-lives · 400 × (1/2)^2 = 100 mg
11. (a) $q = (100.)(4.184)(6.6) = 2760$ J. (b) 0.050 mol water: 2760/0.050 = 55,200 J/mol ≈ 55 kJ/mol released ($\Delta H \approx -55$ kJ/mol). (c) Exothermic.
12. Look for: high pressure favors the side with fewer gas moles (2 vs 4), shifting right; the reaction is exothermic, so low temperature favors ammonia but makes the rate too slow; 450 °C is a compromise between yield and rate; a catalyst speeds both directions and gets to equilibrium faster without changing the yield; removing NH_3 shifts the equilibrium right; collision theory: higher temperature and pressure mean more frequent, energetic collisions.

Printable answer key (PDF, for the parent)

9.1 Thermochemistry: Lavoisier and Laplace's Ice Calorimeter and Hess's Law (1840)

OBJECTIVES

- Distinguish heat and temperature; classify reactions as exothermic or endothermic.
- Calculate heat with $q = mc\Delta T$ and use a calorimeter.
- Use Hess's law to combine reaction enthalpies.

> **Warm-up.** Hold an instant cold pack and a hand warmer (if you have them), or dissolve calcium chloride and then ammonium nitrate fertilizer (adult) in water. Which got warm and which got cold? Where did the heat go?

WATCH

Food Calorimetry Lab: Explanation

Tyler DeWitt · 7:25

https://www.youtube.com/watch?v=uv7sVIqgyqU

> **WHEN TO WATCH**
> After the calorimeter story.

BACKUP VIDEO

Hess's law

Khan Academy · 5:44

https://www.youtube.com/watch?v=QU3zVec5I_M

With Hess's law.

KEY TERMS

exothermic
Releasing heat to the surroundings; ΔH negative.

endothermic
Absorbing heat from the surroundings; ΔH positive.

enthalpy change (ΔH)

The heat released or absorbed by a reaction at constant pressure.

specific heat

The heat needed to raise 1 g of a substance by 1 °C; water 4.184 J/(g·°C).

calorimeter

An insulated device for measuring heat.

Hess's law

The total ΔH of a reaction is the same whether it happens in one step or several.

CORE IDEAS

The story. In the winter of 1782–1783 Antoine Lavoisier and the mathematician Pierre-Simon Laplace built an ice calorimeter: a chamber surrounded by packed ice. Whatever heat a reaction or a living guinea pig released melted ice, and weighing the meltwater measured the heat. They showed that the heat of a guinea pig's breathing roughly equaled the heat of burning the same amount of carbon: respiration is a slow combustion. In 1840 Germain Hess, a Swiss-born chemist in St. Petersburg, found that the heat of a reaction is the same no matter how many steps it takes.

How they figured it out. Energy is conserved. Breaking bonds absorbs energy; forming bonds releases it. If the bonds formed are stronger than those broken, the reaction releases heat (exothermic). Heat flowing out of a reaction flows into the water of a calorimeter, so measuring the water's temperature change measures the heat.

The chemistry. $q = mc\Delta T$. In a coffee-cup calorimeter, $q_{\text{reaction}} = -q_{\text{water}}$. Enthalpy change ΔH is reported per mole of reaction as written: $\mathrm{CH_4 + 2O_2 \rightarrow CO_2 + 2H_2O}$, ΔH = −890 kJ. Hess's law: if equations add up to a target equation, their ΔH values add the same way; reversing an equation changes the sign, and multiplying it multiplies ΔH.

Where it is used. Food Calories are kilocalories measured by burning food in a bomb calorimeter. Fuels are compared by heat of combustion. Instant cold packs use endothermic dissolving of ammonium nitrate; hand warmers use the exothermic rusting of iron powder.

> **Safety first.** Adult lights the nut; do this away from anyone with nut allergies. Ammonium nitrate is an oxidizer: keep it away from heat and fuels and dispose of the solution down the drain with plenty of water.

LAB

Coffee-Cup Calorimetry and the Burning Peanut

You will measure heats of dissolving and estimate the energy in a nut.

You will need

- Two nested foam cups with a lid, a thermometer, a 0.1 g scale
- Calcium chloride ice melt, ammonium nitrate (cold-pack contents, adult), water
- A peanut or cashew, a paper clip stand, a small metal can, a lighter (adult)

Steps

1. Put 100. g of water in the cups and record the temperature. Add 5.0 g of calcium chloride, stir, and record the highest temperature.
2. Repeat with ammonium nitrate and record the lowest temperature.
3. Weigh a nut on a paper clip stand. With an adult, light it under a can holding 100. g of water. Record the temperature rise and weigh the leftover nut.
4. Calculate q in each case.

Check your work

- ☐ Which dissolving was exothermic? Endothermic?
- ☐ Why is your nut's energy value lower than the label?

LAB NOTEBOOK

1. Temperature data and q calculations for all three trials.
2. Sources of error in the nut experiment.

KEY FORMULAS

$$q = mc\Delta T \qquad q_{\text{rxn}} = -q_{\text{water}}$$

$$1\ \text{cal} = 4.184\ \text{J} \qquad 1\ \text{Cal (food)} = 1\ \text{kcal}$$

From history. Lavoisier and Laplace's calorimeter survives in the Musée des Arts et Métiers in Paris. Lavoisier was executed in 1794 during the Revolution; the mathematician Lagrange said it took only a moment to cut off that head, and a hundred years might not produce another like it.

WORKED EXAMPLES

Example 1. A 1.00 g peanut burns under a can holding 100. g of water, raising it from 20.0 °C to 45.0 °C. Estimate the energy in food Calories per gram.

▼ **Show solution**

1. $q = (100.)(4.184)(25.0) = 10{,}460$ J
2. $10{,}460 / 4184 = 2.50$ kcal

Answer: About 2.5 Cal/g (heat is lost, so the true value, about 6, is higher)

Example 2. Given C + O_2 → CO_2, ΔH = −394 kJ and CO + ½O_2 → CO_2, ΔH = −283 kJ, find ΔH for C + ½O_2 → CO.

▼ **Show solution**

1. Keep the first; reverse the second: CO_2 → CO + ½O_2, ΔH = +283.
2. Add: C + ½O_2 → CO.

Answer: ΔH = −111 kJ

PRACTICE

1. Easy Is burning wood exothermic or endothermic?

▼ **Show answer**

Exothermic.

2. **Easy** What is the specific heat of water?

   ▼ **Show answer**

   4.184 J/(g·°C)

3. **Medium** 177 g of water warms from 19.5 °C to 34.3 °C. How much heat did it absorb? ($c = 4.184$ J/(g·°C))

   ▼ **Show answer**

   1.10×10^4 J

4. **Hard** 2110 J of heat is added to 103 g of aluminum ($c = 0.897$ J/(g·°C)). By how much does its temperature rise?

   ▼ **Show answer**

   22.8 °C

5. **Hard** 1358 J of heat is added to 65 g of iron ($c = 0.449$ J/(g·°C)). By how much does its temperature rise?

   ▼ **Show answer**

   46.5 °C

COMMON MISTAKES

- Confusing heat (energy) with temperature.
- Forgetting the sign: heat released by the reaction is gained by the water.
- Not reversing the sign of ΔH when reversing an equation.

> **Make it a game · Hot or Cold?.** The parent names a process (burning wood, melting ice, freezing water, photosynthesis, respiration, a cold pack, dynamite). The student calls exothermic or endothermic and gives the sign of ΔH.

> **Scripture.** "For our God is a consuming fire." (Hebrews 12:29, KJV)
>
> God is a consuming fire. Combustion is the most familiar exothermic reaction and a picture Scripture uses of God's holiness.

9.2 Reaction Rates and Catalysts

OBJECTIVES

- Explain reaction rate with collision theory and activation energy.
- Describe how temperature, concentration, surface area and catalysts affect rate.
- Read an energy diagram.

> **Warm-up.** Drop a whole antacid tablet in cold water and another, crushed, in hot water. Time how long each takes to stop fizzing.

WATCH

GCSE Chemistry - Factors Affecting Rate of Reaction & Collision Theory

Cognito · 5:38

https://www.youtube.com/watch?v=JK7yPzO9POU

> **WHEN TO WATCH**
> After the antacid race.

BACKUP VIDEO

Activation Energy (and exploding bags of Chlorine) - Periodic Table of Videos

Periodic Videos · 10:26

https://www.youtube.com/watch?v=hLbAarRWJic

With the catalyst discussion.

KEY TERMS

reaction rate
The change in amount of a reactant or product per unit time.

collision theory
Particles must collide with enough energy and the right orientation to react.

activation energy

The minimum energy needed for a collision to produce a reaction.

catalyst

A substance that speeds a reaction by lowering its activation energy without being used up.

enzyme

A protein catalyst in living things.

energy diagram

A graph of energy during a reaction, showing reactants, products and the activation hill.

CORE IDEAS

The story. In 1835 Berzelius coined the word catalysis for reactions sped up by substances that are not consumed. In 1889 Arrhenius showed that a rise of 10 °C roughly doubles many reaction rates and explained it by an energy barrier only some collisions can cross. Wilhelm Ostwald, who studied catalysts and their effect on rates, won the 1909 Nobel Prize and developed the process for making nitric acid from ammonia over a platinum catalyst.

How they figured it out. Particles react only when they collide hard enough and in the right orientation. Higher concentration means more collisions; more surface area exposes more particles; higher temperature makes particles move faster, so collisions are more frequent and, more importantly, more of them carry enough energy to clear the activation-energy hill. A catalyst provides a different path with a lower hill.

The chemistry. On an energy diagram, reactants sit on the left, products on the right, and the activation energy E_a is the height of the hill above the reactants. ΔH is the difference between products and reactants. A catalyst lowers E_a for both the forward and reverse reactions but does not change ΔH or the final amounts at equilibrium.

Where it is used. Refrigeration slows the reactions that spoil food. Flour and sawdust explosions happen because powder has huge surface area. Catalytic converters in cars use platinum, palladium and rhodium to turn carbon monoxide and unburned fuel into carbon dioxide and water. Enzymes speed reactions in the body by factors of millions (Biology Unit 2).

LAB

Antacid Rate Races

You will vary temperature, surface area and concentration and time an antacid tablet's reaction.

You will need

- Effervescent antacid tablets (8), clear cups, a thermometer, a timer
- Ice water, room-temperature water, hot tap water, vinegar

Steps

1. Time a whole tablet in 100 mL of ice, room and hot water.
2. At room temperature, compare a whole, a broken (4 pieces) and a crushed tablet.
3. Compare 100 mL of water with 100 mL of half-vinegar solution.
4. Graph reaction time versus temperature.

Check your work

- ☐ Which factor had the biggest effect?
- ☐ Explain each result with collision theory.

LAB NOTEBOOK

1. Three data tables and the temperature graph.
2. Collision-theory explanations of each result.

> **From history.** Paul Sabatier won the 1912 Nobel Prize for using finely divided nickel to add hydrogen to oils, the basis for making margarine and shortening.

WORKED EXAMPLES

Example 1. Hydrogen peroxide decomposes slowly. Explain why adding a pinch of yeast makes it foam immediately.

▼ **Show solution**

1. Yeast contains the enzyme catalase, a catalyst.
2. It lowers the activation energy for $2\mathrm{H_2O_2} \rightarrow 2\mathrm{H_2O} + \mathrm{O_2}$.
3. More collisions succeed per second, so oxygen bubbles form quickly.

Answer: The enzyme catalase catalyzes the decomposition

Example 2. A reaction's rate doubles for every 10 °C. How much faster is it at 50 °C than at 20 °C?

▼ **Show solution**

1. 30 °C rise = three doublings
2. $2^3 = 8$

Answer: 8 times faster

PRACTICE

1. Easy Name four factors that affect reaction rate.

   ▼ **Show answer**

   Temperature, concentration, surface area, catalysts.

2. Easy What is activation energy?

   ▼ **Show answer**

   The minimum energy colliding particles need to react.

3. Medium Why does crushing a tablet make it react faster?

   ▼ **Show answer**

   More surface area exposes more particles to collisions.

COMMON MISTAKES

- Saying catalysts are used up.

- Saying a catalyst changes ΔH or the yield.
- Thinking temperature speeds reactions only by making collisions more frequent; the bigger effect is more energetic collisions.

> **Make it a game · Faster or Slower?.** The parent describes a change (putting milk in the fridge, chewing food, adding a catalyst, using a log instead of kindling). The student says faster or slower and names the factor.

> **Scripture.** "To every thing there is a season, and a time to every purpose under the heaven:" (Ecclesiastes 3:1, KJV)
>
> To every thing there is a season. Chemists study the time reactions take and learn to speed or slow them.

UNIT 9 · ENERGY, RATES, EQUILIBRIUM AND THE NUCLEUS · LESSON 9.3 OF 36 · 4 SESSIONS OF 45 MINUTES

9.3 Equilibrium and Le Chatelier's Principle (1884)

OBJECTIVES

- Describe dynamic equilibrium in reversible reactions.
- Write equilibrium expressions and interpret K.
- Predict shifts with Le Chatelier's principle.

> **Warm-up.** Two people pass buckets of water between two pools at the same rate. Do the water levels change? Is anyone resting? Now one person speeds up. What happens?

WATCH

GCSE Chemistry - Reversible Reactions & Dynamic Equilibrium

Cognito · 6:28

https://www.youtube.com/watch?v=5rC6f_P_A48

> **WHEN TO WATCH**
> After the definition of equilibrium.

BACKUP VIDEO

GCSE Chemistry - Le Chatelier's Principle - Position of Equilibrium in Reversible Reactions

Cognito · 4:13

https://www.youtube.com/watch?v=KZPO23kia_o

With Le Chatelier’s principle.

KEY TERMS

reversible reaction

A reaction that can proceed in both directions, shown with ⇌.

dynamic equilibrium

The state in which forward and reverse rates are equal, so amounts stay constant.

equilibrium constant (K)

The ratio of product to reactant concentrations at equilibrium, each raised to its coefficient.

Le Chatelier’s principle

A system at equilibrium shifts to partly undo a stress placed on it.

stress

A change in concentration, pressure or temperature applied to an equilibrium.

shift

A change in the equilibrium position toward reactants or products.

CORE IDEAS

The story. On Napoleon’s expedition to Egypt in 1798, the chemist Claude Berthollet noticed sodium carbonate crusts around salt lakes, though in the lab sodium carbonate and calcium chloride react the other way. He concluded that the amount

of a substance can push a reaction backward. In 1864 the Norwegians Cato Guldberg and Peter Waage stated the law of mass action. In 1884 Henri Le Chatelier in Paris gave the rule that bears his name.

How they figured it out. In a closed container, as products build up, the reverse reaction speeds up until the two rates are equal. At that point nothing seems to change, yet both reactions continue: a dynamic equilibrium. The ratio $K = \frac{[\mathrm{C}]^c[\mathrm{D}]^d}{[\mathrm{A}]^a[\mathrm{B}]^b}$ for $a\mathrm{A} + b\mathrm{B} \rightleftharpoons c\mathrm{C} + d\mathrm{D}$ is constant at a given temperature. Large K favors products; small K favors reactants.

The chemistry. Le Chatelier: adding a reactant shifts right; removing a product shifts right; increasing pressure (by compressing) shifts toward fewer gas moles; raising temperature shifts in the endothermic direction (treat heat as a reactant or product). Catalysts do not shift equilibrium; they only reach it sooner. Solids and pure liquids are left out of K.

Where it is used. The Haber process (Lesson 6.4) uses high pressure and ammonia removal. Your blood carries oxygen through the equilibrium of hemoglobin, which shifts to release oxygen in hard-working muscles. Carbonated drinks keep their fizz under pressure. Cobalt chloride humidity cards change color as an equilibrium shifts with moisture.

> **Safety first.** Cobalt chloride is toxic: handle the card by its edges and wash hands. Do not touch the colored area.

LAB

Shifting an Equilibrium with Color

You will push a color-changing equilibrium back and forth with temperature and concentration.

You will need

- Red cabbage indicator, vinegar, baking soda solution (pH equilibrium of the indicator)
- Optional: cobalt chloride humidity indicator card from a hardware store
- Ice water, hot tap water, droppers, clear cups

Steps

1. Add indicator to water and add vinegar a drop at a time until the color is just between purple and pink. This is an equilibrium of the indicator's acid and base forms.
2. Add more vinegar: record the color. Then add baking soda solution: record. This is a concentration stress.
3. Breathe on a cobalt chloride card or hold it over steam, then dry it with a hair dryer. Record colors.
4. Explain each change with Le Chatelier's principle.

Check your work

☐ Which direction did each stress shift the equilibrium?

☐ Is the color change reversible? What does that show?

LAB NOTEBOOK

1. A table of stress, color change and direction of shift.
2. Equations for the equilibria you studied.

> **From history.** Le Chatelier also studied cement, metal alloys and the dangers of mine explosions. He once said that the greatest mistake of his career was failing to pursue ammonia synthesis after an accident in his lab, leaving the discovery to Haber.

WORKED EXAMPLES

Example 1. Write K for $\mathrm{N_2}(g) + 3\mathrm{H_2}(g) \rightleftharpoons 2\mathrm{NH_3}(g)$.

▼ **Show solution**

1. Products over reactants, each to the power of its coefficient.

Answer: $K = \dfrac{[\mathrm{NH_3}]^2}{[\mathrm{N_2}][\mathrm{H_2}]^3}$

Example 2. For $2\mathrm{NO_2}(g) \rightleftharpoons \mathrm{N_2O_4}(g)$ + heat (brown ⇌ colorless), predict the color change when the flask is put in ice water and when it is compressed.

▼ **Show solution**

1. Cooling removes heat; the system replaces it by shifting toward the exothermic side, $\mathrm{N_2O_4}$.
2. Compressing favors fewer gas moles (1 vs 2), again $\mathrm{N_2O_4}$.

Answer: Both make the gas paler

PRACTICE

1. Easy What does the double arrow ⇌ mean?

   ▼ **Show answer**

   The reaction is reversible.

2. Easy At equilibrium, what is equal?

   ▼ **Show answer**

   The forward and reverse rates.

3. Medium Why is equilibrium called dynamic?

   ▼ **Show answer**

   Both reactions continue at equal rates, so amounts stay constant though change continues.

COMMON MISTAKES

- Thinking reactions stop at equilibrium.
- Thinking equilibrium means equal amounts of reactants and products.

- Including solids or liquids in K.

> **Make it a game · Which Way?.** The parent names an equilibrium and a stress. The student points left or right to show the shift and explains in one sentence.

> **Scripture.** “Thy faithfulness is unto all generations: thou hast established the earth, and it abideth.” (Psalm 119:90, KJV)
>
> God established the earth and it abides. Equilibrium, a balance that returns when disturbed, is one form of that stability.

UNIT 9 · ENERGY, RATES, EQUILIBRIUM AND THE NUCLEUS · LESSON 9.4 OF 36 · 4 SESSIONS OF 45 MINUTES

9.4 Nuclear Chemistry: Becquerel (1896), the Curies (1898) and Half-Life

OBJECTIVES

- Describe the discovery of radioactivity.
- Write nuclear equations for alpha, beta and gamma decay.
- Solve half-life problems and describe uses and dangers of radiation.

> **Warm-up.** A bowl of 100 pennies is shaken and dumped; every penny showing heads is removed. Predict how many remain after 1, 2 and 3 dumps. Try it.

WATCH

Nuclear Half Life: Intro and Explanation

Tyler DeWitt · 5:53

https://www.youtube.com/watch?v=opjJ-3Tkfyg

> **WHEN TO WATCH**
> After the story.

BACKUP VIDEO

Nuclear Half Life: Calculations

Tyler DeWitt · 8:04

https://www.youtube.com/watch?v=WAsmY4ocWSA

With half-life problems.

KEY TERMS

radioactivity

The spontaneous emission of radiation by unstable nuclei.

alpha particle

A helium-4 nucleus, $^{4}_{2}\mathrm{He}$, stopped by paper.

beta particle

A fast electron, $^{0}_{-1}e$, emitted when a neutron becomes a proton; stopped by aluminum.

gamma ray

High-energy electromagnetic radiation from a nucleus; reduced by thick lead or concrete.

half-life

The time for half of a radioactive sample to decay.

isotope

Atoms of the same element with different numbers of neutrons.

CORE IDEAS

The story. In February 1896 Henri Becquerel in Paris wrapped photographic plates in black paper and set uranium salts on them, planning to expose the salts to sunlight. Paris was cloudy for days, so he put everything in a drawer. Developing the plates anyway, he found strong images: uranium gave off rays by itself. Marie

Skłodowska Curie, a Polish doctoral student, measured the rays with an electrometer her husband Pierre had built and called the effect radioactivity. In 1898, grinding and dissolving tons of pitchblende ore in a leaky shed, the Curies discovered polonium (named for Poland) and radium.

How they figured it out. Rutherford separated the rays with magnets and absorbers: alpha (positive, heavy, stopped by paper), beta (negative, light, stopped by aluminum foil) and gamma (no charge, very penetrating). In 1902 Rutherford and Frederick Soddy showed that radioactive atoms change into other elements, the alchemists' transmutation in a way they never imagined. They also found that each isotope decays with a fixed half-life.

The chemistry. Nuclear equations balance both mass numbers (top) and atomic numbers (bottom). Alpha: $^{238}_{92}\mathrm{U} \rightarrow {}^{234}_{90}\mathrm{Th} + {}^{4}_{2}\mathrm{He}$. Beta: $^{14}_{6}\mathrm{C} \rightarrow {}^{14}_{7}\mathrm{N} + {}^{0}_{-1}e$. Gamma rays often accompany them and change neither number. After each half-life, half the remaining nuclei have decayed: $N = N_0(\tfrac{1}{2})^{t/t_{1/2}}$.

Where it is used. Iodine-131 treats thyroid disease; technetium-99m images bones and hearts; cobalt-60 sterilizes medical tools and kills tumors; smoke detectors use americium-241. Nuclear power plants split uranium (Physics Unit 11). Radiation can also damage cells and DNA, so exposure is limited by time, distance and shielding. Marie Curie's notebooks are still too radioactive to handle without protection.

THINKING ABOUT TIME

What the sign says. Radiometric dating is often quoted as proving that rocks are millions or billions of years old.

What we observe. We observe present decay rates in the laboratory; they are steady under ordinary conditions. Any age calculated from them also requires assumptions about how much parent and daughter material a rock started with and whether material entered or left it since. Carbon-14, with a 5,730-year half-life, is used for once-living things only.

How Christians read it. Christians agree that half-lives are real and measurable. Young-earth creationists argue that the starting assumptions are unprovable, point to dates that disagree for the same rock, and some propose that decay was faster in the past. Old-earth Christians accept the long ages and read Genesis 1 in other ways. Both seek to honor Scripture and the evidence.

Try it. Reread the penny model: how would your "age" estimate change if some pennies were removed before you started counting?

See High School Geology for a fuller treatment of dating methods.

LAB

Half-Life with Pennies (or M&Ms)

You will model radioactive decay with coins and graph the decay curve.

You will need

- 100 pennies (or candies with a printed side), a shoebox with a lid, graph paper

Steps

1. Put 100 pennies heads-up in the box. Close, shake well, and open.
2. Remove all pennies that landed tails (the “decayed” nuclei). Count and record the remainder.
3. Repeat until no pennies remain. Each shake is one half-life.
4. Graph pennies remaining versus number of shakes. Repeat the whole run once more and average.

Check your work

- ☐ How many shakes until about 25 remained?
- ☐ Is the graph a straight line? Why not?
- ☐ Can you predict which penny decays next?

LAB NOTEBOOK

1. Data for two runs and the average.
2. The decay curve with a half-life marked.

> **From history.** Marie Curie is the only person to win Nobel Prizes in two sciences: physics (1903, with Pierre and Becquerel) and chemistry (1911). She drove mobile X-ray vans to the front in World War I.

Another video. Types of Nuclear Radiation. A short extra on the same idea, after the core reading.



For parents: Fermilab, 9:23. Transcript checked. Fermilab's Dr. Don Lincoln on alpha, beta, gamma and neutron radiation and how to shield each, with a real (weak) source and a Geiger counter. Light jokes (Godzilla, comic books, and a closing quip about a Facebook status and getting back together with an ex). Only this video from the channel is approved. Fermilab · 9:23

PRACTICE

1. Easy Who discovered radioactivity, and how?

   ▼ **Show answer**

   Becquerel, 1896, when uranium salts fogged photographic plates in a drawer.

2. Easy Which two elements did the Curies discover in 1898?

   ▼ **Show answer**

   Polonium and radium.

3. Easy cobalt-60 has a half-life of 5.27 years. How much of a 100 mg sample remains after 21.08 years?

   ▼ **Show answer**

   6.25 mg

4. Medium A sample has dropped to 1/16 of its original activity. If the half-life is 3.82 days, how long has it decayed?

   ▼ **Show answer**

   15.28 days

5. Hard Using $N = N_0(1/2)^{t/t_{1/2}}$, how much of a 250.0 g sample of radon-222 ($t_{1/2} = 3.82$ days) remains after 11.0 days?

   ▼ **Show answer**

   34.0 g

COMMON MISTAKES

- Thinking a sample is all gone after two half-lives.
- Not balancing atomic numbers in nuclear equations.
- Confusing nuclear changes with chemical changes (chemical changes do not alter the nucleus).

> **Make it a game · Balance the Nucleus.** The parent writes a nuclear equation with one blank. The student fills in the missing particle or nucleus by balancing the top and bottom numbers.

> **Scripture.** "Of old hast thou laid the foundation of the earth: and the heavens are the work of thy hands. They shall perish, but thou shalt endure: yea, all of them shall wax old like a garment; as a vesture shalt thou change them, and they shall be changed:" (Psalm 102:25–26, KJV)
>
> The heavens and earth wear out like a garment. Radioactive decay is one small sign that creation is not static.

COURSE EXAM

Midterm Exam

Covers Units 1–5. Time: 100 minutes.

You have 100 minutes. A scientific calculator and a periodic table are allowed; no notes. For every calculation write the equation, substitute with units, and give the

answer with its unit and sensible significant figures; correct work earns partial credit. The exam has 22 problems worth a total of 100 points.

---

1. **5 pts** **Lesson 1.1**

   A theory is contradicted by a careful, repeated experiment. What should happen to the theory?

2. **5 pts** **Lesson 1.2**

   How many milliliters are in 0.250 L?

3. **5 pts** **Lesson 1.2**

   A road race is 12 miles long. Convert to kilometers (1 mi = 1.609 km) and then to meters.

4. **5 pts** **Lesson 1.3**

   Why might Galileo have thought the overflow method too imprecise for the crown?

5. **5 pts** **Lesson 1.3**

   What is the mass of 16.0 cm^3 of silver (density 10.5 g/cm^3)?

6. **5 pts** **Lesson 1.4**

   What separation method recovers salt from salt water?

7. **4 pts** **Lesson 2.1**

   Iron filings (5.58 g) are heated with sulfur. All the iron reacts, forming 8.79 g of iron sulfide. What mass of sulfur reacted?

8. **4 pts** **Lesson 2.1**

   16.83 g of steel wool burns completely in air, and the product weighs 22.96 g. What mass of oxygen combined with the iron?

9. **4 pts** **Lesson 2.2**

   A sample of carbon dioxide from a volcano and one from a soda can both contain 27.3% carbon. Another gas is 42.9% carbon. Is it carbon dioxide?

10. **4 pts** **Lesson 2.2**

Carbon forms two oxides. In the first, 3.7 g of carbon combines with 4.93 g of oxygen; in the second, 3.7 g of carbon combines with 9.86 g of oxygen. Find the ratio of the oxygen masses and say which law it illustrates.

11. **4 pts** **Lesson 2.4**

How many electrons does $^{32}_{16}\mathrm{S}^{2-}$ have?

12. **4 pts** **Lesson 2.4**

Natural chlorine has two isotopes: mass 34.969 u (75.78%) and mass 36.966 u (24.22%). Calculate the average atomic mass.

13. **4 pts** **Lesson 3.1**

An electron drops from a higher to a lower level and emits 486 nm light. Would a drop with a larger energy gap emit longer or shorter wavelengths?

14. **4 pts** **Lesson 3.1**

Find the energy of one photon of 486 nm light. ($h = 6.626 \times 10^{-34}$ J·s, $c = 3.00 \times 10^{8}$ m/s)

15. **4 pts** **Lesson 3.2**

How many unpaired electrons does oxygen have? Explain.

16. **4 pts** **Lesson 3.2**

How many valence electrons does lithium (Z = 3) have?

17. **4 pts** **Lesson 5.1**

Who gave chemistry its letter symbols?

18. **4 pts** **Lesson 5.1**

Balance: $\mathrm{Zn + HCl \rightarrow ZnCl_2 + H_2}$

19. **4 pts** **Lesson 5.2**

Which would react more vigorously with water: calcium or potassium? Why?

20. **4 pts** **Lesson 5.3**

Iron(III) chloride and sodium hydroxide are mixed. Predict the precipitate and write the net ionic equation.

21. **4 pts** **Lesson 5.4**

Why must aluminum be produced by electrolysis rather than by heating its ore with carbon?

ESSAY

22. **10 pts**

Choose one chemist from Units 1–5 (Boyle, Lavoisier, Dalton, Thomson, Rutherford, Mendeleev, Lewis or Faraday). Describe the key experiment or idea, what it changed, and how later work built on it.

ANSWER KEY

1. It must be revised or replaced; theories stay open to correction by evidence.
2. 250. mL
3. $19.3\ \text{km} = 1.93 \times 10^{4}\ \text{m}$
4. The difference in volume between a gold crown and one with some silver is only a few cubic centimeters, too small to see as an overflow; a balance under water magnifies the difference.
5. 168 g
6. Evaporation (or distillation, which also recovers the water).
7. 3.21 g
8. 6.13 g
9. No; a different percent composition means a different compound (it is carbon monoxide).
10. 2 : 1, the law of multiple proportions (CO and CO_2)
11. 18
12. 35.45 u
13. Shorter wavelengths (bluer light).
14. 4.09×10^{-19} J
15. Two; $2p^4$ puts one pair in one orbital and single electrons in the other two.
16. 1
17. Berzelius (1813).
18. $Zn + 2HCl \rightarrow ZnCl_2 + H_2$
19. Potassium; it is higher in the activity series (lower ionization energy).
20. $Fe^{3+} + 3OH^- \rightarrow Fe(OH)_3(s)$, a rust-brown solid.
21. Aluminum is very active; carbon cannot reduce its oxide at practical temperatures, so electricity is needed.
22. Look for: accurate dates and details; the experiment or reasoning described in the student's own words; a clear statement of what changed (e.g. Lavoisier: mass is conserved, oxygen explains combustion); one connection to later work; correct chemistry vocabulary.

Printable answer key (PDF, for the parent)

COURSE EXAM

Final Exam

Covers Units 1–9. Time: 100 minutes.

You have 100 minutes. A scientific calculator and a periodic table are allowed; no notes. For every calculation write the equation, substitute with units, and give the answer with its unit and sensible significant figures; correct work earns partial credit. The exam has 36 problems worth a total of 100 points.

---

1. **3 pts** **Lesson 1.1**
   List the main sections of a lab notebook entry used in this course.

2. **3 pts** **Lesson 1.2**
   Explain how the Mars Climate Orbiter was lost.

3. **3 pts** **Lesson 1.2**
   A road race is 8 miles long. Convert to kilometers (1 mi = 1.609 km) and then to meters.

4. **3 pts** **Lesson 1.3**
   Write the formula for density.

5. **3 pts** **Lesson 1.3**
   What is the mass of 36.2 cm^3 of zinc (density 7.13 g/cm^3)?

6. **3 pts** **Lesson 1.4**
   Who invented chromatography, and what did he first separate?

7. **3 pts** **Lesson 2.1**
   Why do chemical equations have to be balanced?

8. **3 pts** **Lesson 2.1**

   19.6 g of steel wool burns completely in air, and the product weighs 25.80 g. What mass of oxygen combined with the iron?

9. **3 pts** **Lesson 2.2**

   In what book and year did Dalton publish his atomic theory?

10. **3 pts** **Lesson 2.2**

    Carbon forms two oxides. In the first, 1.73 g of carbon combines with 2.31 g of oxygen; in the second, 1.73 g of carbon combines with 4.62 g of oxygen. Find the ratio of the oxygen masses and say which law it illustrates.

11. **3 pts** **Lesson 2.4**

    Give two uses of mass spectrometry.

12. **3 pts** **Lesson 2.4**

    Natural lithium has two isotopes: mass 6.015 u (7.59%) and mass 7.016 u (92.41%). Calculate the average atomic mass.

13. **3 pts** **Lesson 3.1**

    What did Bunsen design that made flame tests easier to see?

14. **3 pts** **Lesson 3.1**

    Light from a flame test has wavelength 434 nm. Find its frequency. ($c = 3.00 \times 10^8$ m/s)

15. **3 pts** **Lesson 3.2**

    State the Pauli exclusion principle.

16. **3 pts** **Lesson 3.2**

    How many valence electrons does iron (Z = 26) have?

17. **3 pts** **Lesson 5.1**

    Is $2H_2O \rightarrow 2H_2 + O_2$ balanced? Check atoms.

18. **3 pts** **Lesson 5.1**

    Balance: $Al + O_2 \rightarrow Al_2O_3$

19. **3 pts** **Lesson 5.3**

Is ammonium phosphate soluble?

20. **3 pts** **Lesson 6.1**

How many moles is 3.41×10^{24} atoms of iron?

21. **2 pts** **Lesson 6.2**

What is the percent by mass of C in CO_2?

22. **2 pts** **Lesson 6.3**

For $N_2 + 3H_2 \rightarrow 2NH_3$, what mass of NH_3 forms from 8.2 g of N_2?

23. **2 pts** **Lesson 6.4**

For $2Na + Cl_2 \rightarrow 2NaCl$, 30.7 g of Na reacts with 57.7 g of Cl_2. (a) Which is the limiting reactant? (b) What is the theoretical yield of NaCl? (c) If 73.1 g is actually collected, what is the percent yield?

24. **2 pts** **Lesson 7.1**

Convert 2.78 atm to kPa and to mmHg (1 atm = 101.3 kPa = 760 mmHg).

25. **2 pts** **Lesson 7.2**

A sealed can of gas is at 196 kPa and 20 °C. What is its pressure at 60 °C?

26. **2 pts** **Lesson 7.3**

What volume does 3.07 mol of gas occupy at STP (0 °C, 1 atm; 22.4 L/mol)?

27. **2 pts** **Lesson 7.4**

A mixture contains nitrogen at 218 mmHg, oxygen at 159 mmHg and water vapor at 21 mmHg. What is the total pressure?

28. **2 pts** **Lesson 8.1**

49.3 g of $CaCl_2$ is dissolved to make 405 mL of solution. Find the molarity.

29. **2 pts** **Lesson 8.2**

132 g of $CaCl_2$ is dissolved in 0.6 kg of water. Estimate the new freezing point. ($K_f = 1.86$ °C·kg/mol; $i = 3$)

30. **2 pts** **Lesson 8.3**

Why do farmers add lime to acidic soil?

31. **2 pts** **Lesson 8.4**

A solution has pH 9.8. Find $[\mathrm{H}^+]$.

32. **2 pts** **Lesson 9.1**

1441 J of heat is added to 46 g of aluminum ($c = 0.897$ J/(g·°C)). By how much does its temperature rise?

33. **2 pts** **Lesson 9.2**

Explain why a catalyst speeds the reverse reaction too.

34. **2 pts** **Lesson 9.3**

Increasing the pressure on $\mathrm{H_2}(g) + \mathrm{I_2}(g) \rightleftharpoons 2\mathrm{HI}(g)$ has no effect on the position. Why?

35. **2 pts** **Lesson 9.4**

A sample has dropped to 1/64 of its original activity. If the half-life is 5.27 years, how long has it decayed?

ESSAY

36. **10 pts**

Trace the idea of the atom from Dalton (1808) to the nucleus and the periodic table. Name at least four scientists and the evidence each contributed, and explain how chemists know about particles too small to see.

ANSWER KEY

1. Title and date, question, hypothesis, procedure, data table with units, results/graph, discussion of error and conclusion.
2. One team used pound-force seconds and another expected newton-seconds; the unit mismatch sent the craft too low into Mars's atmosphere.
3. $12.9\ \text{km} = 1.29 \times 10^{4}\ \text{m}$
4. $\rho = m / V$
5. 258 g
6. Mikhail Tsvet, about 1903; plant pigments (chlorophylls and carotenes).
7. Atoms are neither created nor destroyed, so each kind of atom must appear in equal numbers on both sides.
8. 6.20 g
9. *A New System of Chemical Philosophy*, 1808.
10. 2 : 1, the law of multiple proportions (CO and CO_2)
11. Any two: drug testing, airport explosive detection, dating artifacts, identifying proteins, measuring isotope ratios.
12. 6.940 u
13. A burner with a nearly colorless flame (the Bunsen burner).
14. $6.91 \times 10^{14}\ \text{Hz}$
15. No more than two electrons per orbital, and they must have opposite spins.
16. 2
17. Yes: H 4 = 4, O 2 = 2.
18. $4Al + 3O_2 \rightarrow 2Al_2O_3$
19. Yes (ammonium compounds are soluble).
20. 5.66 mol
21. 27.3%
22. 9.97 g
23. (a) Na; (b) 78.0 g; (c) 93.7%
24. $282\ \text{kPa},\ 2.11 \times 10^{3}\ \text{mmHg}$
25. 223 kPa
26. 68.8 L

27. 398 mmHg
28. 1.10 M
29. $-11.1\ ^\circ\text{C}$
30. Lime is basic and neutralizes the acid.
31. 1.6×10^{-10} M
32. $34.9\ ^\circ\text{C}$
33. It lowers the same energy hill, which is crossed in both directions.
34. Both sides have 2 moles of gas.
35. 31.62 years
36. Look for: Dalton (fixed and multiple proportions), Thomson (electron, cathode rays), Rutherford (gold foil, nucleus), Bohr or Moseley (energy levels, atomic number), Mendeleev (periodic law); the point that indirect evidence from repeatable experiments builds a reliable model; clear organization.

Printable answer key (PDF, for the parent)