Last updated: November 1, 2026

🗺️ Four-year roadmap · Year 1, Fall

About this course

This free course has 5 units and 3 labs. Those are the real counts, not a longer syllabus cut short.

A check is one number or code result on a course page that a script recomputes before the page is published. The printed number and the recomputed number have to match.

> This is a free self-study course outline written by Saving Planets. It is not a university course, it carries no academic credit, and no school is involved.
>
> Its scope follows the first semester of college general chemistry as taught in open materials such as OpenStax Chemistry 2e and MIT OpenCourseWare, which we link to throughout. Those organizations do not endorse this page.
>
> Work through it in order. Each unit has goals, key ideas, equations, step-by-step examples, practice with answers, common mistakes, flashcards and a quiz. Your progress is saved only in this browser.

How this course helps save the planet

Chemistry sits behind many of the solutions on Saving Planets: measuring greenhouse gases from orbit, testing water, turning CO_2 into stone, and building cleaner fuels and rockets.

Each unit ends with a short, sourced section showing how its ideas help both people and nature.

For instructors

You can teach this course, or let a student work through it alone. It is free. Saving Planets does not give a grade or school credit.

- This page is the whole course: 5 units and 3 labs (CHEM 110 + CHEM 110L).
- Plan on 15 weeks, about 6 to 9 hours a week, including the lab weeks.
- Follow the 15-week schedule. Read the unit, do the practice on paper, then open the answer. In a lab week, do the lab or the home version.
- A check is one number or code result on the page that a script recomputes before we publish. The number you see and the recomputed number have to match.

- Use Print this page, or download the PDF study guide. The formula sheet is near the end.
- Ask: how does this unit help people and the planet? Each unit has a green box with a sourced fact.
- Copy the teacher guide template and write in your class name, the time you have, and the questions you will ask. Open the teacher guide template

Before you start

Prerequisites

- Algebra: solving for one unknown, rearranging formulas, ratios and percentages.
- Scientific notation and logarithms (base 10). You will use log for pH and absorbance.
- Units: converting grams to kilograms, mL to L, °C to K.
- A scientific calculator. A phone calculator in scientific mode works.

How to study

- Read the unit, then close it and try the worked examples yourself before checking each step.
- Do every practice problem on paper before opening the answer. Getting it wrong first and fixing it is how the method sticks.
- Always carry units through a calculation. If the units do not cancel to the answer you want, the setup is wrong.
- Check significant figures: your answer should not be more precise than your least precise measurement.
- Use the flashcards for 10 minutes a day (spaced practice beats one long cram).
- Take the unit quiz only when you think you are ready. Below 80 percent, review the misconceptions and try again tomorrow.
- Explain one idea out loud each day, as if teaching it. If you get stuck, that is the part to review.

Lab equipment (for the wet labs)

- Digital balance (to 0.001 g)
- pH meter or probe
- Spectrophotometer or colorimeter with cuvettes
- Burette, volumetric pipets and flasks
- Crucible, lid, clay triangle and burner
- Safety goggles, gloves and lab coat

Suggested 15-week schedule

About 6 to 9 hours a week: reading, practice problems and, in lab weeks, the lab.

Week	Unit	Topics	Reading
1	Atomic Theory & Quantum Numbers	Atoms, isotopes and average atomic mass	OpenStax 2.2–2.3
2	Atomic Theory & Quantum Numbers	Light, photons and the Bohr model of hydrogen	OpenStax 6.1–6.2
3	Atomic Theory & Quantum Numbers	Quantum numbers, electron configurations, periodic trends	OpenStax 6.3–6.5
4	Stoichiometry	The mole, molar mass and formulas	OpenStax 3.1–3.2
5	Stoichiometry	Balancing equations and reaction stoichiometry · Lab A	OpenStax 4.1, 4.3
6	Stoichiometry	Limiting reactant, yield, molarity and dilution · Lab B	OpenStax 3.3, 4.4–4.5
7	Molecular Geometry & Bonding	Ionic and covalent bonds, electronegativity, Lewis structures	OpenStax 7.1–7.4
8	Molecular Geometry & Bonding	VSEPR shapes and molecular polarity	OpenStax 7.6
9	Molecular Geometry & Bonding	Intermolecular forces · Midterm practice set	OpenStax 10.1
10	States of Matter: Gases & Solutions	Gas laws and the ideal gas law	OpenStax 9.1–9.3
11	States of Matter: Gases & Solutions	Gas mixtures, partial pressure, kinetic-molecular theory	OpenStax 9.3–9.5
12	States of Matter: Gases & Solutions	Solutions and gas solubility · Lab C	OpenStax 11.1–11.3
13	Thermodynamics	Energy, heat, specific heat and calorimetry	OpenStax 5.1–5.2
14	Thermodynamics	Enthalpy, Hess’s law, intro to entropy and Gibbs energy	OpenStax 5.3, 16.1–16.4
15	Review	Review, flashcards and the final practice set	All units

Your progress

Tick units and labs as you finish them, and take the quizzes. Progress and quiz scores are saved only in this browser, never sent to us.

Units

Atomic Theory & Quantum Numbers

What atoms are made of, how electrons are arranged, and why each element gives off its own colors of light.

Learning objectives

By the end of this unit you can:

- Find protons, neutrons and electrons from an isotope symbol, and compute an average atomic mass.
- Use $c = \lambda\nu$ and $E = h\nu$ to move between wavelength, frequency and photon energy.
- Find the wavelength of light given off when a hydrogen electron drops between levels.
- List the allowed quantum numbers (n, l, m_l, m_s) and spot impossible sets.
- Write electron configurations and orbital diagrams, including ions and noble-gas shorthand.
- Explain periodic trends in atomic radius and ionization energy.

Key concepts

- An atom has a tiny nucleus of protons and neutrons with electrons around it. The atomic number Z counts protons; the mass number A counts protons plus neutrons.
- Isotopes are atoms of the same element with different numbers of neutrons. The atomic mass on the periodic table is a weighted average of the isotopes found in nature.
- Light behaves as a wave (wavelength λ, frequency ν) and as particles called photons, each carrying energy $E = h\nu$.
- Electrons in an atom can only have certain energies. When an electron drops to a lower level it gives off a photon whose energy equals the gap, so every element has its own line spectrum.
- Four quantum numbers describe each electron: n (shell, 1, 2, 3…), l (subshell shape, 0 to n − 1, named s, p, d, f), m_l (orientation, −l to +l) and m_s (spin, +½ or −½).
- Pauli exclusion principle: no two electrons in an atom share all four quantum numbers, so an orbital holds at most two electrons, with opposite spins.
- Aufbau principle: fill the lowest-energy orbitals first (1s, 2s, 2p, 3s, 3p, 4s, 3d, 4p…). Hund’s rule: in equal-energy orbitals, place electrons singly with parallel spins before pairing.
- Across a period, atoms get smaller and ionization energy generally rises because the nuclear charge felt by the outer electrons grows. Down a group, atoms get larger and ionization energy falls.

Key equations

$$c = \lambda\nu$$

Speed of light = wavelength × frequency ($c = 2.998 \times 10^8$ m/s)

$$E = h\nu = \frac{hc}{\lambda}$$

`E=hν=(hc)/()`

Energy of one photon (h = 6.626 × 10⁻³⁴ J·s)

$$E_n = -\frac{k}{n^2}, \qquad k = 2.179 \times 10^{-18}\ \text{J}$$

`E_n=-(k)/(n^2), k=2.179×10^-18\ J`

Bohr energy levels of the hydrogen atom

$$|\Delta E| = k\left(\frac{1}{n_{\text{low}}^2} - \frac{1}{n_{\text{high}}^2}\right)$$

`|Δ E|=k(1n_low^2-1n_high^2)`

Energy of the photon for a jump between two hydrogen levels

$$\bar{m} = \sum_i (\text{fraction}_i)(m_i)$$

`=_i (fraction_i)(m_i)`

Average atomic mass from isotope abundances

Worked examples, step by step

Example 1.1 What wavelength of light does a hydrogen atom give off when its electron drops from n = 3 to n = 2?

1. Energy gap: |ΔE| = k(1/2² − 1/3²) = 2.179 × 10⁻¹⁸ J × (1/4 − 1/9) = 2.179 × 10⁻¹⁸ J × 5/36 = 3.026 × 10⁻¹⁹ J.
2. Rearrange E = hc/λ to λ = hc/E.
3. λ = (6.626 × 10⁻³⁴ J·s × 2.998 × 10⁸ m/s) ÷ 3.026 × 10⁻¹⁹ J = 6.564 × 10⁻⁷ m.
4. Convert: 6.564 × 10⁻⁷ m = 656 nm.

Answer: 656 nm, red light. This is the bright red line in the hydrogen spectrum that astronomers use to spot hydrogen gas in space.

Example 1.2 Write the electron configuration of iron (Fe, Z = 26) and of the Fe²⁺ ion. How many unpaired electrons does Fe have?

1. Fill in order with 26 electrons: 1s² 2s² 2p⁶ 3s² 3p⁶ 4s² 3d⁶.

2. Shorthand: argon has 18 electrons, so Fe = [Ar] $4s^2$ $3d^6$.
3. Hund's rule for $3d^6$: five d orbitals get one electron each, then one pairs up. That leaves 4 unpaired electrons.
4. Transition metals lose their 4s electrons first when they form ions, so Fe^{2+} = [Ar] $3d^6$.

Answer: Fe = [Ar] $4s^2$ $3d^6$ with 4 unpaired electrons; Fe^{2+} = [Ar] $3d^6$.

Example 1.3 Chlorine is 75.76% chlorine-35 (34.969 u) and 24.24% chlorine-37 (36.966 u). What is its average atomic mass?

1. Convert percentages to fractions: 0.7576 and 0.2424.
2. Multiply each mass by its fraction: 0.7576 × 34.969 u = 26.49 u; 0.2424 × 36.966 u = 8.961 u.
3. Add: 26.49 u + 8.961 u = 35.45 u.

Answer: 35.45 u, which matches the periodic table value for chlorine.

Practice problems

Work each on paper first, then check.

1. How many protons, neutrons and electrons are in one atom of carbon-14?

   6 protons, 8 neutrons (14 − 6) and 6 electrons.

2. A green laser has a wavelength of 532 nm. Find its frequency and the energy of one photon.

   $\nu = c/\lambda = (2.998 \times 10^8\ m/s) \div (532 \times 10^{-9}\ m) = 5.64 \times 10^{14}$ Hz. $E = h\nu = 6.626 \times 10^{-34}\ J{\cdot}s \times 5.64 \times 10^{14}\ Hz = 3.73 \times 10^{-19}$ J.

3. List the allowed values of m_l when $l = 2$.

   −2, −1, 0, +1, +2 (five d orbitals).

4. Which set of quantum numbers is impossible: (a) $n = 2, l = 1, m_l = -1$; (b) $n = 2, l = 2, m_l = 0$; (c) $n = 3, l = 0, m_l = 0$?

   (b). l must be less than n, so n = 2 allows only l = 0 or 1.

5. What is the maximum number of electrons in the n = 3 shell?

   $2n^2 = 2 \times 3^2 = 18$ (3s holds 2, 3p holds 6, 3d holds 10).

6. Write the configuration of sulfur (Z = 16) and count its unpaired electrons.

   [Ne] $3s^2$ $3p^4$. The four 3p electrons fill as ↑↓ ↑ ↑, so 2 are unpaired.

7. Which is larger, a sodium atom or a chlorine atom? Why?

Sodium. Both are in period 3, but chlorine has 6 more protons pulling the same shell of electrons closer.

Common misconceptions

Myth: Electrons orbit the nucleus in neat circles, like planets.

Actually: An orbital is a region where an electron is likely to be found, not a path. The Bohr model's circles work only for hydrogen's energy levels.

Myth: The 4s electrons stay put when a transition metal forms an ion, because 4s filled first.

Actually: Transition metals lose 4s electrons before 3d ones. Fe^{2+} is [Ar] $3d^6$, not [Ar] $4s^2$ $3d^4$.

Myth: Higher-frequency light has a longer wavelength.

Actually: Frequency and wavelength are inversely related ($c = \lambda\nu$). Higher frequency means shorter wavelength and more energy per photon.

Myth: The atomic mass on the periodic table is the mass of every atom of that element.

Actually: The table shows a weighted average of the natural isotopes. Individual atoms have the mass of their own isotope.

Flashcards

Atomic number (Z)

Number of protons in the nucleus; it defines the element.

Isotope

Atoms of one element with different numbers of neutrons.

E = hν

Energy of one photon equals Planck's constant times its frequency.

Pauli exclusion principle

No two electrons in an atom have the same four quantum numbers; max two per orbital.

Hund's rule

Fill equal-energy orbitals singly, spins parallel, before pairing.

Allowed l values

0 up to n − 1 (s = 0, p = 1, d = 2, f = 3).

Ionization energy

Line spectrum

Energy needed to remove an electron from a gaseous atom; rises across a period.

The set of wavelengths an element gives off; a fingerprint used to identify it.

Self-check quiz

1. Which photon carries the most energy?

Red light, 700 nm

Green light, 530 nm

Ultraviolet light, 250 nm

Infrared light, 1,000 nm

Ultraviolet light, 250 nm. $E = hc/\lambda$, so the shortest wavelength (250 nm) has the most energy.

2. How many orbitals are in a p subshell?

1

3

5

7

3. $l = 1$ gives $m_l = -1, 0, +1$: three orbitals.

3. What is the ground-state configuration of oxygen (Z = 8)?

$1s^2\ 2s^2\ 2p^4$

$1s^2\ 2s^4\ 2p^2$

$1s^2\ 2p^6$

$1s^2\ 2s^2\ 2p^6$

$1s^2\ 2s^2\ 2p^4$. 2 + 2 + 4 = 8 electrons, filling 1s, then 2s, then 2p.

4. Across period 3 from Na to Ar, first ionization energy generally...

decreases

increases

stays the same

drops to zero at Ar

increases. Nuclear charge grows while the shell stays the same, so electrons are held more tightly.

5. Which quantum number set is allowed?

$n = 1, l = 1$

$n = 3, l = 2, m_l = -3$

$n = 4, l = 3, m_l = +2$

$n = 2, l = 0, m_l = +1$

$n = 4, l = 3, m_l = +2$. For $l = 3$, m_l can be -3 to $+3$, and $l = 3$ is less than $n = 4$.

Try this quiz again

How this helps save the planet

Every gas absorbs and gives off light at its own wavelengths, the idea behind atomic line spectra. Satellite spectrometers use those fingerprints to measure gases from orbit.

NASA's EMIT instrument on the International Space Station uses imaging spectroscopy to find large methane plumes, which helps people find and fix leaks. That protects the climate for people and nature alike.

NASA's OCO-2 satellite measures carbon dioxide by how much sunlight CO_2 absorbs at specific wavelengths.

On Saving Planets: EMIT: methane from orbit · OCO-2: measuring CO_2

Sources: NASA Earth science: climate change (https://science.nasa.gov/climate-change/)

Free learning resources

- OpenStax Chemistry 2e, 6.3 Development of Quantum Theory (https://openstax.org/books/chemistry-2e/pages/6-3-development-of-quantum-theory): Free textbook section with examples and end-of-chapter problems.
- OpenStax Chemistry 2e, 6.4 Electron Configurations (https://openstax.org/books/chemistry-2e/pages/6-4-electronic-structure-of-atoms-electron-configurations): Aufbau, Hund's rule and exceptions.
- PhET: Build an Atom (https://phet.colorado.edu/en/simulations/build-an-atom): Free simulation: add protons, neutrons and electrons.
- PhET: Models of the Hydrogen Atom (https://phet.colorado.edu/en/simulations/models-of-the-hydrogen-atom): Compare the Bohr and quantum models.
- MIT OpenCourseWare 5.111SC Principles of Chemical Science (https://ocw.mit.edu/courses/5-111sc-principles-of-chemical-science-fall-2014/): Free video lectures and problem sets.

☐ **I finished this unit**

Stoichiometry

The math of chemical reactions: counting atoms by weighing, balancing equations, and predicting how much product you get. You will use this later for combustion and rocket propulsion.

Learning objectives

By the end of this unit you can:

- Convert between mass, moles and number of particles using molar mass and Avogadro's number.
- Balance chemical equations by inspection, including combustion reactions.
- Use mole ratios to find the mass of a product or reactant.
- Identify the limiting reactant and compute theoretical and percent yield.
- Compute molarity and solve dilution problems.

Key concepts

- A mole is 6.02214076 × 10^{23} particles (Avogadro's number, exact by definition). The molar mass in g/mol equals the formula mass in atomic mass units.
- A balanced equation has the same number of each kind of atom on both sides. Change only the coefficients, never the subscripts.
- Coefficients give mole ratios, not mass ratios. Always convert grams to moles before using them.
- The limiting reactant runs out first and sets the maximum amount of product (the theoretical yield). Percent yield compares what you actually got with that maximum.
- Molarity (M) is moles of solute per liter of solution. Diluting keeps the moles of solute the same, so $M_1V_1 = M_2V_2$.
- In complete combustion, a hydrocarbon reacts with oxygen to give carbon dioxide and water. The oxygen comes from the air, which is why the CO_2 produced weighs more than the fuel burned.

Key equations

$$n = \frac{m}{M}$$

`n=m/M`

Moles = mass ÷ molar mass

$$N = n\,N_A, \qquad N_A = 6.02214076 \times 10^{23}\ \mathrm{mol^{-1}}$$

`N=n N_A, N_A=6.02214076×10^23\ mol^-1`

Number of particles

$$\text{percent yield} = \frac{\text{actual yield}}{\text{theoretical yield}} \times 100\%$$

```
percent yield= yieldtheoretical yield×100\%
```

Percent yield

$$M = \frac{n_{\text{solute}}}{V_{\text{solution}}\,(\text{L})}$$

```
M=_soluteV_solution\ (L)
```

Molarity

$$M_1 V_1 = M_2 V_2$$

```
M_1V_1=M_2V_2
```

Dilution

Worked examples, step by step

Example 2.1 Methane burns: $CH_4 + O_2 \rightarrow CO_2 + H_2O$. Balance it, then find the mass of CO_2 made by burning 1.00 kg of methane.

1. Balance C: 1 CH_4 gives 1 CO_2. Balance H: 4 H need 2 H_2O. Balance O: 2 (in CO_2) + 2 (in 2 H_2O) = 4 O atoms, so 2 O_2.
2. Balanced: $CH_4 + 2\ O_2 \rightarrow CO_2 + 2\ H_2O$.
3. Moles of CH_4: 1000 g ÷ 16.04 g/mol = 62.34 mol.
4. Mole ratio CO_2 : CH_4 = 1 : 1, so 62.34 mol CO_2.
5. Mass of CO_2: 62.34 mol × 44.01 g/mol = 2,744 g.

Answer: About 2.74 kg of CO_2 from 1.00 kg of methane. The extra mass is oxygen taken from the air.

Example 2.2 Hydrogen–oxygen rocket engines run on $2\ H_2 + O_2 \rightarrow 2\ H_2O$. If 10.0 g of H_2 reacts with 64.0 g of O_2, which is limiting, how much water forms, and what is left over?

1. Moles: H_2 = 10.0 g ÷ 2.016 g/mol = 4.960 mol; O_2 = 64.0 g ÷ 32.00 g/mol = 2.000 mol.
2. To use all 2.000 mol O_2 you need 2 × 2.000 = 4.000 mol H_2. You have 4.960 mol, so O_2 runs out first: O_2 is limiting.
3. Water: 2 mol H_2O per mol O_2, so 4.000 mol H_2O × 18.02 g/mol = 72.1 g.
4. Left over: 4.960 − 4.000 = 0.960 mol H_2 × 2.016 g/mol = 1.94 g H_2.
5. Check: 72.1 g + 1.94 g = 74.0 g, the same as 10.0 g + 64.0 g. Mass is conserved.

Answer: O_2 is limiting; 72.1 g of water forms and 1.94 g of H_2 is left over.

Example 2.3 How do you make 250.0 mL of 0.100 M NaOH from a 1.00 M stock solution?

1. Use $M_1V_1 = M_2V_2$ with M_1 = 1.00 M, M_2 = 0.100 M, V_2 = 250.0 mL.
2. V_1 = (0.100 M × 250.0 mL) ÷ 1.00 M = 25.0 mL.
3. Measure 25.0 mL of stock with a volumetric pipet into a 250.0 mL volumetric flask and add water to the mark.

Answer: 25.0 mL of stock, diluted to 250.0 mL.

Practice problems

Work each on paper first, then check.

1. Balance: $C_3H_8 + O_2 \rightarrow CO_2 + H_2O$ (propane combustion).

   $C_3H_8 + 5\ O_2 \rightarrow 3\ CO_2 + 4\ H_2O$.

2. Balance the combustion of octane, C_8H_{18}.

   $2\ C_8H_{18} + 25\ O_2 \rightarrow 16\ CO_2 + 18\ H_2O$.

3. What is the molar mass of octane, C_8H_{18}?

   8 × 12.011 + 18 × 1.008 = 114.23 g/mol.

4. How many molecules are in 18.02 g of water?

   18.02 g ÷ 18.02 g/mol = 1.000 mol, so 6.022×10^{23} molecules.

5. What is the molarity of 5.844 g of NaCl dissolved to make 250.0 mL of solution?

   5.844 g ÷ 58.44 g/mol = 0.1000 mol; 0.1000 mol ÷ 0.2500 L = 0.4000 M.

6. 25.0 mL of 6.00 M HCl is diluted to 150.0 mL. What is the new concentration?

   M_2 = (6.00 M × 25.0 mL) ÷ 150.0 mL = 1.00 M.

7. A reaction's theoretical yield is 72.1 g, but you collect 65.0 g. What is the percent yield?

   65.0 ÷ 72.1 × 100% = 90.2%.

Common misconceptions

Myth: The reactant with the smaller mass is the limiting reactant.

Actually: Compare moles against the mole ratio. In the rocket example, H_2 had less mass but O_2 was limiting.

Myth: You can balance an equation by changing subscripts, like writing H_2O_2 instead of H_2O.

Actually: Subscripts define the substance. Changing them makes a different chemical. Change only coefficients.

Myth: Burning 1 kg of fuel cannot make more than 1 kg of products.

Actually: Oxygen from the air joins the fuel, so the products weigh more than the fuel alone (mass is still conserved overall).

Myth: Molarity uses the volume of water added.

Actually: Molarity uses the final volume of the solution, which is why you fill a volumetric flask to the mark.

Flashcards

Mole

$6.02214076 \times 10^{23}$ particles (exact).

Molar mass

Mass of one mole of a substance, in g/mol.

Limiting reactant

The reactant used up first; it sets the theoretical yield.

Theoretical yield

Maximum product predicted from the limiting reactant.

Molarity

Moles of solute per liter of solution (mol/L).

Dilution

$M_1V_1 = M_2V_2$; moles of solute stay the same.

Complete combustion of a hydrocarbon

Fuel + $O_2 \rightarrow CO_2 + H_2O$.

Mole ratio

Ratio of coefficients in a balanced equation.

Self-check quiz

1. How many moles are in 88.0 g of CO_2 (44.01 g/mol)?

0.500 mol

1.00 mol

2.00 mol

3.87×10^3 mol

2.00 mol. 88.0 g ÷ 44.01 g/mol = 2.00 mol.

2. In $C_3H_8 + 5\ O_2 \rightarrow 3\ CO_2 + 4\ H_2O$, how many moles of O_2 react with 2 mol of C_3H_8?

5

10

2.5

7

10. The ratio is 5 O_2 per C_3H_8, so 2 × 5 = 10 mol.

3. What does $M_1V_1 = M_2V_2$ keep constant?

Volume

Concentration

Moles of solute

Mass of water

Moles of solute. Adding solvent changes volume and concentration, but not the moles of solute.

4. The theoretical yield is set by...

the excess reactant

the limiting reactant

the catalyst

the heaviest product

the limiting reactant. Product stops forming when the limiting reactant is used up.

5. What is the molar mass of H_2O?

10.0 g/mol

16.00 g/mol

17.01 g/mol

18.02 g/mol

18.02 g/mol. 2 × 1.008 + 16.00 = 18.02 g/mol.

Try this quiz again

How this helps save the planet

Stoichiometry is how engineers and climate scientists turn fuel use into emissions. As the worked example shows, burning 1 kg of methane releases about 2.74 kg of CO_2.

The same math sizes rocket propellant loads, compares cleaner fuels, and checks how much CO_2 a capture project removes.

The U.S. Environmental Protection Agency notes that carbon dioxide enters the atmosphere mostly through burning fossil fuels, solid waste and wood. Counting it accurately is the first step to cutting it, which protects both people and ecosystems.

On Saving Planets: Carbfix: turning CO_2 into stone · Artemis Program

Sources: U.S. EPA: Overview of Greenhouse Gases (https://www.epa.gov/ghgemissions/overview-greenhouse-gases) · NASA Glenn: Rocket thrust (for students) (https://www.grc.nasa.gov/www/k-12/airplane/rockth.html)

Free learning resources

- OpenStax Chemistry 2e, Chapter 3: Composition of Substances and Solutions (https://openstax.org/books/chemistry-2e/pages/3-introduction): Moles, formulas and molarity.
- OpenStax Chemistry 2e, 4.3 Reaction Stoichiometry (https://openstax.org/books/chemistry-2e/pages/4-3-reaction-stoichiometry): Mole ratios and yields with examples.
- PhET: Balancing Chemical Equations (https://phet.colorado.edu/en/simulations/balancing-chemical-equations): Practice balancing with instant feedback.
- PhET: Reactants, Products and Leftovers (https://phet.colorado.edu/en/simulations/reactants-products-and-leftovers): See limiting reactants visually.

- Khan Academy: Chemistry (https://www.khanacademy.org/science/chemistry): Free videos and practice on stoichiometry.

☐ **I finished this unit**

Molecular Geometry & Bonding

Why atoms bond, how to draw molecules, what shapes they take, and the forces between molecules that decide whether something is a gas, liquid or solid.

Learning objectives

By the end of this unit you can:

- Classify bonds as ionic, polar covalent or nonpolar covalent using electronegativity.
- Draw Lewis structures and check them with formal charge.
- Predict electron-pair and molecular geometry with VSEPR theory.
- Decide whether a molecule is polar from its bonds and shape.
- Rank substances by boiling point using intermolecular forces.

Key concepts

- Ionic bonds form when electrons transfer between atoms with very different electronegativities (for example Na and Cl). Covalent bonds share electron pairs.
- Electronegativity is an atom's pull on shared electrons. Unequal sharing makes a polar bond with partial charges δ+ and δ−.
- Lewis structures show valence electrons as bonds and lone pairs. Most main-group atoms end with eight electrons around them (the octet rule); hydrogen ends with two.
- Formal charge helps choose the best Lewis structure: the best one has formal charges as close to zero as possible.
- VSEPR: electron groups (bonds and lone pairs) around a central atom spread out as far as possible. 2 groups: linear, 180°. 3: trigonal planar, 120°. 4: tetrahedral, 109.5°. 5: trigonal bipyramidal. 6: octahedral.
- Lone pairs take up more room than bonds, so they squeeze bond angles: water is bent at about 104.5°, ammonia is trigonal pyramidal at about 107°.
- A molecule is polar if its bond dipoles do not cancel. CO_2 has polar bonds but is linear, so it is nonpolar overall; H_2O is bent, so it is polar.
- Intermolecular forces, weakest to strongest for similar sizes: London dispersion (all molecules, larger with more electrons), dipole–dipole (polar molecules), hydrogen bonding (H on N, O or F). Stronger forces mean higher boiling points.

Key equations

$$\mathrm{FC} = V - N - \frac{B}{2}$$

```
FC=V-N-B/2
```

Formal charge: V = valence electrons, N = nonbonding electrons, B = bonding electrons

$$\Delta\chi = |\chi_A - \chi_B|$$

Electronegativity difference; larger means a more polar bond

Worked examples, step by step

Example 3.1 Draw the Lewis structure of CO_2 and predict its shape and polarity.

1. Count valence electrons: C has 4, each O has 6: 4 + 2 × 6 = 16.
2. Put C in the center and connect each O with a double bond: O=C=O uses 8 electrons.
3. Place the remaining 8 as two lone pairs on each O. Every atom now has an octet.
4. Formal charges: C = 4 − 0 − 8/2 = 0; each O = 6 − 4 − 4/2 = 0. A good structure.
5. Two electron groups on C and no lone pairs: linear, 180°. The two C=O dipoles point in opposite directions and cancel.

Answer: O=C=O, linear and nonpolar overall, even though each bond is polar.

Example 3.2 Predict the shape, bond angle and polarity of water.

1. Valence electrons: 2 × 1 + 6 = 8.
2. Two O–H bonds use 4 electrons; the other 4 are two lone pairs on O.
3. Four electron groups: tetrahedral electron geometry. With two lone pairs, the molecular shape is bent.
4. Lone pairs push the bonds closer: angle about 104.5°, less than 109.5°.
5. The bond dipoles do not cancel in a bent shape, so water is polar and can hydrogen bond.

Answer: Bent, about 104.5°, polar.

Example 3.3 Which has the higher boiling point: methane (CH_4) or octane (C_8H_{18})? Why?

1. Both are nonpolar hydrocarbons, so the main force is London dispersion.
2. Octane has many more electrons and a longer shape, so its dispersion forces are much stronger.

Answer: Octane. That is why octane is a liquid at room temperature while methane is a gas.

Practice problems

Work each on paper first, then check.

1. Predict the shape and bond angle of NH_3.

   Four electron groups with one lone pair: trigonal pyramidal, about 107°.

2. Predict the shape of CH_4 and say whether it is polar.

   Tetrahedral, 109.5°. The four C–H dipoles cancel, so it is nonpolar.

3. Is BF_3 polar?

   No. It is trigonal planar with three identical B–F bonds whose dipoles cancel.

4. What is the shape of SF_6?

   Octahedral (six bonding groups, no lone pairs on S), 90° bond angles.

5. Find the formal charge on N in the ammonium ion, NH_4^+.

   FC = 5 − 0 − 8/2 = +1.

6. Why does H_2O boil at 100 °C while the heavier H_2S boils far lower?

   Water molecules hydrogen bond to each other; H_2S cannot hydrogen bond as strongly, so less energy is needed to separate its molecules.

7. Is SO_2 polar or nonpolar?

   Polar. Sulfur has a lone pair, so SO_2 is bent and its bond dipoles do not cancel.

Common misconceptions

Myth: A molecule with polar bonds must be polar.
Actually: Shape matters. CO_2 and CCl_4 have polar bonds but symmetric shapes, so they are nonpolar.

Myth: Boiling breaks the bonds inside molecules.
Actually: Boiling water separates H_2O molecules from each other. The O–H bonds inside each molecule stay intact.

Myth: Lone pairs do not affect shape because you cannot see them in the formula.
Actually: Lone pairs repel more strongly than bonds and bend the shape, as in H_2O and NH_3.

Myth: Hydrogen bonds are bonds to hydrogen atoms in any molecule.
Actually: They need H bonded to N, O or F, attracted to a lone pair on another N, O or F.

Flashcards

Electronegativity

Octet rule

How strongly an atom pulls shared electrons toward itself.

Main-group atoms tend to end with eight valence electrons.

Formal charge

FC = V − N − B/2; best structures keep it near zero.

VSEPR, 4 groups

Tetrahedral, 109.5°.

Water's shape

Bent, about 104.5°.

Hydrogen bond

Attraction between H on N/O/F and a lone pair on another N/O/F.

London dispersion

Weak attraction from temporary dipoles; grows with more electrons.

Polar molecule

Bond dipoles do not cancel because of its shape.

Self-check quiz

1. What is the molecular shape of a central atom with 3 bonds and 1 lone pair?

Trigonal planar

Trigonal pyramidal

Bent

Tetrahedral

Trigonal pyramidal. Four electron groups, one a lone pair: trigonal pyramidal, like NH_3.

2. Which molecule is nonpolar?

H_2O

NH_3

CO_2

SO_2

CO_2. Linear CO_2 has opposite bond dipoles that cancel.

3. Which intermolecular force is present in all molecules?

Hydrogen bonding

Dipole–dipole

London dispersion

Ion–dipole

London dispersion. Every molecule has electrons that can form temporary dipoles.

4. What is the ideal bond angle in a tetrahedral molecule?

90°

109.5°

120°

180°

109.5°. Four groups spread out in 3D at 109.5°.

5. How many valence electrons are in the Lewis structure of CO_2?

12

14

16

22

16. 4 from C plus 6 from each O.

Try this quiz again

How this helps save the planet

Molecular shape explains the greenhouse effect. Molecules such as CO_2, CH_4 and H_2O can bend or stretch in ways that change their dipole, so they absorb infrared heat radiating from Earth. N_2 and O_2, which make up most of the air, cannot.

Hydrogen bonding makes ice less dense than liquid water, so ice floats. Lakes freeze from the top down, and fish and other life survive the winter in the water below.

Intermolecular forces also guide cleanups: oil is nonpolar and does not mix with water, which is why spills spread as surface slicks.

On Saving Planets: GOSAT-GW: greenhouse gases from space · Lakes hub

Sources: U.S. EPA: Overview of Greenhouse Gases (https://www.epa.gov/ghgemissions/overview-greenhouse-gases)

Free learning resources

- OpenStax Chemistry 2e, Chapter 7: Chemical Bonding and Molecular Geometry (https://openstax.org/books/chemistry-2e/pages/7-introduction): Lewis structures, formal charge and VSEPR.
- OpenStax Chemistry 2e, 7.6 Molecular Structure and Polarity (https://openstax.org/books/chemistry-2e/pages/7-6-molecular-structure-and-polarity): Shapes and dipoles with worked examples.
- OpenStax Chemistry 2e, 10.1 Intermolecular Forces (https://openstax.org/books/chemistry-2e/pages/10-1-intermolecular-forces): Dispersion, dipole–dipole and hydrogen bonding.
- PhET: Molecule Shapes (https://phet.colorado.edu/en/simulations/molecule-shapes): Build molecules in 3D and watch lone pairs change the shape.
- PhET: Molecule Polarity (https://phet.colorado.edu/en/simulations/molecule-polarity): Change electronegativity and see bond and molecular dipoles.

☐ **I finished this unit**

States of Matter: Gases & Solutions

How gases respond to pressure and temperature, how gas mixtures work, and how substances (including oxygen and CO_2) dissolve in water.

Learning objectives

By the end of this unit you can:

- Use Boyle's, Charles's and the combined gas law, always in kelvin.
- Solve problems with the ideal gas law, PV = nRT, including molar mass from density.
- Use Dalton's law of partial pressures.
- Explain gas behavior with kinetic-molecular theory.
- Describe solubility, and use Henry's law for gases dissolving in liquids.

Key concepts

- Gas laws need absolute temperature: T(K) = T(°C) + 273.15.
- At constant temperature, pressure and volume are inversely related (Boyle). At constant pressure, volume is proportional to kelvin temperature (Charles).
- The ideal gas law, PV = nRT, ties pressure, volume, moles and temperature together. At 0 °C and 1 atm, one mole of an ideal gas takes up 22.4 L.
- In a gas mixture, each gas exerts its own partial pressure, and they add up to the total (Dalton's law). Dry air is about 20.95% oxygen by volume.
- Kinetic-molecular theory: gas particles move fast and randomly, collide elastically, and their average kinetic energy is proportional to kelvin temperature.
- A solution is a uniform mixture. "Like dissolves like": polar solvents such as water dissolve polar and ionic substances.
- The amount of a gas that dissolves in a liquid rises with that gas's partial pressure (Henry's law) and usually falls as the water warms.

Key equations

$$PV = nRT$$

`PV=nRT`

Ideal gas law

$$R = 0.08206\ \frac{\text{L·atm}}{\text{mol·K}} = 8.314\ \frac{\text{J}}{\text{mol·K}}$$

`R=0.08206\ ·atmmol·K=8.314\ ·K`

Gas constant

$$\frac{P_1 V_1}{T_1} = \frac{P_2 V_2}{T_2}$$

`(P_1V_1)/(T_1)=(P_2V_2)/(T_2)`

Combined gas law (fixed amount of gas)

$$P_{\text{total}} = P_A + P_B + \ldots, \qquad P_A = x_A \, P_{\text{total}}$$

`P_total=P_A+P_B+, P_A=x_A P_total`

Dalton’s law; x is the mole fraction

$$M = \frac{dRT}{P}$$

`M=dRT/P`

Molar mass from gas density d

$$C_g = k_H \, P_g$$

`C_g=k_H P_g`

Henry’s law: dissolved gas concentration

Worked examples, step by step

Example 4.1 What volume does 1.00 mol of an ideal gas occupy at 25.0 °C and 1.00 atm?

1. Convert temperature: 25.0 + 273.15 = 298.2 K.
2. V = nRT/P = 1.00 mol × 0.08206 L·atm/(mol·K) × 298.2 K ÷ 1.00 atm.
3. V = 24.5 L.

Answer: 24.5 L, a bit more than the 22.4 L at 0 °C because warmer gas expands.

Example 4.2 A weather balloon holds 2.00 L of helium at 1.00 atm and 300 K. It rises to where the pressure is 0.500 atm and the temperature is 250 K. What is its new volume?

1. Combined gas law solved for V_2: $V_2 = V_1 \times (P_1/P_2) \times (T_2/T_1)$.
2. V_2 = 2.00 L × (1.00/0.500) × (250/300).
3. V_2 = 2.00 L × 2 × 0.8333 = 3.33 L.

Answer: 3.33 L. Lower pressure lets it expand more than the cold shrinks it.

Example 4.3 Estimate how much oxygen dissolves in water at 25 °C open to air at 1.00 atm. Henry's constant for O_2 is about 1.3×10^{-3} mol/(L·atm).

1. Partial pressure of O_2: 0.2095 × 1.00 atm = 0.21 atm.
2. C = k_H × P = 1.3×10^{-3} mol/(L·atm) × 0.21 atm = 2.7×10^{-4} mol/L.
3. Convert to mass: 2.7×10^{-4} mol/L × 32.00 g/mol = 8.7×10^{-3} g/L = 8.7 mg/L.

Answer: About 9 mg of oxygen per liter: a small amount, and all that fish and other aquatic life have to breathe.

Practice problems

Work each on paper first, then check.

1. Convert 25 °C and −40 °C to kelvin.

   298 K and 233 K (add 273.15).

2. How many moles of gas fill 10.0 L at 0 °C (273.15 K) and 1.00 atm?

   n = PV/RT = (1.00 × 10.0) ÷ (0.08206 × 273.15) = 0.446 mol.

3. At constant temperature, what happens to a gas volume if the pressure doubles?

   It halves (Boyle's law).

4. A gas has a density of 1.25 g/L at 0 °C and 1.00 atm. Find its molar mass.

   M = dRT/P = 1.25 × 0.08206 × 273.15 ÷ 1.00 = 28.0 g/mol (consistent with N_2).

5. A tank holds 2.0 mol N_2 and 0.50 mol O_2 at 5.0 atm total. What is the O_2 partial pressure?

   $x(O_2)$ = 0.50 ÷ 2.5 = 0.20; $P(O_2)$ = 0.20 × 5.0 atm = 1.0 atm.

6. Why does a can of soda go flat faster when warm?

   Gas solubility falls as temperature rises, so dissolved CO_2 escapes more quickly.

Common misconceptions

Myth: You can use °C in gas laws.

Actually: Always use kelvin. At 0 °C the volume is not zero; at 0 K it would be.

Myth: 22.4 L/mol is true for any gas at any conditions.

Actually: It applies only to an ideal gas at 0 °C and 1 atm. At 25 °C it is about 24.5 L.

Myth: Heavier gas molecules move faster at the same temperature.

Actually: At the same temperature all gases have the same average kinetic energy, so heavier molecules move slower.

Myth: Gases dissolve better in warm water, like sugar does.

Actually: Most gases are less soluble in warmer water.

Flashcards

Ideal gas law

PV = nRT.

R

0.08206 L·atm/(mol·K) or 8.314 J/(mol·K).

Boyle's law

$P_1V_1 = P_2V_2$ at constant n and T.

Charles's law

$V_1/T_1 = V_2/T_2$ at constant n and P.

Dalton's law

Total pressure is the sum of the partial pressures.

Henry's law

Dissolved gas concentration = k_H × partial pressure.

Kelvin

T(K) = T(°C) + 273.15.

Molar volume

22.4 L/mol for an ideal gas at 0 °C and 1 atm.

Self-check quiz

1. Which temperature scale must gas-law calculations use?

Celsius

Fahrenheit

Kelvin

Any of these

Kelvin. Gas laws are proportional to absolute temperature.

2. At constant pressure, a gas is heated from 200 K to 400 K. Its volume...

halves

doubles

stays the same

quadruples

doubles. Charles's law: V is proportional to T in kelvin.

3. Air at 1.00 atm is about 21% O_2. The partial pressure of O_2 is about...

0.21 atm

0.79 atm

1.00 atm

21 atm

0.21 atm. $P(O_2) = 0.21 \times 1.00$ atm.

4. What happens to dissolved oxygen as a lake warms?

It rises

It falls

It does not change

It turns into CO_2

It falls. Gas solubility in water usually decreases with temperature.

5. What is R in L·atm/(mol·K)?

8.314

0.08206

22.4

6.022×10^{23}

0.08206. 8.314 is the value in J/(mol·K).

Try this quiz again

How this helps save the planet

Dissolved gases decide what can live in water. Because warm water holds less oxygen, warming lakes and rivers can stress fish, and water-quality teams measure dissolved oxygen to protect them.

Henry's law also explains ocean acidification. As CO_2 in the air rises, more dissolves in seawater and forms carbonic acid. NOAA reports that the ocean absorbs about 30% of the CO_2 released into the atmosphere and that surface ocean pH has fallen by about 0.1 since the industrial revolution, roughly a 30% increase in acidity.

That matters to people as well as reefs: NOAA notes that oysters, corals and other shell builders struggle in more acidic water, and many fisheries and coastal jobs depend on them.

On Saving Planets: Oceans hub · PACE: ocean color and plankton

Sources: NOAA: Ocean acidification (education) (https://www.noaa.gov/education/resource-collections/ocean-coasts/ocean-acidification)

Free learning resources

- OpenStax Chemistry 2e, Chapter 9: Gases (https://openstax.org/books/chemistry-2e/pages/9-introduction): Gas laws, mixtures and kinetic-molecular theory.
- OpenStax Chemistry 2e, Chapter 11: Solutions and Colloids (https://openstax.org/books/chemistry-2e/pages/11-introduction): Solubility and Henry's law.
- PhET: Gas Properties (https://phet.colorado.edu/en/simulations/gas-properties): Pump gas, heat it and watch pressure change.
- NOAA: Ocean acidification resources (https://www.noaa.gov/education/resource-collections/ocean-coasts/ocean-acidification): Lessons, data and labs on CO_2 in seawater.

☐ **I finished this unit**

Thermodynamics

Chemical energy: heat, enthalpy, why some reactions release energy, and a first look at entropy and Gibbs free energy.

Learning objectives

By the end of this unit you can:

- Use q = mcΔT and calorimetry to find heat flow and specific heat.
- Classify processes as exothermic or endothermic and give the sign of ΔH.
- Use Hess's law and standard enthalpies of formation to find ΔH of a reaction.
- State the first law of thermodynamics with the chemistry sign convention.
- Use ΔG = ΔH − TΔS to judge whether a process is spontaneous.

Key concepts

- Energy is conserved (the first law): ΔU = q + w. In chemistry, q is heat added to the system and w is work done on the system.
- Exothermic reactions release heat to the surroundings (ΔH < 0); endothermic reactions absorb heat (ΔH > 0).
- Specific heat c is the heat needed to warm 1 g by 1 °C. Water's is high, 4.184 J/(g·°C), which is why oceans and lakes warm and cool slowly.
- Enthalpy is a state function, so ΔH depends only on start and end. That is why Hess's law works: you can add reactions and add their ΔH values.
- Standard enthalpy of formation (ΔH°f) is the enthalpy change to make 1 mol of a compound from its elements in their standard states. For an element in its standard state, ΔH°f = 0.
- Entropy (S) measures how spread out energy and matter are. The Gibbs free energy change, ΔG = ΔH − TΔS, is negative for a spontaneous process at constant temperature and pressure.

Key equations

$$q = mc\Delta T$$

`q=mcΔ T`

Heat for a temperature change

$$q_{\text{lost}} = -q_{\text{gained}}$$

`q_lost=-q_gained`

Calorimetry: heat lost by the hot object equals heat gained by the cold one

$$\Delta U = q + w, \qquad w = -P\Delta V$$

Δ U=q+w, w=-PΔ V

First law (chemistry sign convention)

$$\Delta H^\circ_{\text{rxn}} = \sum n\,\Delta H^\circ_f(\text{products}) - \sum n\,\Delta H^\circ_f(\text{reactants})$$

Δ H^_rxn= n Δ H^_f(products)- n Δ H^_f(reactants)

Reaction enthalpy from formation enthalpies

$$\Delta G = \Delta H - T\Delta S$$

Δ G=Δ H-TΔ S

Gibbs free energy

Worked examples, step by step

Example 5.1 How much heat does it take to warm 250.0 g of water from 20.0 °C to 100.0 °C?

1. ΔT = 100.0 − 20.0 = 80.0 °C.
2. q = mcΔT = 250.0 g × 4.184 J/(g·°C) × 80.0 °C = 83,680 J.
3. In kilojoules: 83.7 kJ.

Answer: 83.7 kJ.

Example 5.2 Find ΔH° for burning methane, $CH_4(g) + 2\,O_2(g) \rightarrow CO_2(g) + 2\,H_2O(l)$, using ΔH°f values: $CH_4(g)$ −74.6, $CO_2(g)$ −393.5, $H_2O(l)$ −285.8 kJ/mol; $O_2(g)$ 0.

1. Products: −393.5 + 2 × (−285.8) = −965.1 kJ.
2. Reactants: −74.6 + 2 × 0 = −74.6 kJ.
3. ΔH° = −965.1 − (−74.6) = −890.5 kJ.

Answer: About −890 kJ per mole of methane: strongly exothermic.

Example 5.3 Compare the heat released per gram by methane and by hydrogen. For $H_2(g) + ½\,O_2(g) \rightarrow H_2O(l)$, ΔH° = −285.8 kJ.

1. Methane: 890.5 kJ ÷ 16.04 g = 55.5 kJ/g.
2. Hydrogen: 285.8 kJ ÷ 2.016 g = 141.8 kJ/g.

3. Products: methane makes CO_2; hydrogen makes only water.

Answer: Hydrogen releases about 2.6 times more heat per gram, and burning it makes no CO_2. Whether hydrogen is clean overall depends on how it was produced.

Example 5.4 A reaction has $\Delta H = +40.0$ kJ and $\Delta S = +120$ J/K. Is it spontaneous at 298 K? Above what temperature does it become spontaneous?

1. Use the same units: $\Delta S = 0.120$ kJ/K.
2. At 298 K: $\Delta G = 40.0 - 298 \times 0.120 = 40.0 - 35.8 = +4.2$ kJ. Positive, so not spontaneous.
3. Set $\Delta G = 0$: $T = \Delta H / \Delta S = 40.0 \div 0.120 = 333$ K.

Answer: Not spontaneous at 298 K; spontaneous above about 333 K (60 °C).

Practice problems

Work each on paper first, then check.

1. How much heat raises 1.00 kg of water by 10.0 °C?

   $q = 1000 \text{ g} \times 4.184 \times 10.0 = 41{,}840$ J = 41.8 kJ.

2. Is melting ice endothermic or exothermic? What is the sign of q for the ice?

   Endothermic; q is positive because the ice absorbs heat.

3. A 50.0 g metal at 100.0 °C is dropped into 100.0 g of water at 20.0 °C. Both end at 22.0 °C. Find the metal's specific heat.

   Heat gained by water = $100.0 \times 4.184 \times 2.0 = 836.8$ J. Metal: $836.8 = 50.0 \times c \times 78.0$, so $c = 0.215$ J/(g·°C).

4. Use Hess's law: $C(s) + O_2 \rightarrow CO_2$, $\Delta H = -393.5$ kJ; $CO + \frac{1}{2} O_2 \rightarrow CO_2$, $\Delta H = -283.0$ kJ. Find ΔH for $C(s) + \frac{1}{2} O_2 \rightarrow CO$.

   Reverse the second and add: $-393.5 + 283.0 = -110.5$ kJ.

5. What is $\Delta H°f$ for $O_2(g)$?

   Zero: it is an element in its standard state.

6. A process has $\Delta H < 0$ and $\Delta S > 0$. When is it spontaneous?

   At every temperature, because $\Delta G = \Delta H - T\Delta S$ is always negative.

Common misconceptions

Myth: Heat and temperature are the same thing.

Actually: Temperature measures average particle energy; heat is energy transferred. A bathtub of warm water holds more heat than a cup of boiling water.

Myth: Exothermic reactions always happen on their own.

Actually: Spontaneity depends on ΔG, which includes entropy. Some exothermic reactions are not spontaneous at high temperatures.

Myth: A spontaneous reaction is a fast one.

Actually: Spontaneous tells you the direction, not the speed. Diamond turning into graphite is spontaneous but far too slow to notice.

Myth: Breaking bonds releases energy.

Actually: Breaking bonds always takes energy. Energy is released when new, stronger bonds form.

Flashcards

q = mcΔT

Heat = mass × specific heat × temperature change.

Exothermic

Releases heat; ΔH < 0.

Endothermic

Absorbs heat; ΔH > 0.

Hess's law

ΔH of a sum of reactions is the sum of their ΔH values.

ΔH°f of an element

Zero in its standard state.

First law

ΔU = q + w; energy is conserved.

Gibbs energy

ΔG = ΔH − TΔS; negative means spontaneous.

Specific heat of water

4.184 J/(g·°C).

Self-check quiz

1. The sign of ΔH for an exothermic reaction is…

positive

negative

zero

it depends on temperature

negative. The system loses heat to the surroundings.

2. Water's specific heat is...

1.00 J/(g·°C)

4.184 J/(g·°C)

8.314 J/(g·°C)

0.449 J/(g·°C)

4.184 J/(g·°C). 4.184 J warms 1 g of water by 1 °C.

3. Hess's law works because enthalpy is...

always negative

a state function

equal to entropy

measured in kelvin

a state function. ΔH depends only on the start and end states.

4. A reaction is spontaneous when...

$\Delta G < 0$

$\Delta G > 0$

$\Delta H > 0$

$\Delta S = 0$

$\Delta G < 0$. Negative Gibbs energy change means the process can proceed on its own.

5. Per gram, which fuel releases more heat when burned to liquid water?

Methane

Hydrogen

They are equal

Neither releases heat

Hydrogen. About 142 kJ/g for hydrogen vs. 55.5 kJ/g for methane.

Try this quiz again

How this helps save the planet

Thermochemistry is the accounting behind clean energy. Comparing kilojoules per gram and CO_2 per kilojoule, as in the worked examples, shows why hydrogen made with clean power can cut emissions.

The U.S. Department of Energy notes that hydrogen used in a fuel cell produces only water, and that hydrogen can be made from many sources, including solar and wind power.

Water's high specific heat lets oceans store huge amounts of heat, which steadies coastal climates for people and wildlife.

On Saving Planets: Sentinel-6B: tracking sea level · Earth Observation hub

Sources: U.S. DOE: Hydrogen fuel basics (https://www.energy.gov/eere/fuelcells/hydrogen-fuel-basics)

Free learning resources

- OpenStax Chemistry 2e, Chapter 5: Thermochemistry (https://openstax.org/books/chemistry-2e/pages/5-introduction): Calorimetry, enthalpy and Hess's law.
- OpenStax Chemistry 2e, Chapter 16: Thermodynamics (https://openstax.org/books/chemistry-2e/pages/16-introduction): Entropy and Gibbs free energy.
- PhET: Energy Forms and Changes (https://phet.colorado.edu/en/simulations/energy-forms-and-changes): Watch heat flow between objects.
- MIT OpenCourseWare 3.091 Introduction to Solid-State Chemistry (https://ocw.mit.edu/courses/3-091-introduction-to-solid-state-chemistry-fall-2018/): Free lectures that also cover bonding and energy.

☐ **I finished this unit**

Problem Sets

These practice problems live inside each unit. Work on paper first, then open the answer.

1. Atomic Theory & Quantum Numbers
   7 problems
2. Stoichiometry
   7 problems
3. Molecular Geometry & Bonding
   7 problems
4. States of Matter: Gases & Solutions
   6 problems
5. Thermodynamics
   6 problems

The lab (CHEM 110L)

The lab is the hands-on part of the course, a 1-credit component in a typical college schedule. Here it is free and for practice only.

You put lecture ideas into practice with digital sensors, meters and precision glassware. Do the wet-lab versions only in a supervised school or college lab; each exercise also has a safe home or virtual version.

Safety reference: American Chemical Society: Chemical safety (https://www.acs.org/chemical-safety.html)

Lab A: Empirical formula of magnesium oxide

WEEK 5

Goal: Heat magnesium in air, weigh the product, and work out the simplest whole-number ratio of magnesium to oxygen.

$$\frac{n_{\mathrm{Mg}}}{n_{\mathrm{O}}} = \frac{m_{\mathrm{Mg}}/24.31}{m_{\mathrm{O}}/16.00}$$

```
_Mgn_O=_Mg/24.31m_O/16.00
```

Divide each mole amount by the smaller one, then round to whole numbers

Procedure

1. Heat an empty crucible and lid to constant mass; cool and weigh.
2. Add a coiled strip of clean magnesium ribbon and weigh again.
3. Heat with the lid slightly open so air gets in but smoke does not escape, until the metal no longer glows.
4. Cool, add a few drops of water, and heat again gently (this turns any magnesium nitride into the oxide), then heat strongly.
5. Cool and weigh. Repeat heating until two weighings agree within 0.002 g.

Data table

Measurement	Mass (g)
Crucible + lid	25.120
Crucible + lid + Mg	25.360
Crucible + lid + product	25.520
Mg (by difference)	0.240
Oxygen gained	0.160

Sample data for practice only.

Sample calculation

1. Moles Mg = 0.240 g ÷ 24.31 g/mol = 9.87 × 10⁻³ mol.
2. Moles O = 0.160 g ÷ 16.00 g/mol = 1.00 × 10⁻² mol.
3. Divide by the smaller: Mg 1.00, O 1.01, which rounds to 1 : 1.
4. Empirical formula: MgO.

> **Safety**

- Burning magnesium gives off very bright light, including ultraviolet. Do not stare at it.
- Use tongs; crucibles stay hot long after they stop glowing.
- Never try this at home: magnesium fires are hard to put out, and water makes them worse.

Home or virtual alternative

- Work the same calculation from a percent composition, as in the final practice set (40.0% C, 6.71% H, 53.3% O gives CH_2O).
- OpenStax Chemistry 2e section 3.2 has more empirical-formula problems with answers.

OpenStax 3.2 Determining Empirical and Molecular Formulas (https://openstax.org/books/chemistry-2e/pages/3-2-determining-empirical-and-molecular-formulas)

Planet link: Working out the formula of an unknown compound is a first step in identifying a pollutant, a mineral or a new material.

☐ **I completed this lab (wet or virtual)**

Lab B: Acid–base titration and dilutions

WEEK 6

Goal: Find the concentration of a hydrochloric acid solution by titrating it with sodium hydroxide of known concentration.

$$M_a V_a = M_b V_b \quad (1:1 \text{ acid–base})$$

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M_aV_a=M_bV_b(1 : 1 acid–base)
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For HCl + NaOH → NaCl + H_2O, moles of acid = moles of base at the equivalence point

Procedure

1. Rinse the burette with a little of the 0.1000 M NaOH, then fill it and record the starting reading to 0.01 mL.
2. Pipet 25.00 mL of the HCl into a flask and add 2–3 drops of phenolphthalein indicator.
3. Add NaOH while swirling. Near the end, add it drop by drop.
4. Stop at the first faint pink that lasts 30 seconds. Record the final reading.
5. Repeat until two trials agree within 0.10 mL. With a pH probe, also plot pH against volume to see the steep jump at the equivalence point.

Data table

Trial	Initial reading (mL)	Final reading (mL)	NaOH used (mL)
1 (rough)	0.50	23.10	22.60
2	0.20	22.65	22.45
3	1.05	23.52	22.47

Sample data for practice only. Average the trials that agree (2 and 3).

Sample calculation

1. Average NaOH volume: (22.45 + 22.47) ÷ 2 = 22.46 mL = 0.02246 L.
2. Moles NaOH = 0.1000 mol/L × 0.02246 L = 2.246×10^{-3} mol.
3. The reaction is 1 : 1, so moles HCl = 2.246×10^{-3} mol.
4. Concentration of HCl = 2.246×10^{-3} mol ÷ 0.02500 L = 0.08984 M.

> **Safety**
>
> - Wear goggles: NaOH and HCl can injure eyes and skin even when dilute.
> - If either touches skin, rinse with plenty of water and tell the instructor.
> - Fill the burette below eye level, using a funnel, and remove the funnel before reading.

Home or virtual alternative

- PhET Acid-Base Solutions and pH Scale let you test strong and weak acids and bases with a virtual pH meter.
- Kitchen indicator: red cabbage juice turns red-pink in acids like vinegar and green in bases like baking soda solution. This is qualitative only.

PhET: Acid-Base Solutions (https://phet.colorado.edu/en/simulations/acid-base-solutions) · PhET: pH Scale (https://phet.colorado.edu/en/simulations/ph-scale) · PhET: Molarity (https://phet.colorado.edu/en/simulations/molarity)

Planet link: pH is one of the most important water-quality measurements. The U.S. Geological Survey notes that normal rain has a pH of about 5.6 and that water draining from an abandoned coal mine can reach pH 2, acidic enough to kill fish downstream. USGS Water Science School: pH and Water (https://www.usgs.gov/special-topics/water-science-school/science/ph-and-water)

☐ **I completed this lab (wet or virtual)**

Lab C: Spectrophotometric analysis of a dye solution

WEEK 12

Goal: Make a calibration curve from dye standards and use it to find the concentration of an unknown, using the Beer–Lambert law.

$$A = \varepsilon b c, \qquad A = -\log_{10} T$$

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A= bc, A=-_10T
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A = absorbance, ε = molar absorptivity, b = path length (cm), c = concentration (mol/L), T = fraction of light transmitted

Procedure

1. Warm up the spectrophotometer for the time in its manual.
2. Scan one standard to find the wavelength where the dye absorbs most (λmax).
3. Zero the instrument with a cuvette of the blank (the solvent, here water).
4. Prepare five standards from a stock solution by dilution ($M_1V_1 = M_2V_2$).
5. Measure each standard's absorbance at λmax, rinsing the cuvette with the next solution each time. Handle cuvettes by the frosted sides.
6. Measure the unknown the same way.
7. Plot absorbance (y) against concentration (x), fit a straight line through the origin, and use the slope to find the unknown's concentration.

Data table

Standard	Concentration (μM)	Absorbance
1	2.00	0.130
2	4.00	0.261
3	6.00	0.389
4	8.00	0.522
5	10.00	0.650
Unknown	?	0.455

Sample data for practice only, not a published value for any real dye.

Sample calculation

1. Slope of the line through the origin: about 0.0650 per μM (for example 0.650 ÷ 10.00 μM).
2. With b = 1.00 cm, ε = 0.0650 $\mu M^{-1}\ cm^{-1}$ = 6.50 × $10^4\ M^{-1}\ cm^{-1}$.
3. Unknown: c = A ÷ (εb) = 0.455 ÷ 0.0650 μM^{-1} = 7.00 μM.

4. Fraction of light the unknown transmits: T = 10^(−0.455) = 0.351, or 35.1%.

> **Safety**
>
> - Wear goggles; many dyes stain skin and clothes.
> - Wipe up spills at once and keep liquids away from the instrument.
> - Dispose of solutions as your lab instructor directs.

Home or virtual alternative

- PhET Beer's Law Lab: change concentration, path length and wavelength, and record absorbance to build your own calibration curve.
- Kitchen version: make a series of food-coloring dilutions in identical clear glasses and rank them by eye against a white card. This shows the idea but does not give real absorbance numbers.

PhET: Beer's Law Lab (https://phet.colorado.edu/en/simulations/beers-law-lab) · PhET: Concentration (https://phet.colorado.edu/en/simulations/concentration)

Planet link: The same idea helps satellites judge ocean and lake health. NASA's ocean color missions measure the colors of light leaving the water to estimate chlorophyll, a sign of plankton. NASA Ocean Color (https://oceancolor.gsfc.nasa.gov/)

☐ **I completed this lab (wet or virtual)**

Exam prep practice sets

Time yourself: about 10 minutes per problem. Then compare with the full solutions.

Midterm practice set (Units 1–3)

1. A hydrogen electron drops from n = 4 to n = 2. Find the wavelength of the light.

    1. $|\Delta E| = 2.179 \times 10^{-18}\ \text{J} \times (1/4 - 1/16) = 2.179 \times 10^{-18}\ \text{J} \times 0.1875 = 4.086 \times 10^{-19}\ \text{J}$.
    2. $\lambda = hc/E = (6.626 \times 10^{-34} \times 2.998 \times 10^{8}) \div 4.086 \times 10^{-19} = 4.862 \times 10^{-7}\ \text{m}$.

    Answer: 486 nm (blue-green).

2. Write the ground-state configuration of copper (Z = 29).

    1. The simple filling order predicts [Ar] $4s^2\ 3d^9$, but a full 3d subshell is more stable.
    2. One 4s electron moves to 3d.

    Answer: [Ar] $4s^1\ 3d^{10}$, a known exception.

3. How much CO_2 is produced by burning 1.00 kg of octane (C_8H_{18}, 114.23 g/mol)?

    1. Balanced: 2 C_8H_{18} + 25 O_2 → 16 CO_2 + 18 H_2O, so 8 mol CO_2 per mol octane.
    2. Moles octane = 1000 g ÷ 114.23 g/mol = 8.754 mol.
    3. CO_2 = 8 × 8.754 = 70.04 mol × 44.01 g/mol = 3,082 g.

    Answer: About 3.08 kg of CO_2.

4. 28.0 g of N_2 reacts with 9.00 g of H_2: N_2 + 3 H_2 → 2 NH_3. Find the limiting reactant and the mass of NH_3.

    1. N_2: 28.0 ÷ 28.02 = 0.999 mol. H_2: 9.00 ÷ 2.016 = 4.46 mol.
    2. N_2 needs 3 × 0.999 = 3.00 mol H_2; 4.46 mol is available, so N_2 is limiting.
    3. NH_3 = 2 × 0.999 = 2.00 mol × 17.03 g/mol = 34.0 g.

    Answer: N_2 is limiting; 34.0 g NH_3.

5. Give the shape and polarity of SO_2 and of CCl_4.

    1. SO_2: S has two bonding groups and one lone pair, so it is bent and the dipoles do not cancel.
    2. CCl_4: four identical bonds, tetrahedral, dipoles cancel.

    Answer: SO_2 bent and polar; CCl_4 tetrahedral and nonpolar.

6. Rank by boiling point, lowest first: H_2O, He, CH_4.

    1. He: only weak dispersion, very few electrons.
    2. CH_4: dispersion forces, more electrons than He.
    3. H_2O: hydrogen bonding.

    Answer: He < CH_4 < H_2O.

Final practice set (all units)

1. What volume does 88.0 g of CO_2 occupy at 25.0 °C and 1.00 atm?

    1. n = 88.0 ÷ 44.01 = 2.00 mol.
    2. V = nRT/P = 2.00 × 0.08206 × 298.15 ÷ 1.00.

    Answer: 48.9 L.

2. How many grams of methane must burn (ΔH = −890.5 kJ/mol) to warm 500.0 g of water from 15.0 °C to 85.0 °C, if all the heat goes into the water?

    1. q = 500.0 × 4.184 × 70.0 = 146,440 J = 146.4 kJ.
    2. Moles CH_4 = 146.4 ÷ 890.5 = 0.1644 mol.
    3. Mass = 0.1644 × 16.04 = 2.64 g.

    Answer: 2.64 g of methane.

3. 20.00 mL of sulfuric acid is neutralized by 31.60 mL of 0.1000 M NaOH. Find the acid's molarity.

    1. H_2SO_4 + 2 NaOH → Na_2SO_4 + 2 H_2O (two acidic hydrogens).
    2. Moles NaOH = 0.1000 × 0.03160 = 3.160 × 10^{-3} mol, so moles H_2SO_4 = 1.580 × 10^{-3} mol.
    3. M = 1.580 × 10^{-3} ÷ 0.02000 L.

    Answer: 0.07900 M. A common mistake is to forget the 1 : 2 ratio and get 0.1580 M.

4. A dye has ε = 1.20 × 10^4 M^{-1} cm^{-1}. A sample in a 1.00 cm cuvette reads A = 0.300. Find its concentration.

    1. c = A ÷ (εb) = 0.300 ÷ (1.20 × 10^4 × 1.00).

    Answer: 2.50 × 10^{-5} M.

5. A compound is 40.0% C, 6.71% H and 53.3% O by mass, with a molar mass of 180.16 g/mol. Find its empirical and molecular formulas.

    1. In 100 g: C 40.0 ÷ 12.011 = 3.33 mol; H 6.71 ÷ 1.008 = 6.66 mol; O 53.3 ÷ 16.00 = 3.33 mol.
    2. Divide by 3.33: 1 : 2 : 1, so CH_2O (30.03 g/mol).
    3. 180.16 ÷ 30.03 = 6.

    Answer: Empirical CH_2O; molecular $C_6H_{12}O_6$ (glucose).

6. Use Hess's law to find ΔH for C(s) + ½ O_2 → CO, given C + O_2 → CO_2 (−393.5 kJ) and CO + ½ O_2 → CO_2 (−283.0 kJ).

    1. Keep the first reaction; reverse the second (+283.0 kJ).
    2. Add them; CO_2 cancels.

    Answer: −110.5 kJ.

7. For a process with ΔH = −10.0 kJ and ΔS = −50.0 J/K, find the temperature below which it is spontaneous.

    1. ΔG < 0 when ΔH − TΔS < 0. Both are negative, so it works at low T.
    2. T = ΔH/ΔS = −10.0 kJ ÷ −0.0500 kJ/K = 200 K.

    Answer: Spontaneous below 200 K.

Formula sheet

$$c = \lambda\nu \qquad E = h\nu = hc/\lambda$$

$$E_n = -2.179 \times 10^{-18}\ \mathrm{J}/n^2 \qquad n = m/M$$

$$M = n/V$$

$$M_1V_1 = M_2V_2$$

$$PV = nRT$$

$$P_1V_1/T_1 = P_2V_2/T_2$$

$$P_A = x_A P_{\text{tot}}$$

$$C_g = k_H P_g$$

$$q = mc\Delta T$$

$$\Delta U = q + w$$

$$\Delta H^\circ = \sum n\Delta H_f^\circ(\text{prod}) - \sum n\Delta H_f^\circ(\text{reac})$$

$$\Delta G = \Delta H - T\Delta S$$

$$A = \varepsilon bc$$

$$\text{FC} = V - N - B/2$$

Constants

Constant	Value
Planck constant, h	6.62607015 × 10⁻³⁴ J·s (exact)
Speed of light, c	2.99792458 × 10⁸ m/s (exact)
Avogadro constant, N_A	6.02214076 × 10²³ mol⁻¹ (exact)
Gas constant, R	8.314 J/(mol·K) = 0.08206 L·atm/(mol·K)
Hydrogen energy constant, k	2.179 × 10⁻¹⁸ J
Specific heat of water	4.184 J/(g·°C)
Standard pressure	1 atm = 101.325 kPa = 760 mmHg
Absolute zero	0 K = −273.15 °C

Source: NIST: Fundamental physical constants (https://physics.nist.gov/cuu/Constants/)

Periodic table

Use one of these free interactive tables alongside the course.

- PubChem Periodic Table (U.S. National Library of Medicine) (https://pubchem.ncbi.nlm.nih.gov/periodic-table/): Interactive, with data for every element.
- Ptable (https://ptable.com/): Fast interactive table with orbitals, isotopes and trends.
- NIST Periodic Table of the Elements (https://www.nist.gov/pml/periodic-table-elements): Printable reference table from the U.S. standards agency.

Glossary

Absorbance
How much light a sample absorbs; $A = -\log_{10} T$.

Avogadro's number
$6.02214076 \times 10^{23}$, the number of particles in one mole.

Calorimetry
Measuring heat flow from temperature changes.

Electron configuration
The arrangement of an atom's electrons in orbitals.

Empirical formula
The simplest whole-number ratio of atoms in a compound.

Endpoint
The point in a titration where the indicator changes color.

Enthalpy (H)
Heat content of a system at constant pressure.

Entropy (S)
A measure of how spread out energy and matter are.

Equivalence point
The point where exactly enough titrant has been added to react with the sample.

Gibbs free energy (G)
$G = H - TS$; its change predicts spontaneity.

Hydrogen bond
Strong attraction between H on N, O or F and a lone pair on another N, O or F.

Ideal gas
A model gas whose particles have no volume and no attractions.

Isotope
Atoms of one element with different neutron numbers.

Limiting reactant
The reactant that runs out first.

Molarity
Moles of solute per liter of solution.

Mole
The SI unit for amount of substance.

Orbital
A region where an electron is likely to be found.

Partial pressure
The pressure one gas in a mixture would exert alone.

Percent yield
Actual yield ÷ theoretical yield × 100%.

Photon
A particle of light with energy $E = h\nu$.

Quantum numbers
n, l, m_l, m_s: the four values that describe an electron.

Specific heat
Heat needed to raise 1 g of a substance by 1 °C.

Stoichiometry
The quantitative relationships between amounts in reactions.

Titration
Adding a solution of known concentration to find the concentration of another.

VSEPR
Valence-shell electron-pair repulsion: electron groups spread out to set molecular shape.

Self-assessment rubric

No grade or credit is given. Use this to judge honestly where you are in each unit.

Level	What it looks like
4 · Mastery	Solves new, unseen problems correctly, explains why each step works, and catches their own errors.
3 · Proficient	Solves the practice problems with correct units and significant figures; occasional slips.
2 · Developing	Follows worked examples but gets stuck on new problems or mixes up units.
1 · Beginning	Recognizes the terms but cannot yet set up the calculations.

Suggested weighting if you want a score

Part	Weight
Practice problems	30%
Lab write-ups (or virtual labs)	20%
Midterm practice set	20%
Final practice set	30%

Aim for level 3 or better in every unit before moving to the next course.

Related courses

- Physics I + Lab PHYS 110 + PHYS 110L
- Calculus I MATH 110
- Calculus II MATH 111
- All courses