Teaching Innovation Grant

Final Report Cover Sheet

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Final Report

Proposal Title: Cultivating Resilience in General Chemistry I: Supplemental Instruction to Promot

Final reports should include the following items:

a) Progress made on the project

b) Future work planned for the project

c) Summary of assessment data to date (including student feedback)

d) How funds have been expended

e) How project results have been shared (conference, blog, report, workshop, etc.)

f) Future plan to disseminate results on this campus and externally

g) **Optional**: An appendix may be included with class assignments, form Bs, or other teacher products generated. If you are interested in sharing these items with the SHSU teaching community, please consider adding them to the Engaging Classrooms Library as well. Check out the web page for more details.

Please submit final reports, including this cover sheet, as a PDF to TLC@shsu.edu

Cultivating Resilience in General Chemistry I: Supplemental Instruction to Promote Growth Mindset and Self-Efficacy Final Report of Activities

A. Progress made on the project

The project was completed in full. Supplemental Instruction (SI) sessions for General Chemistry I (CHEM 1411) were designed and implemented during the Fall 2025 and Spring 2026 semesters. Sessions were offered during six weeks of each semester, with up to 12 SI sessions scheduled across the week to increase student access and accommodate varied student schedules. Sessions were held on multiple days and at multiple times, allowing students to attend sessions that best fit their availability.

Student participation was tracked in both semesters. Sessions were not mandatory, and students chose whether to attend the free SI sessions. Participation was encouraged through emails sent by the PI team and the General Chemistry I lecture instructors. Most faculty also offered extra credit to students who attended, and careful attendance records were maintained throughout the project. In Fall 2025, there were 354 total SI attendance instances, with 85 students participating out of 156 enrolled students, for a 54% participation rate. In Spring 2026, there were 169 total SI attendance instances, with 87 students participating out of 363 enrolled students, for a 24% participation rate.

A team of undergraduate SI leaders supported the implementation of the sessions by facilitating problem-solving activities, guiding students through the newly developed materials, and helping create a supportive peer-learning environment. The SI leaders were advanced undergraduate students with experience in chemistry coursework and could serve as near-peer mentors for CHEM 1411 students. Across both semesters, SI leaders completed 396 total hours of project work, including 120 hours in Fall 2025 and 276 hours in Spring 2026. These hours included scheduled SI sessions, preparation for student support, and coordination with the PI team to ensure that the sessions aligned with course content and project goals.

The project also produced substantial new instructional materials that can be reused in future semesters. Problem sets were created for the 10 chapters in CHEM 1411, totaling 94 pages of new student-facing materials. These problem sets were designed to align with the General Chemistry I lecture sequence and to provide students with additional structured practice on core course concepts. SI leader answer keys were also created to support consistent facilitation across sessions and to help undergraduate leaders guide students through problem-solving strategies rather than simply providing answers.

In addition to the written instructional materials, the project team conducted video interviews with two senior faculty members and four undergraduate students at SHSU. Each interview lasted approximately 30 minutes and was edited into 10–12 minutes of useful material, yielding approximately 60 minutes of edited interview content. The interviews focused on students' and faculty members' experiences with learning chemistry, including challenges they encountered, strategies they used to improve, and the importance of persistence, problem-solving, and a growth mindset. These videos will be made available in future course iterations via Blackboard Ultra and the Department of Chemistry website, ensuring the materials continue to support student motivation and self-efficacy beyond the grant period.

B. Future work planned for the project

The project team plans to continue using the learning materials and video interviews developed through this TIG in upcoming semesters. The problem sets, SI leader answer keys, and video interview materials will be shared with General Chemistry I laboratory teaching assistants so they can incorporate them into their scheduled CHEM 1411 office hours each semester.

This approach will allow the project materials to remain in regular use beyond the grant period while continuing to provide students with structured academic support. The materials will help teaching assistants guide students through chemistry-specific problem-solving strategies, effective study practices, and growth mindset concepts. The video interviews will also be used to reinforce messages about persistence, resilience, and self-efficacy in chemistry learning.

By integrating these resources into existing office-hour structures, the department can sustain the project's benefits without requiring the same level of additional grant-funded staffing. The team also plans to revise the materials as needed based on student and teaching assistant feedback so they remain aligned with course content and student needs.

C. Summary of assessment data to date (including student feedback)

C.1. Overall mindset beliefs, and chemistry self-efficacy in control and study groups

Assessment data were collected during the TIG-funded SI implementation period in Fall 2025 and Spring 2026. Surveys were administered at the end of three of the six SI sessions each semester. Two instruments were used: a modified Dweck Mindset Instrument (DMI) and a Chemistry Self-Efficacy instrument. These instruments were used to calculate three measures: the percentage of fixed mindset beliefs (%Fixed), the percentage of growth mindset beliefs (%Growth), and a chemistry self-efficacy percentage score (%CSE).

For comparison, the same instruments had been used during the prior academic year, Fall 2024 and Spring 2025, with CHEM 1411 students who did not participate in the TIG-funded SI sessions (Table 1). This prior-year student population was used as the control group. In that case, the instruments were administered near the end of the semester during laboratory sessions. The study group consists of students who completed surveys after participating in SI sessions during Fall 2025 and Spring 2026.

Table 1: Comparison of mindset beliefs and chemistry self-efficacy between control and study groups.

Group	Count	%Fixed	%Growth	%CSE
Control	485	23%	77%	64.9
Study	125	23%	77%	56.0
Total	611			

The results showed no apparent difference in overall mindset beliefs between the control and study groups. Both groups showed 23% fixed-mindset beliefs and 77% growth-mindset beliefs. However, the study group reported a lower Chemistry Self-Efficacy score than the control group, with an average %CSE of 56.0 compared with 64.9 for the control group.

One possible interpretation is that students who completed the SI problem sets may have developed a more realistic understanding of the level of problem-solving skill required in General Chemistry I. After working through structured problem sets during the SI sessions, these students may have been better able to evaluate their own confidence and abilities in chemistry, leading to

more calibrated self-efficacy responses. In contrast, the control group completed the survey near the end of the semester and may have reported greater confidence after completing the course, potentially reflecting an overestimation of their chemistry problem-solving abilities.

A positive outcome from these assessment data is the high percentage of growth mindset beliefs observed in both groups. Across the control and study groups, 77% of responses reflected growth mindset views, suggesting that most CHEM 1411 students believe their ability to learn chemistry can improve through effort, practice, feedback, and effective learning strategies. This is encouraging because students who view scientific ability as something that can be developed may be more likely to persist through challenging course material, seek help when needed, and respond productively to mistakes.

The PI also considers this high growth-mindset percentage to be connected, at least in part, to broader student-success initiatives at SHSU, including UNIV 1101 and similar courses or programs where growth-mindset concepts have been introduced and discussed with students over the past several years. Although the SI sessions did not appear to shift the overall mindset percentage compared with the control group, the consistently high growth mindset scores indicate that many students already hold learning beliefs supportive of long-term success in chemistry and other STEM fields.

A deeper analysis of the study group showed important differences based on the number of SI sessions students attended. When the study group was disaggregated by the number of SI sessions attended, a positive trend became apparent. Students who attended more SI sessions reported stronger growth-mindset beliefs and higher chemistry self-efficacy scores (Table 2).

Table 2: Mindset beliefs and chemistry self-efficacy by SI session attendance.

Group	Count	Sessions Attended	%Fixed	%Growth	%CSE
Control	485	0	23%	77%	64.9
Study	53	0–2	26%	74%	54.0
Study	33	3–4	23%	77%	57.1
Study	39	5–6	17%	83%	63.8

These results suggest that greater participation in the SI sessions was associated with more positive student beliefs about their ability to grow and improve in chemistry. The trend is especially encouraging because students who attended 5–6 sessions not only reported the highest growth mindset percentage, but also reached a chemistry self-efficacy score similar to that of the control group. This indicates that regular engagement with the SI sessions may have helped students develop more confidence in their chemistry abilities while also strengthening their belief that chemistry learning can improve through effort, practice, and effective strategies. The decrease in fixed mindset beliefs among students who attended 5–6 sessions is also notable. This group reported only 17% fixed mindset beliefs, compared with 23% in the control group and 26% among students who attended 0–2 sessions.

C.2. Relationship between SI participation, student beliefs, and course performance

The project team examined the relationship between SI participation, mindset, chemistry self-efficacy, and course performance (Table 3). These results show a strong positive relationship between the number of SI sessions attended and student course performance. Students who attended 0–2 SI sessions had ABC and DFW rates of 18% and 82%, respectively. In comparison,

students who attended 3–4 sessions had an ABC rate of 56%, while students who attended 5–6 sessions had the strongest outcome, with an ABC rate of 73% and a DFW rate of only 27%. This represents a substantial improvement in course success for students who participated more consistently in the SI sessions.

Table 3: Relationship between SI session attendance, student beliefs, and course performance.

SI Sessions Attended	Count	%Fixed	%Growth	%CSE	Exam 1	Exam 2	Exam 3	Final Exam	Final Grade	%ABC	%DFW
0–2	92	20%	80%	53.0	47.1	33.2	28.9	33.7	46.6	18%	82%
3–4	16	27%	73%	58.6	56.5	37.0	38.8	51.3	60.4	56%	44%
5–6	48	20%	80%	62.1	56.3	51.9	50.9	67.6	79.4	73%	27%
Total	156	22%	78%	59.8	50.9	39.4	36.7	46.0	58.1	39%	61%

Note. Some students who attended SI sessions did not fully complete the assessment surveys. As a result, their responses were not included in the calculation of %Fixed, %Growth, and %CSE for the attendance subgroups. Course performance data were calculated using the available grade records for each subgroup.

Course grades followed the same pattern. Students who attended 0–2 sessions earned an average final course grade of 46.6, while students who attended 3–4 sessions earned an average final grade of 60.4. Students who attended 5–6 sessions earned an average final grade of 79.4, which is more than 30 points higher than the lowest-attendance group. Similar trends were observed across exams, especially later in the semester. For example, the average final exam score increased from 33.7 for students attending 0–2 sessions to 67.6 for students attending 5–6 sessions.

The chemistry self-efficacy results also increased with attendance. Students who attended 0–2 sessions reported a %CSE score of 53.0, compared with 58.6 for students attending 3–4 sessions and 62.1 for students attending 5–6 sessions. This suggests that students who participated more regularly in SI sessions developed greater confidence in their ability to understand and complete chemistry-related tasks.

The mindset results remained generally positive across all groups, with growth mindset percentages ranging from 73% to 80%. Although mindset did not show the same clear stepwise increase as course performance and chemistry self-efficacy in this data set, students across all attendance groups continued to report high levels of growth mindset beliefs. This supports the interpretation that many students already viewed chemistry learning as something that can improve with effort and practice. In contrast, the SI sessions appeared to have a stronger, direct relationship with chemistry confidence and course performance.

Overall, these findings suggest that regular participation in SI sessions was associated with improved chemistry self-efficacy, stronger exam performance, higher final course grades, and lower DFW rates. Because attendance was voluntary, these results should be interpreted as correlational rather than strictly causal. However, the pattern is consistent and encouraging: students who attended more SI sessions were more successful in CHEM 1411. This supports the value of structured, peer-led SI sessions as an effective strategy for improving student success in General Chemistry I.

C.3. Student feedback

Student responses to the prompt "*What's one strategy, concept, or insight from the SI session you're taking away that will help on exams?*" suggest that the SI sessions were most helpful in strengthening practical exam-preparation skills, especially problem setup, dimensional analysis,

unit conversions, and formula selection. Many students reported taking away strategies such as reading questions more carefully, identifying what information is given, determining what the problem is asking, writing out variables, organizing work before solving, and checking units throughout the process.

A major theme was the value of step-by-step problem solving. Students repeatedly mentioned that breaking problems into smaller parts, writing out conversion factors, drawing diagrams when useful, and avoiding shortcuts helped them feel more prepared for exams. Several students also noted that they learned to slow down, avoid second-guessing themselves, and work through problems more deliberately.

Another frequent theme was improved confidence with chemistry-specific skills and concepts. Students identified dimensional analysis, stoichiometry, molarity, dilution, calorimetry, thermochemistry, gas laws, wavelength calculations, Lewis structures, formal charge, VSEPR, and molecular geometry as topics they understood better after the sessions. This suggests that the SI sessions supported both foundational problem-solving habits and content-specific review.

Students also valued the applied and collaborative nature of the sessions. Several responses mentioned that real-world examples, group work, asking questions, and working through practice problems helped them understand not only how to solve problems, but why the procedures mattered. Many students recognized that practice problems were more useful for exam preparation than simply rereading notes or lecture slides.

Overall, the student feedback indicates that the SI sessions helped students develop more organized, reflective, and strategic approaches to chemistry problem solving. The responses also show that students became more aware of the importance of practice, persistence, asking for help, and understanding concepts rather than only memorizing formulas. These outcomes align well with the project goals of improving chemistry self-efficacy, supporting a growth mindset, and helping students prepare more effectively for General Chemistry I exams.

D. How funds have been expended

Grant funds were expended in accordance with the approved project budget. Undergraduate SI leaders worked a total of 396 hours on activities directly related to the project. Their work included preparing for SI sessions, such as reviewing problem sets and SI leader answer keys; facilitating scheduled SI sessions with CHEM 1411 students; supporting students during problem-solving activities; and assisting with the creation and recording of faculty and student interviews. These undergraduate student hours were compensated at $10.00 per hour, totaling $3,960, which matches the budgeted amount for undergraduate student support.

The remaining funds are allocated for faculty compensation. Three faculty members will receive $2,000 stipends each, totaling $6,000, in recognition of their work designing the project, developing instructional materials, coordinating SI leader recruitment and implementation, overseeing assessment and attendance data collection, and preparing the final report. These faculty stipends will be processed once the final report is approved.

In total, the project expenditures account for the full approved grant budget of $9,960.

E. How project results have been shared (conference, blog, report, workshop, etc.)

The project results will be shared through a conference presentation at the upcoming Biennial Conference on Chemical Education 2026, to be held in Madison, Wisconsin, from July 26-30, 2026. The PI of the grant will present the project's results and conclusions, with a focus on the impact of the SI sessions on students' growth mindset, chemistry self-efficacy, and course performance in General Chemistry I.

An abstract for this presentation has been accepted, and the PI has secured travel funds to attend the conference and present the project findings. This presentation will allow the project team to share the SI model's design, implementation, assessment results, and lessons learned with chemistry educators from other institutions. It will also provide an opportunity to receive feedback from colleagues and identify potential improvements for future iterations of the project.

F. Future plan to disseminate results on this campus and externally

The PI plans to further disseminate the results of the TIG-funded project through a communication paper for the Journal *Educación Química*. The paper will describe the design and implementation of the SI sessions, the growth mindset and chemistry self-efficacy components of the project, the instructional materials developed, and the assessment results related to student participation and course performance.

Work on the manuscript will begin during Summer 2026, and the paper will be completed in Fall 2026. The PI will be on sabbatical during Fall 2026 and plans to use part of that time to complete and submit the publication.

On campus, the project materials and results will continue to be shared within the Department of Chemistry. The problem sets, SI leader answer keys, and video interview materials will be made available for future CHEM 1411 support efforts, including use by laboratory teaching assistants during scheduled office hours. The project team also plans to share lessons learned with faculty involved in General Chemistry I instruction so that the materials and strategies developed through the TIG can continue supporting student success beyond the grant period.

G. Appendix information includes the following:

An appendix includes supporting materials generated and used during the project. These materials document the implementation of the TIG-funded work and provide examples of the instructional and assessment resources developed.

The appendix includes the following items:

- Conference abstract accepted for presentation at the Biennial Conference on Chemical Education (BCCE 2026).
- Problem sets and the SI leader's answer key materials created for the 10 chapters covered in CHEM 1411.
- Faculty and student interview protocols used to guide the video interviews focused on persistence, problem-solving, and growth mindset in chemistry.
- An example Supplemental Instruction survey used to collect Dweck Mindset Instrument (DMI) and Chemistry Self-Efficacy data from participating students.

Appendix

Final Report of Activities

Conference: 29th Biennial Conference on Chemical Education (BCCE 2026), July 26 – July 30, 2026. Madison, WI.

Title:

Cultivating Resilience in General Chemistry I: Supplemental Instruction to Promote Growth Mindset and Self-Efficacy

Authors:
Adrian Villalta-Cerdas, Donovan Haines, Dustin Gross

High attrition rates in STEM disciplines continue to challenge efforts to build a resilient scientific workforce. Gateway courses such as General Chemistry I often present academic and psychological barriers that contribute to elevated DFW rates and decreased persistence. This project addresses these challenges by integrating growth mindset principles and chemistry-specific self-efficacy development into a structured Supplemental Instruction (SI) model serving approximately 350 students per semester at Sam Houston State University.

The initiative involves developing and implementing semi-structured SI sessions aligned with the General Chemistry I course content. Sessions combine three core components: (1) growth mindset modules featuring student and faculty narratives that model resilience and emphasize the malleability of intelligence in chemistry; (2) explicit instruction in chemistry-specific learning strategies, including metacognitive approaches and problem-solving frameworks; and (3) guided, curriculum-aligned problem-solving activities that foster collaborative learning. Advanced undergraduate chemistry students are recruited and trained as SI leaders to facilitate sessions, promote a supportive learning environment, and model adaptive academic behaviors. Participation in SI sessions is incentivized through structured course credit to encourage consistent engagement. To evaluate the intervention's impact, the project employs a mixed-methods assessment framework. Changes in students' mindset beliefs are measured using the Dweck Mindset Instrument (DMI), while shifts in chemistry self-efficacy are assessed using the Chemistry Self-Efficacy (CSE) instrument. Academic outcomes are analyzed through comparisons of exam scores, laboratory performance, and overall course grades, alongside participation tracking. Qualitative data from student reflections and feedback surveys provide additional insight into students' experiences and the perceived value of the SI sessions.

Preliminary findings from the first implementation cycle (Fall 2025 and Spring 2026) indicate positive shifts in growth-oriented beliefs and chemistry self-efficacy among participating students, accompanied by improved performance on course assessments relative to non-participants.

Problem 1A - Water Planning for a Desert Hike

During a summer research trip to the Atacama Desert in Chile, scientists must carefully plan their water supplies. Safety guidelines recommend 1.2 liters of water per 3.0 kilometers of hiking under typical daytime conditions.

Main Calculation
If a geologist plans a 22.5 km hike, how many liters of water should they carry?
(Useful conversion: 1 mile = 1.609 km)

Critical Thinking
- What assumptions are built into this estimate?
- Would the actual water need increase or decrease if the terrain were steeper, the humidity higher, or the temperature lower? Explain briefly.

Extension — Connecting to Energy and Real-World Context
- Water isn't just hydration, it's part of your body's cooling system. When you sweat, evaporation removes heat from your body (an endothermic process).
- Estimate roughly how much heat energy could be absorbed if the geologist's volume of water (from above) completely evaporated.
  (Hint: the enthalpy of vaporization for water is about 40.7 kJ/mol; 1 mol = 18.0 g of water)
- What design choices could engineers make in water packs or clothing to improve cooling efficiency while minimizing weight?

Discussion
- In extreme environments—like deserts, high mountains, or space missions—balancing mass, energy, and safety is critical.
- If you were designing supplies for explorers, how would you prioritize between carrying extra water, recycling body moisture, or developing advanced cooling gear?
- What trade-offs (cost, technology, risk) would influence your decision?

Problem 1B - Safe Caffeine Intake

Health guidelines suggest that a moderate caffeine intake should not exceed 3.0 mg of caffeine per kilogram of body mass per day.

Main Calculation
If an adult weighs 165 lb, calculate the maximum recommended caffeine intake (in milligrams) for that person.
(1 lb = 453.6 g)

Critical Thinking
- What assumptions are built into this guideline?
- Would this "safe limit" differ for children, pregnant individuals, or people with heart conditions? Why might mg/kg limits be used instead of a single universal daily limit?

Extension — Connecting to Energy and Real-World Context
An average cup of coffee contains about 95 mg of caffeine, and an energy drink about 160 mg.
- Using your result, estimate how many cups of coffee or cans of energy drink the person could consume before exceeding the guideline.
- Then, discuss what *non-chemical* factors (sleep, hydration, medication, stress) could influence how caffeine affects the body—even at or below this calculated limit.

Discussion
Society often normalizes high caffeine intake in work and study culture.
- If you were designing workplace or campus policies around energy drinks and coffee consumption, how would you balance personal choice, health, and productivity?
- What trade-offs would exist between safety education and regulation?

Problem 1C - Nanodrop Vitamin Supplement

A company claims its new "nanodrop vitamin formula" contains vitamin C (ascorbic acid, $C_6H_8O_6$) at a concentration of 18X, meaning the average mass of vitamin C per drop is 10^{-18} times the total mass of the drop.
Each drop of the supplement has a mass of 0.030 g.

Main Calculation
How many drops would a person have to consume to expect to ingest at least one molecule of vitamin C?
Molar mass of vitamin C = 176.1 g/mol; 1 mol = 6.02 × 10^{23} molecules.

Critical Thinking
- What assumptions are you making in this calculation about uniform mixing and distribution of molecules?
- If you took one drop, would it actually contain *any* molecules at all? Why or why not?

Extension — Connecting to Energy and Real-World Context
Many "nano" or "homeopathic" products claim to deliver benefits despite extreme dilution.
- Estimate the mass of vitamin C in a typical 500 mg tablet and compare it to your calculated total mass for one molecule.
- What does this tell you about scale and scientific communication in consumer health claims?

Discussion Prompt
- If consumers believe such diluted products work, should regulations require quantitative labeling (e.g., molecules per serving) or rely on buyer awareness?
- How can chemistry educators help the public interpret claims that involve scientific units or exponential notation?

Additional Problem 1D – Fuel Efficiency Conversion

A car gets 35 miles per gallon (mpg). Convert this value to liters per 100 km, a common international unit for fuel economy.

Useful conversions:
$1\text{mile} = 1.609\ \text{km}$; $1\text{gallon} = 3.785\ \text{L}$

Answer: ≈ 6.7 L per 100 km

Additional Problem 1E - Counting Atoms in a Microgram

How many silver atoms are in 1.00 μg of pure silver (Ag)?

Molar mass of Ag = 107.9 g/mol
$N_A = 6.02 \times 10^{23}$ atoms/mol

Answer: ≈ 5.6 × 10^{15} atoms

SI Leader Answer Key
Problem 1A - Water Planning for a Desert Hike

Step-by-step solution

Main calculation (liters needed)

1. Convert the recommendation to a per-kilometer rate:

$$1.2\ \text{L} / 3.0\ \text{km} = 0.40\ \text{L/km}$$

2. Multiply by the planned distance (22.5 km):

$$(0.40\ \text{L/km}) \times (22.5\ \text{km}) = 9.0\ \text{L}$$

Note: The mile→km conversion is not needed here—good chance to discuss identifying relevant vs. extra data.

Extension (evaporative cooling energy)

1. Assume density of water 1 kg/L

$$9.0\ \text{L} \rightarrow 9.0\ \text{kg} = 9000\ \text{g}$$

2. Convert mass to moles:

$$9000\ \text{g} \times (1\ \text{mol} / 18\ \text{g}) = 500\ \text{mol}$$

3. Use $\Delta H_{vap} = 40.7$ kJ/mol

$$500\ \text{mol} \times (40.7\ \text{kJ} / 1\ \text{mol}) = 2.04 \times 10^4\ \text{kJ}$$

Critical thinking prompts (for discussion)

- Assumptions: steady pace, level terrain, similar heat/humidity as guideline, no wind-chill or shade effects, no resupply, average physiology.
- Qualitative effects:
  • Steeper/hotter/drier → more water.
  • Higher humidity/lower temperature/slower pace → less water (though humidity can reduce evaporative cooling efficiency).

Key talking points for SI leaders

Likely mistakes / misconceptions

- Ratio handling: Using 1.2 L per km instead of per 3.0 km (missing the division).
- Over-rounding: Rounding intermediate steps too early; maintain guard digits until the end.
- Irrelevant conversions: Trying to convert to miles even though the distance is already in km.
- Evaporation math: Forgetting density (L→kg), skipping g→mol, or using the wrong molar mass for water.

Why each step matters conceptually

- Dimensional analysis: Turning a verbal guideline into a unit ratio (L/km) and chaining units cleanly is Chapter-1 core skill.
- Assumptions awareness: The "right" number depends on environment and physiology; numbers are models, not guarantees.
- Energy linkage: Translating mass of water to latent heat connects measurement units to thermal regulation (bridging to thermochemistry later).

Real-world connection

- Field planning: 9 L ≈ 9 kg—mass matters for safety and pacing.
- Design implications: Cooling efficiency can be improved with reflective clothing, wicking layers, shaded rest, evaporative surfaces, insulated reservoirs, and route/time-of-day planning.
- Risk framing: Underestimating water needs in hot, arid environments is dangerous; teach students to validate assumptions and add safety margins.

SI Leader Answer Key
Problem 1B - Safe Caffeine Intake

Step-by-step solution
Given

- Limit: 3.0 mg caffeine · kg^{-1} · day^{-1}
- Mass: 165 lb
- Conversion: 1 lb = 453.6 g

(a) Convert body mass to kilograms

$$165 \text{ lb} \times \frac{453.6 \text{ g}}{1 \text{ lb}} = 7.4844 \times 10^4 \text{ g}$$

$$7.4844 \times 10^4 \text{ g} \times \frac{1 \text{ kg}}{1000 \text{ g}} = 74.844 \text{ kg}$$

(b) Apply the mg/kg guideline

$$\text{Dose} = (3.0 \text{ mg} \cdot \text{kg}^{-1}) \times (74.844 \text{ kg}) = 224.532 \text{ mg}$$

(c) Extension conversions (cups / energy drinks)

- Coffee: 95 mg per cup

$$\frac{224.532 \text{ mg}}{95 \text{ mg/cup}} = 2.36 \text{ cups} \ (\approx 2 \text{ cups})$$

- Energy drink: 160 mg per can

$$\frac{224.532 \text{ mg}}{160 \text{ mg/can}} = 1.40 \text{ cans} \ (\approx 1 \text{ can})$$

Key talking points for SI leaders

Common mistakes / misconceptions

- Mixing up units: Using mg directly on pounds without converting lb → kg.
- Misreading the guideline: Treating "3.0 mg/kg" as a flat 3.0 mg total.
- Sig figs drift: Over-reporting precision (e.g., giving 224.532 mg as final).
- Serving estimates: Forgetting to round *down* when mapping to real beverages.

Why each step matters conceptually

- Dimensional analysis: This is a classic chain—lb → g → kg → mg via a per-kg ratio; it cements unit fluency early in the course.
- Proportional reasoning: Dose scales with body mass; mg/kg highlights how a single rule adapts to individuals.
- Significant figures: Communicates uncertainty and prevents false precision in health contexts.

Real-world connection / discussion cues

- Assumptions: Healthy adult metabolism, average sensitivity, no interacting meds; real limits vary (children, pregnancy, cardiac issues).
- Behavioral context: Even "within limit," sleep debt, dehydration, or anxiety can magnify caffeine's effects.
- Risk communication: When converting to cups/cans, emphasize variability in actual caffeine content and the prudence of rounding down.

SI Leader Answer Key
Problem 1C - Nanodrop Vitamin Supplement

Step-by-step solution
Given

- Labeling: 18X ⇒ mass of vitamin C per drop = $10^{-18}\times$ (mass of drop)
- Mass of 1 drop = 0.030 g
- Molar mass (vitamin C) $M = 176.1$ g\cdotpmol^{-1}
- Avogadro's number $N_A = 6.02 \times 10^{23}$ molecules\cdotpmol^{-1}

(a) Mass of vitamin C in one drop

$$m_{\text{vitC, drop}} = (0.030\ \text{g}) \times 10^{-18} = 3.0 \times 10^{-20}\ \text{g}$$

(b) Moles of vitamin C in one drop

$$n_{\text{drop}} = \frac{3.0 \times 10^{-20}\ \text{g}}{176.1\ \text{g/mol}} = 1.70 \times 10^{-22}\ \text{mol}$$

(c) Molecules of vitamin C in one drop

$$N_{\text{drop}} = n_{\text{drop}}\, N_A = (1.70 \times 10^{-22}\ mol)(6.02 \times 10^{23}\ molecules\ /\ 1\ mol)$$
$$\approx 1.03 \times 10^{2}\ \text{molecules}$$

(d) Drops needed for at least one molecule (expected value)

$$\text{Drops} = \frac{1\ \text{molecule}}{1.03 \times 10^{2}\ \text{molecules/drop}} \approx 9.7 \times 10^{-3}\ \text{drops}$$

Physically, you can't take a fraction of a drop, so one drop already exceeds one molecule on average.
Answer (main calculation): 1 drop (since ~1.0×10^{2} molecules per drop).

Key talking points for SI leaders

Likely mistakes / misconceptions

- Misreading "18X": Forgetting it means a mass fraction of $10^{-18}$of the *drop mass*, not 18-fold dilution.
- Skipping unit bridges: Not converting mass → moles → molecules using $M$and N_A.
- Exponent slips: Losing powers of ten when multiplying by 10^{-18}or by N_A.
- "One molecule" logic: Thinking you must have >1 drop because the concentration is "tiny". The math shows ~100 molecules per drop despite the extreme label.

Why each step matters conceptually

- Dimensional analysis: Enforces chaining *mass fraction → mass → moles → molecules*.
- Orders of magnitude: Students practice combining very small/very large numbers (exponents) without intuition-only reasoning.

- Statistical thinking: The "at least one molecule" idea is fundamentally probabilistic; the expected count per serving frames that.

Connection to real-world context

- Label interpretation: "X" scales can hide how minuscule (or not) the actual content is. Here, even $10^{-18}$of a macroscopic mass can still contain ~10^{2} molecules because $N_A$is huge.
- Comparison to a tablet (extension prompt):
  - 500 mg tablet = 0.500 gvitamin C.
  - Mass per molecule: $m_{\text{one mol.}} = \frac{176.1\text{ g}}{6.02\times10^{23}} \approx 2.9 \times 10^{-22}$ g/molecule.
  - Molecules in tablet: $\frac{0.500}{176.1}$ mol $\times\, 6.02 \times 10^{23} \approx 1.7 \times 10^{21}$molecules.
  - This stark gap helps students reason about scale and the importance of quantitative labeling in consumer health claims.

Discussion

- Ask students how quantitative labels (e.g., molecules per serving) might change consumer understanding.
- Explore how educators can demystify scientific notation and Avogadro-scale quantities for public decision-making.

SI Leader Answer Key
Additional Problem 1D – Fuel Efficiency Conversion

Given:

35 miles per gallon
1 mile = 1.609 km
1 gallon = 3.785 L

Step-by-step solution
(a) Express mpg as a ratio:

35 miles / 1 gallon

Convert to kilometers per liter:

$$35\,\frac{\text{miles}}{\text{gallon}} \times \frac{1.609\text{ km}}{1\text{ mile}} \times \frac{1\text{ gallon}}{3.785\text{ L}} = 14.96\ \frac{\text{km}}{\text{L}}$$

(b) Convert to liters per 100 km:
Take the reciprocal (to get L/km)

1 L / 14.96 km = 0.0668 L / 1 km

(0.0668 L / 1 km) x 100 km = 6.68 L are necessary for 100 km

Key talking points for SI leaders

- **Common mistakes:**
  - Forgetting to invert the ratio (L/100 km vs. km/L).
  - Dropping conversion factors (especially gallon → liter).
  - Rounding too early or inconsistently.

- **Conceptual focus:**
  - Emphasize that fuel efficiency can be expressed *directly* (mpg) or *inversely* (L/100 km), but both describe the same physical relationship.
  - This is a compound-unit conversion. A good example of dimensional analysis with inverses.

- **Real-world tie-in:**
  - This conversion highlights why different regions (U.S. vs. Europe) use different fuel metrics and how each emphasizes different perspectives on consumption.

SI Leader Answer Key
Additional Problem 1E - Counting Atoms in a Microgram

Given:

$$m = 1.00\ \mu\text{g} = 1.00 \times 10^{-6}\ \text{g}$$
$$M_{\text{Ag}} = 107.9\ \text{g/mol}$$
$$N_A = 6.02 \times 10^{23}\ \text{atoms/mol}$$

Step-by-step solution

(a) Convert mass to moles:

$$n = \frac{m}{M} = \frac{1.00 \times 10^{-6}\ \text{g}}{107.9\ \text{g/mol}} = 9.27 \times 10^{-9}\ \text{mol}$$

(b) Convert moles to atoms:

$$N = n\,N_A = (9.27 \times 10^{-9}\ mol)(6.02 \times 10^{23}\ atom/mol) = 5.58 \times 10^{15}\ \text{atoms}$$

Key talking points for SI leaders

- **Common mistakes:**
  - Confusing micrograms with milligrams (10^{-6} vs. 10^{-3}).
  - Forgetting to convert g → mol → atoms (skipping a step).
  - Misplacing powers of ten in exponent math.

- **Conceptual focus:**
  - Reinforces the idea that even micro-scale masses contain enormous numbers of atoms—a tangible feel for Avogadro's number.
  - Bridges macroscopic mass and microscopic scale, showing how chemists connect lab measurements to particle counts.

- **Real-world tie-in:**
  - Used to explain why trace amounts of metals (like silver or mercury) can still have measurable effects or be detected in environmental samples.
  - Builds intuition for later topics like stoichiometry and molar relationships.

Problem 2A: Nitrogen Oxides and Air Quality

Industrial emissions can form two different nitrogen-oxygen compounds:

- Nitric oxide (NO)
- Nitrogen dioxide (NO_2)

Suppose a fixed mass of nitrogen reacts with 80.0 g of oxygen to form nitric oxide (NO).

Main Calculation
How many grams of oxygen would react with the same mass of nitrogen to form nitrogen dioxide (NO_2), according to the Law of Multiple Proportions?

Critical Thinking

- What assumption do we make when applying the law of multiple proportions to real atmospheric chemistry?
- How might impurities, catalysts, or reaction conditions affect the actual observed mass ratios?

Extension
Nitrogen oxides (collectively called NO_x) are major pollutants formed in car engines, power plants, and industrial processes.

- NO tends to form at lower oxygen availability
- NO_2 tends to form at higher oxygen availability and contributes to smog and acid rain

Explain how controlling the oxygen-to-fuel ratio in engines affects NO vs. NO_2 formation.
Why is understanding this ratio important for energy efficiency and environmental impact?

Discussion
Cities worldwide face trade-offs between:

- reducing vehicle emissions,
- maintaining energy supply,
- and minimizing economic costs.

If you were designing environmental policy, how would you prioritize public health, industrial needs, and fuel costs?
What strategies might strike a balance?

Problem 2B: Lunar Oxygen Isotopes & Origin of the Moon

During a recent lunar mission, scientists analyzed oxygen isotopes in moon rocks to compare them with Earth rocks and study the Moon's origin.

The oxygen in one lunar sample showed the following isotopic composition:

Isotope	Mass (u)	% Abundance
^{16}O	15.9949	99.759%
^{17}O	16.9991	0.038%
^{18}O	17.9992	0.203%

Main Calculation
Calculate the average atomic mass of oxygen in this lunar sample.

Critical Thinking
If Earth rocks have very similar oxygen isotope ratios to these lunar rocks, what does that imply about the Moon's origin?
Why would small deviations still matter to scientists?

Extension
Space missions rely on in-situ resource utilization (ISRU) which means using local materials to support human activity. Oxygen on the Moon could be extracted from rocks for breathing and as rocket oxidizer.

- Why would knowing the exact isotopic makeup of lunar oxygen matter for resource planning, fuel chemistry, or life support systems?
- How might isotope data affect energy requirements for oxygen extraction processes?

Discussion
If future lunar colonies depend on mining oxygen from moon rock, who should control those resources? individual nations, private companies, or international agreements.
What ethical and economic considerations should guide space resource policy?

Problem 2C: Medical Imaging Isotopes

Modern medical imaging sometimes uses stable isotopes of elements to trace biological processes. Imagine a researcher studying iodine uptake in thyroid tissue. Natural iodine consists of two stable isotopes.

One isotope, ^{127}I, has an isotopic mass of 126.9045 amu and a natural abundance of 80.1%. The average atomic mass of iodine in this tissue sample is found to be 126.999 amu.

Main Calculation
Calculate the isotopic mass of the second iodine isotope in this tissue sample.
Assume only two isotopes of iodine are present.

Critical Thinking
- What assumptions do we make when using weighted average atomic mass to back-calculate isotope masses?
- Why might biological materials show slightly different isotope ratios than average Earth samples?

Extension — Real-World Connection
Non-radioactive isotopes are often used as biochemical tracers in medicine, agriculture, and environmental science.
- Why is knowing the exact isotopic composition important for designing tracer experiments?
- How could shifts in isotope ratios help detect metabolic disorders, pollution sources, or food authenticity (e.g., honey purity, origin of crops)?

Discuss how small variations in isotope abundance can carry significant scientific information.

Discussion
If you could add stable isotope tracers to the food or water supply to track nutrient uptake and health outcomes, should society do it?
Consider ethics, privacy, environmental impact, and potential health benefits.

Additional Problem 2D — Law of Multiple Proportions: Sulfur Oxides

Sulfur can form two oxides:

- Sulfur monoxide (SO)
- Sulfur dioxide (SO_2)

If 32.0 g of sulfur combines with 16.0 g of oxygen to form SO, how many grams of oxygen would combine with the same mass of sulfur to form SO_2?

Answer: 32.0 g O

Additional Problem 2E — Average Atomic Mass: Neon Isotopes

Neon occurs naturally as three isotopes:

Isotope	Mass (amu)	Abundance (%)
^{20}Ne	19.992	90.48%
^{21}Ne	20.994	0.27%
^{22}Ne	21.991	9.25%

Calculate the average atomic mass of neon to 4 significant figures.

Answer: 20.20 amu

SI Leader Answer Key
Problem 2A: Nitrogen Oxides and Air Quality

Step-by-step solution

Goal: Same mass of N combines with O to make two oxides. Use the Law of Multiple Proportions: for a fixed mass of one element, the masses of the other element combine in small whole-number ratios.

Approach A (by formula stoichiometry):
- In **NO**, the O:N atom ratio is **1:1** → for a fixed amount of N, oxygen mass is proportional to **1 O atom**.
- In **NO_2**, the O:N atom ratio is **2:1** → oxygen mass is proportional to **2 O atoms**.
- Therefore, for the **same N mass**, the oxygen mass in NO_2 is **twice** that in NO.

Given 80.0 g O(for NO), oxygen for NO_2:

$$m_{\mathrm{O,\,NO_2}} = 2 \times 80.0\ \mathrm{g} = 160.0\ \mathrm{g}$$

Optional cross-check using molar masses:
- In NO: 14.0 g N: 16.0 g O
- In NO_2: 14.0 g N: 32.0 g O
- For the same N mass, O scales $32.0/16.0 = 2 \rightarrow 80.0 \times 2 = 160.0$ g.

Key talking points for SI leaders

Likely mistakes / misconceptions
- Not holding nitrogen constant: Students sometimes keep total mass constant instead of the mass of N, which breaks the law's premise.
- Forgetting whole-number ratio: NO → 1 O atom, NO_2 → 2 O atoms; the oxygen masses must be in a 2:1 ratio for a fixed nitrogen mass.
- Overcomplicating with moles of N: You don't need the actual mass of N—only the relative O per fixed N matters.

Why each step matters conceptually
- Atomic ratios → mass ratios: The law translates integer atom counts (subscripts) into simple mass ratios for a fixed mass of one element.
- Proportional reasoning: Doubling the number of O atoms doubles the mass of O at fixed N; this is foundational for empirical formulas and composition reasoning.

Real-world connection (Extension cues)
- Engine oxygen–fuel control:
  - Lean (more O_2): pushes products toward more oxidized forms (e.g., more NO_2 from NO) but can raise combustion temperature (also promoting NO_x formation thermally).
  - Stoichiometric/tuned control: with exhaust treatment (EGR, three-way catalysts) reduces overall NO_x.

- Why it matters: Balancing the O_2–fuel ratio affects energy efficiency, NO/NO_2 proportions, and downstream pollutants (smog, acid rain). Understanding composition/ratios underpins both chemical and policy decisions.

Discussion

- Ask students how to weigh public health (NO_x reduction) vs costs (fuel, retrofits) and industrial needs.
- Strategies might include engine calibration, after-treatment tech, fuel standards, traffic policy, and economic incentives—each with trade-offs.

SI Leader Answer Key
Problem 2B: Lunar Oxygen Isotopes & Origin of the Moon

Step-by-step solution

Given data

Isotope	Mass (u)	% Abundance
^{16}O	15.9949	99.759%
^{17}O	16.9991	0.038%
^{18}O	17.9992	0.203%

(a) Convert % to decimal fractions

$$x_{16} = 0.99759, x_{17} = 0.00038, x_{18} = 0.00203$$

(b) Weighted average atomic mass

$$\bar{m} = \sum(x_i \cdot m_i) = (0.99759)(15.9949) + (0.00038)(16.9991) + (0.00203)(17.9992)$$
$$\approx 15.95635 + 0.0064597 + 0.0365384 = 15.99935 \text{ u}$$

Key talking points for SI leaders

Common mistakes / misconceptions
- Using a simple (unweighted) average. Emphasize it must be weighted by fractional abundance.
- Forgetting to convert % → decimal (e.g., using 99.759 instead of 0.99759).
- Fractions not summing to 1.000 due to rounding/entry errors—always check normalization.
- Over-rounding intermediate steps, leading to a noticeable bias in the final mass.

Why each step matters conceptually
- Isotopes & averages: The periodic table's atomic masses are weighted averages reflecting natural isotope distributions.
- Mass ↔ composition link: Small changes in isotope fractions shift average mass slightly. This underpins isotope geochemistry and tracer studies.
- Precision: Careful handling of decimals/exponents models good quantitative practice early in the course.

Real-world connection (for Critical Thinking & Extension)
- Moon's origin: If lunar and terrestrial oxygen isotope ratios are nearly identical, it supports a common origin/mixing scenario (e.g., giant impact hypothesis). Small deviations are still scientifically rich. They can indicate fractionation processes, source reservoirs, or thermal histories.
- ISRU implications: Knowing precise isotope make-up can matter for:
    - Resource planning: instrument calibration, expected yields when extracting O from oxides/silicates.
    - Fuel chemistry & life support: isotopic composition can subtly affect diagnostics (e.g., mass spec), monitoring, and accounting; it also guides energy budgeting for extraction

processes (e.g., electrolysis, carbothermal reduction), where composition influences reaction pathways and efficiency.

Discussion

Invite students to connect scientific data → policy/ethics: who governs lunar resources, how to share benefits, and how uncertainty in measurements should influence regulation and international agreements.

SI Leader Answer Key
Problem 2C: Medical Imaging Isotopes

Step-by-step solution

Given

- Isotope 1: ^{127}I mass $m_1 = 126.9045$ amu, abundance $80.1\% = 0.801$
- Isotope 2: unknown mass m_2 =?, abundance $1 - 0.801 = 0.199$
- Average atomic mass (sample): $\bar{m} = 126.999$ amu

Weighted-average equation

$$\bar{m} = (x_1 m_1) + (x_2 m_2) \Rightarrow 126.999 = (0.801)(126.9045) + (0.199)\, m_2$$

Solve for m_2

$$(0.801)(126.9045) = 101.6505045$$
$$126.999 - 101.6505045 = 25.3484955 = 0.199\, m_2$$
$$m_2 = \frac{25.3484955}{0.199} \approx 127.379 \text{ amu}$$

Key talking points for SI leaders

Common mistakes / misconceptions

- Forgetting to convert % to fractions (using 80.1 instead of 0.801).
- Wrong complement abundance (not using $1 - 0.801 = 0.199$).
- Using a simple (unweighted) average of masses instead of weighting by abundances.
- Over-rounding early, which can shift the final result by ~0.01 amu.

Why each step matters conceptually

- Weighted averages reflect how natural samples mix isotopes; the periodic table's atomic masses come from such averages.
- Mass–composition link: Small shifts in abundance change $\bar{m}$subtly. This is the basis of isotope geochemistry and tracer analyses.
- Precision discipline: Keeping guard digits through the algebra avoids compounding rounding error.

Real-world connection

- Tracer design & interpretation: Accurate isotope masses/abundances help calibrate mass spectrometers and interpret uptake pathways in medical/biological studies.
- Natural variability: Biological fractionation or source differences can make tissue ratios differ slightly from terrestrial averages. These small differences can be diagnostic signals (e.g., metabolism, diet, origin, pollution sources).

SI Leader Answer Key
Additional Problem 2D — Law of Multiple Proportions: Sulfur Oxides

Problem Recap

If 32.0 g of sulfur reacts with 16.0 g O to form SO, how many grams of oxygen react with the same amount of sulfur to form SO_2?

Step-by-Step Solution

Step 1 — Ratio for SO

$$\text{S:O} = 32.0\ \text{g}: 16.0\ \text{g}$$

Step 2 — SO_2 has twice as much oxygen per sulfur atom

$$\text{O mass needed} = 2 \times 16.0\ \text{g} = 32.0\ \text{g}$$

Final Answer

$$\boxed{32.0\ \text{g O}}$$

Key Talking Points

- Demonstrates Law of Multiple Proportions: same element (S) combines with different masses of O in simple whole-number ratios (1:2 here).
- Common mistake: trying to do complex stoichiometry instead of recognizing proportional mass relationships.
- Real-world tie-in: sulfur oxides and air pollution — SO_2 contributes to acid rain and particulate formation.

SI Leader Answer Key

Additional Problem 2E — Average Atomic Mass: Neon Isotopes

Problem Recap

Isotope	Mass (amu)	Abundance (%)
^{20}Ne	19.992	90.48%
^{21}Ne	20.994	0.27%
^{22}Ne	21.991	9.25%

Compute the weighted average atomic mass.

Step-by-Step Solution

Convert % → decimal:

$$x_{20} = 0.9048, x_{21} = 0.0027, x_{22} = 0.0925$$

Weighted mass:

$$(0.9048)(19.992) = 18.1056$$
$$(0.0027)(20.994) = 0.056684$$
$$(0.0925)(21.991) = 2.03417$$

Add:

$$18.1056 + 0.056684 + 2.03417 = 20.196454$$

Round to 4 sig figs:

$$\boxed{20.20 \text{ amu}}$$

Note: Some rounding conventions may give ≈ 20.18 amu if rounding earlier. Stress significance of carrying digits.

Key Talking Points

- Emphasize weighted average, NOT a simple average.
- Students often forget to convert % → decimals.
- Reinforces why the periodic table atomic mass for Ne (~20.18) is not an integer, its a natural isotopic mixture.
- Real-world use: isotope ratios help with radiometric dating and atmospheric tracing (e.g., volcanic gas studies).

Problem 3A - Urban Heat and Carbon Emissions

In 2020, a city cleared 150,000 m^2 of forested land to expand housing developments. On average, the forest contained 12.0 kg of carbon per m^2, and it is estimated that 65% of the biomass was burned during the clearing process, releasing carbon dioxide.

1. Main calculation:

Calculate the mass of CO_2 released by this land-clearing event (in kilograms).

2. Critical thinking:

What are the key assumptions in this calculation? How might underestimating or overestimating the percentage of biomass burned affect the result?

3. Extension — connecting to energy/real-world:

The EPA estimates that the average U.S. household emits about 7,500 kg of CO_2 per year from energy use. Compare your calculated result to this value. Roughly how many households' worth of annual emissions does the land clearing equal?

4. Discussion:

Urban development often conflicts with environmental sustainability. How might chemistry-based calculations like this one influence policy decisions about land use?

Problem 3B - Smelting and Environmental Impact

An ore sample contains 15.0% copper(I) sulfide (Cu_2S) by mass. When the ore is heated in oxygen, it reacts to form copper(I) oxide (Cu_2O) and sulfur dioxide (SO_2):

$$2\ Cu_2S(s) + 3\ O_2(g) \rightarrow 2\ Cu_2O(s) + 2\ SO_2(g)$$

Suppose a mining company processes 2.00 kg of this ore, and the reaction runs with a 92.0% yield.

Tasks

1. **Main calculation:**
   How many grams of **Cu_2O** are produced? Show all steps of your dimensional analysis.

2. **Extension – environmental context:**
   OSHA (Occupational Safety and Health Administration) sets the permissible exposure limit (PEL) in workplace air as 5 ppm (or 13 mg/m³) of SO_2 as an 8-hour time-weighted average.

   - How many grams of SO_2 are released in this process at the same yield?
   - If the SO_2 produced during the smelting process were released into a closed volume of 1000 m³ of air, what would be the concentration in mg/m³? Compare your result to the OSHA permissible limit of 5 ppm (≈ 13 mg/m³). Would the concentration be considered safe for human exposure?

3. **Critical thinking:**
   Suppose the ore had been 20.0% Cu_2S instead of 15.0%. How would this affect the yield calculation — proportionally, or in some other way?

4. **Discussion:**
   Copper smelting produces valuable metal but also pollutes the air. If you were designing a policy, what balance would you suggest between yield efficiency and emission controls?

Problem 3C - Breathing and Metabolism

On average, an adult exhales about 1.0 kg of CO_2 per day as a result of cellular respiration. The simplified reaction for the breakdown of glucose is:

$$C_6H_{12}O_6(s) + 6\ O_2(g) \rightarrow 6\ CO_2(g) + 6\ H_2O(l)$$

Tasks

1. **Main calculation:**
   How many grams of glucose ($C_6H_{12}O_6$, 180 g/mol) are needed to produce 1.0 kg of CO_2?

2. **Critical thinking:**
   - If a person eats about 250 g of carbohydrates in a day, how does that compare to the glucose needed to produce the daily CO_2?
   - What assumptions are we making by using glucose as the only energy source?

3. **Extension – health and energy:**
   - The combustion of glucose releases about 16 kJ of energy per gram. Based on your calculation, how much energy (in kJ) is released in one day of respiration?
   - How does that compare to the average daily human energy requirement (~8,700 kJ)?

4. **Discussion:**
   - How do these calculations help us connect chemistry to biology and health? Where might we see oversimplification in this model?

Additional Problem 3D – Rocket Fuel Stoichiometry

Hydrazine (N_2H_4) is used as a rocket propellant. When it burns in oxygen, it produces nitrogen gas and water:

$$N_2H_4(l) + O_2(g) \rightarrow N_2(g) + 2\ H_2O(g)$$

Question:
If a rocket burns 250 g of N_2H_4, how many grams of oxygen are required, and how many grams of water are produced?

Answer: 250 g N_2H_4 requires 250 g O_2 and produces 281 g H_2O.

Additional Problem 3E - Limestone and Carbon Dioxide

Limestone ($CaCO_3$) is used in cement production. When heated, it decomposes into calcium oxide (CaO) and carbon dioxide:

$$CaCO_3(s) \rightarrow CaO(s) + CO_2(g)$$

Question:
If 1.50 kg of $CaCO_3$ is decomposed completely, how many grams of CO_2 are released?
If only 85% of the reaction goes to completion, how much CO_2 is actually produced?

Answer: 1.50 kg $CaCO_3$ releases 660 g CO_2 at 100% yield; 561 g CO_2 at 85%.

SI Leader Answer Key
Problem 3A - Urban Heat and Carbon Emissions

Step 1. Mass of carbon burned

- Area cleared = 150,000 m²
- Carbon density = 12.0 kg C/m²
  Total carbon present = 150,000 x 12.0 = 1.80×10^6 kg C
- Fraction burned = 65% = 0.65
  Carbon burned=$1.80 \times 10^6 \times 0.65 = 1.17 \times 10^6$ kg C

Step 2. Mass of CO_2 released

Reaction:

$$C + O_2 \rightarrow CO_2$$

- M(C) = 12 g/mol
- M(CO_2) = 44 g/mol
- Mass ratio = 44/12 = 3.67
  CO_2 mass = $1.17 \times 10^6 \times 3.67 = 4.30 \times 10^6$ kg CO_2

Step 3. Comparison to household emissions

- Average household = 7,500 kg CO_2/year
  Household equivalents = 4.30×10^6 / 7,500 = 573 households

Key talking points for SI leaders

- **Main calculation (dimensional analysis):**
  Emphasize the stepwise flow: area → carbon mass → burned fraction → moles of carbon → moles of CO_2 → mass of CO_2. This reinforces the importance of unit tracking and converting through stoichiometric ratios, not skipping directly to the end.

- **Critical thinking (assumptions):**
  Highlight the simplifying assumptions:
  - The carbon density is uniform across the land area.
  - All carbon is converted to CO_2 (ignores incomplete combustion).
  - Burning fraction (65%) is an estimate, not an exact measured value.
    Ask students: how would changing these assumptions (e.g., more moisture in biomass, or partial combustion) affect the outcome?

- **Extension (real-world relevance):**
  Stress the comparison to household CO_2 emissions — this makes an abstract mass of CO_2 concrete. Encourage students to reflect on how chemistry connects to energy use, climate, and policy discussions. Units matter here: kg vs tons, or per-day vs per-year values, drastically affect interpretation.

- **Discussion (big-picture context):**
  Guide students to think beyond the math: how can such calculations inform urban planning, emissions reporting, or environmental regulations? Encourage debate: should cities account for CO_2 emissions from deforestation in their climate goals? This broadens the conversation from numbers to societal impact.

SI Leader Answer Key
Problem 3B - Smelting and Environmental Impact

1) Step-by-step solution

Given / Assumptions

- Ore mass = 2.00 kg = 2000 g
- Cu_2S content = 15.0% by mass
- Percent yield = 92.0%
- Balanced reaction:

$$2\ Cu_2S\ (s) + 3\ O_2(g) \rightarrow 2\ Cu_2O(s) + 2\ SO_2(g)$$

- Molar masses ($g \cdot mol^{-1}$):
  Cu_2S = 159.16
  Cu_2O = 143.09
  SO_2 = 64.065

(1) Grams of Cu_2O produced (main calculation)

a. Mass of Cu_2S in the ore

$$m_{Cu2S} = (0.150)\ (2000\ g) = 300.0\ g$$

b. Moles of Cu_2S

$$n_{Cu2S} = 300.0\ g\ /\ 159.16\ g\ mol^{-1} = 1.885\ mol$$

c. Stoichiometry (1:1:1 for Cu_2S : Cu_2O : SO_2)

$$n_{Cu2S} = n_{Cu2O} = 1.885\ mol$$

d. Apply percent yield (92.0%)

$$n_{Cu2O\ actual} = (0.920)\ (1.885) = 1.734\ mol$$

e. Convert to grams of Cu_2O

$$m_{Cu2O} = (1.734\ mol)\ (143.09\ g\ mol^{-1}) = 2.48 \times 10^2\ g$$

(2) Extension — environmental context

a. Grams of SO_2 released (same yield)

$$n_{SO2actual} = (0.920)\ (1.885) = 1.734\ mol$$
$$m_{SO2} = (1.734\ mol)\ (64.065\ g\ mol^{-1}) = 1.11 \times 10^2\ g$$

b. Concentration in a 1000 m^3 closed volume ($mg \cdot m^{-3}$)

$$m_{SO2} = 111\ g = 111{,}000\ mg$$
$$C = 111{,}000\ mg\ /\ 1000\ m^3 = 111\ mg \cdot m^{-3}$$

Compare to OSHA PEL = 13 $mg \cdot m^{-3}$ (≈ 5 ppm):

$111\ mg \cdot m^{-3} \gg 13\ mg \cdot m{-}3$ ➔ Not safe for 8-h exposure

(3) Critical thinking (20.0% vs 15.0% Cu_2S)

If ore were 20.0% Cu_2S (with the same 92% process yield), the theoretical moles of products increase proportionally with the Cu_2S mass. The percent yield then scales the result by the same factor (0.92). So under these assumptions, actual product masses increase proportionally with Cu_2S percentage.

In real operations, yield might also depend on process conditions, but the calculation model here is proportional.

Key talking points for SI leaders

- **Likely mistakes / misconceptions**
  - **Percent by mass**: forgetting to take 15.0% of the ore mass before doing stoichiometry.
  - **Unit conversions**: not converting kg → g for mass; mixing mL, L, m^3 or g ↔ mg incorrectly.
  - **Stoichiometric ratios**: misreading the 2:3:2 coefficients as affecting the $Cu_2S \rightarrow Cu_2O$ or $Cu_2S \rightarrow SO_2$ molar ratio (it's 1:1:1 for each product per Cu_2S).
  - **Percent yield**: applying 92% to the **r**eactant amount or applying it twice; it should scale the theoretical product once.
  - **PEL comparison**: comparing to "5 ppm" directly in ppm without a conversion; the problem provides the equivalence (5 ppm ≈ 13 $mg \cdot m^{-3}$)—compare in $mg \cdot m^{-3}$.

- **Why each step matters conceptually**
  - **Composition** → Reactant moles: Shows the bridge from a mixture (ore) to pure reactant (Cu_2S).
  - **Balanced equation**: Converts moles of reactant to moles of product via coefficients—this is the core of reaction prediction.
  - **Yield**: Reinforces that real processes are not ideal; we scale theory by efficiency.
  - **Mass–concentration link**: Moving from grams to $mg \cdot m^{-3}$ connects lab stoichiometry to exposure standards.

- **Real-world connection**
  - Environmental compliance: Smelting produces valuable Cu_2O, but SO_2 emissions drive air quality and acid rain concerns; plants use scrubbers and process controls to meet OSHA/EPA limits.
  - Policy & design: The numbers support discussions on emission controls, cost–benefit tradeoffs, and community health impacts.

SI Leader Answer Key
Problem 3C - Breathing and Metabolism

1) Step-by-step solution
Given / data

- Target CO_2: 1.0 kg = 1000 g
- Balanced reaction:

$$C_6H_{12}O_6(s) + 6\ O_2(g) \rightarrow 6\ CO_2(g) + 6\ H_2O$$

- Molar masses (use textbook round values unless your course specifies otherwise):
  $M_{CO2} = 44.0\ g\ mol^{-1}$
  $M_{glucose} = 180\ g\ mol^{-1}$

(1) Grams of glucose needed for 1.0 kg CO_2 — *Main calculation*
a. Moles of CO_2 for 1.0 kg

$$n_{CO2} = 1000\ g / 44.0\ g\ mol^{-1} = 22.727\ mol$$

b. Stoichiometric link (from the equation)

1 mol glucose → 6 mol CO2 ➔ $n_{glucose} = n_{CO2} / 6 = 22.727 / 6 = 3.788$ mol

c. Convert moles of glucose to grams

$$m_{glucose} = (3.788\ mol)\ (180\ g\ mol^{-1}) = 681.8\ g$$

(2) Critical thinking comparisons
a. Compare to 250 g carbohydrates/day

- Glucose needed (from above): ≈ 682 g
- Typical carbs eaten: ≈ 250 g
- Observation: 682 g > 250 g. This mismatch suggests:
  - We assumed all CO_2 arises from glucose; in reality, fats and other substrates also produce CO_2.
  - Not all dietary carbohydrate is oxidized directly as glucose; some is stored/converted; protein can also contribute to CO_2.

b. Assumptions made by using glucose as the only fuel

- Glucose is the sole metabolic fuel (ignores fats/proteins).
- Complete oxidation; steady state over the day.
- Biochemical efficiency, nitrogen balance, and storage interconversions are ignored.

(3) Extension — health & energy
Energy released by oxidizing the calculated glucose mass
Given 16 kJ g^{-1} for glucose:

$$q = (681.8\ g)\ (16\ kJ\ g^{-1}) = 10{,}909\ kJ = 1.09 \times 10^4\ kJ$$

Compare to average daily human requirement (~8,700 kJ)

$$10{,}909 / 8{,}700 = 1.25$$

So the simple glucose-only estimate is ~25% higher than the rough daily energy requirement—again highlighting model simplifications (mixed fuels, variable activity, efficiencies).

Key talking points for SI leaders

Likely mistakes / misconceptions

- **Mass–mole confusion:** Dividing 1000 g CO_2 by 180 g/mol (glucose) instead of 44 g/mol (CO_2) in the first step.
- **Stoichiometric ratio error:** Forgetting the 6:1 ratio (6 mol CO_2 per 1 mol glucose).
- **Sig figs / units:** Not converting kg → g or over/under-reporting significant figures.

Why each step matters conceptually

- **Moles as the bridge:** Chemistry happens in moles; converting mass → moles → mass ensures we use the balanced equation correctly.
- **Equation literacy:** The coefficients encode how matter is conserved and rearranged (6 CO_2 for each glucose).
- **Interpreting numbers:** Converting back to grams and then to energy ties abstract stoichiometry to tangible daily quantities.

Real-world connections / model limits

- **Metabolism is mixed-fuel:** Humans oxidize fats and proteins too; assuming glucose only **overestimates** the needed grams of glucose for a given CO_2 mass.
- **Energy budgets vary:** Activity level, basal metabolic rate, thermic effect of food, and metabolic efficiency all shift daily energy needs.
- **Health/nutrition link:** This exercise helps students see how chemistry underpins physiology (respiration ↔ energy), while appreciating where simple models oversimplify real biology.

These talking points should help you guide students beyond the arithmetic and into meaningful interpretation of the chemistry–biology connection.

SI Leader Answer Key
Additional Problem 3D – Rocket Fuel Stoichiometry

Reaction (balanced):

$$N_2H_4(l) + O_2(g) \rightarrow N_2(g) + 2\ H_2O(g)$$

Molar masses ($g \cdot mol^{-1}$):
$M_{N2H4} = 32.05$
$M_{O2} = 32.00$
$M_{H2O} = 18.02$

1) Step-by-step solution
a. Moles of hydrazine

$$n_{N2H4} = 250\ g\ /\ 32.05\ g\ mol^{-1} = 7.80\ mol$$

b. Stoichiometry
From the equation: 1 N_2H_4 : 1 O_2 : 2 H_2O

$$n_{O2} = 7.80\ mol$$
$$n_{H2O} = 2\ (7.80) = 15.6\ mol$$

c. Convert to grams

$$m_{O2} = (7.80)\ (32.00) = 2.50 \times 10^2\ g$$
$$m_{H2O} = (15.6)\ (18.02) = 2.81 \times 10^2\ g$$

Key talking points for SI leaders

- **Likely mistakes:** Using wrong molar mass; forgetting the 2 in front of H_2O; rounding too early.
- **Concept emphasis:** Balanced coefficients drive mole ratios; convert mass → moles → mass.
- **Real-world tie-in:** Hydrazine is energetic/toxic; combustion products & required oxidizer matter for engine design and safety.

SI Leader Answer Key
Additional Problem 3E - Limestone and Carbon Dioxide

Reaction (balanced):

$$CaCO_3(s) \rightarrow CaO(s) + CO_2(g)$$

Molar masses ($g \cdot mol^{-1}$):
$M_{CaCO3} = 100.09$
$M_{CO2} = 44.01$

1) Step-by-step solution
a. Moles of $CaCO_3$ (100% conversion)

$$n_{CaCO3} = 1.50\ kg / 100.09\ g\ mol^{-1} = 1500\ g / 100.09 = 14.99\ mol$$

b. Stoichiometry to CO_2 (1:1)

$$n_{CO2} = 14.99\ mol$$

c. Mass of CO_2 (100% yield)

$$m_{CO2} = (14.99)\ (44.01) = 6.60 \times 10^2\ g$$

d. Mass at 85% completion

$$m_{CO2} = 0.85 \times 6.60 \times 10^2 = 5.61 \times 10^2\ g$$

Key talking points for SI leaders

- **Likely mistakes:** Forgetting to convert kg → g; mixing up molar masses; applying percent completion to reactant instead of product.
- **Concept emphasis:** Decomposition is 1:1 in moles ($CaCO_3 \rightarrow CO_2$); percent completion/percent yield scales the theoretical product.
- **Real-world tie-in:** Cement production is a major CO_2 source; this links simple stoichiometry to industrial emissions and climate impact.

Problem 4A - Water Treatment and Sodium Levels

A water treatment facility is preparing a solution of sodium sulfide (Na_2S) to be used in a chemical process. They start with 50.0 mL of a 0.874 M Na_2S solution and dilute it with water to a final volume of 250.0 mL.

Tasks

1. **Main calculation:**
   What is the final concentration of sodium ions (Na^+) in the diluted solution?

2. **Critical thinking:**
   - Why is it important to consider that Na_2S dissociates into 2 Na^+ ions per formula unit?
   - What mistake might someone make if they ignored this?

3. **Extension – real-world context:**
   The EPA recommends that drinking water sodium concentration stay below 20 mg/L for individuals on a low-sodium diet.
   - Compare the sodium concentration you calculated (in mg/L) to this guideline.
   - Would the diluted solution be considered safe for drinking? Why or why not?

4. **Discussion:**
   If you were working in an environmental chemistry lab, what kinds of errors (calculation, unit conversion, or chemical assumption) could lead to underestimating sodium in water samples?

Problem 4B - Checking Nitrate Acidity in Runoff

A lab is monitoring nitrate-rich runoff near an agricultural site. A 25.0 mL water sample is known to contain a strong monoprotic acid due to nitric acid (from nitrate formation). The sample is titrated with 0.0200 M NaOH(aq), and the endpoint is reached after 12.50 mL of NaOH is delivered.

$$HNO_3(aq) + NaOH(aq) \rightarrow NaNO_3(aq) + H_2O(l)$$

Tasks

1. **Concentration of the acid:**
   Determine the molarity of the acid in the original 25.0 mL sample.

2. **pH connection:**
   Assuming the acid is HNO_3 and behaves as a strong acid, estimate the initial pH of the sample *before* titration (ignore activity effects). pH = - log($[H^+]$)

3. **What if it's diprotic?**
   Suppose the same titration data came from a sample where the acid was actually sulfuric acid (strong, diprotic). What would be the molarity of H_2SO_4 under the same volumes and NaOH concentration?

4. **Discussion:**
   Given that drinking water standards typically restrict pH to ~6.5–8.5, what risks might arise from the runoff water? If you were an environmental chemist reporting these results, how would you explain to non-scientist stakeholders (farmers, community leaders, regulators) what the measured pH means in terms of water safety?

Problem 4C - Hemoglobin and Oxygen Transport

A typical adult has about 6.0 L of blood, and the hemoglobin content is about 16 g per 100 mL of blood. The molar mass of hemoglobin is approximately 64,000 g/mol.

Tasks

1. **Molar concentration:**
   Calculate the molar concentration of hemoglobin in blood.

2. **Total moles of hemoglobin:**
   How many moles of hemoglobin are present in a typical adult?

3. **Critical thinking:**
   Each hemoglobin molecule can bind up to 4 molecules of O_2. Based on your answer above, estimate how many moles of O_2 could theoretically be carried by the hemoglobin in a human body.

4. **Extension – biological connection:**
   At rest, an adult consumes about 250 mL O_2 per minute. Roughly how long would the total hemoglobin O_2 supply last if no new oxygen entered the body? What assumptions are we making here? Use the conversion factor 1.00 mol gas = 22.4 L gas at STP.

5. **Discussion:**
   Why is it important for chemists and biologists to understand the limits of simplified models when connecting molecular quantities to real physiology?

Additional Problem 4D – Dilution and Ionic Concentration

You have 35.0 mL of 0.250 M $CaCl_2(aq)$. This solution is diluted to a final volume of 150.0 mL.

Question:
What is the concentration of chloride ions (Cl^-) in the diluted solution?

Answer: 0.117 M Cl^-

Additional Problem 4E – Acid-Base Titration

A 20.0 mL sample of $H_2SO_4(aq)$ of unknown concentration is titrated with 0.100 M NaOH(aq). It takes 32.0 mL of NaOH to reach the endpoint.

$$H_2SO_4(aq) + 2\ NaOH(aq) \rightarrow Na_2SO_4(aq) + 2\ H_2O(l)$$

Question:
What is the molarity of the H_2SO_4 solution?

Answer: 0.0800 M H_2SO_4

SI Leader Answer Key
Problem 4A - Water Treatment and Sodium Levels

Step-by-step solution
Given / data

- Stock solution: 0.874 M Na_2S
- Aliquot volume: 50.0 mL = 0.0500 L
- Final (diluted) volume: 250.0 mL = 0.2500 L
- Sodium atomic mass (for mg/L conversion): 22.99 $g \cdot mol^{-1}$
- Note: $Na_2S \rightarrow 2\ Na^+ + S^{2-}$ (complete dissociation, ideal assumption)

(1) Final $[Na^+]$ after dilution
a. Diluted molarity of Na_2S ($M_1V_1 = M_2V_2$)

$$M_{Na_2S,\ final} = (0.874\ M)\ (0.0500\ L) / 0.2500\ L = 0.1748\ M = 0.175\ M$$

b. Convert Na_2S molarity to Na^+ molarity (2 Na^+ per Na_2S)

$$[Na^+] = 2 \times 0.1748 = 0.3496\ M = 0.350\ M$$

(2) Compare to EPA sodium guideline (mg/L)
c. Convert molarity of Na^+ to mg/L

$$mg/L\ Na+ = (0.3496\ mol/L) \times (22.99\ g/mol) \times 1000\ mg / 1\ g$$
$$mg/L\ Na+ = 8.037 \times 10^3\ mg/L = 8.04 \times 10^3\ mg/L$$

EPA note (for low-sodium diets): ≤ 20 $mg \cdot L^{-1}$ recommended.
This solution (~8037 $mg \cdot L^{-1}$) is far above that level → not safe for drinking (and is not intended for consumption; it's a process solution).

Key talking points for SI leaders

Likely mistakes / misconceptions

- **Forgetting dissociation:** Treating Na_2S as giving 1 Na^+ instead of 2 Na^+ (missing the factor of 2).
- **Dilution misstep:** Using $M_1V_1 = M_2V_2$ with mL without consistent units, or plugging 250 instead of 0.250 L.
- **Wrong species for mg/L:** Converting the salt (Na_2S) to mg/L instead of the ion of interest (Na^+), or using the molar mass of Na_2S instead of Na.
- **Rounding too early:** Rounding intermediate steps can shift the final digits; keep guard digits until the end.

Why each step matters conceptually

- **$M_1V_1 = M_2V_2$** captures how moles are conserved during dilution (moles of solute unchanged; concentration changes with volume).
- **Dissociation stoichiometry** links formula units to ionic concentrations. Key for any ionic solution problem (and later, for conductivity, colligative effects, equilibria).
- **Unit translation to mg/L** connects laboratory molarity to public health standards, which are usually mass-per-volume.

Real-world connection

- **Process vs. potable water:** Industrial or treatment-process solutions can be orders of magnitude more concentrated than drinking water standards; understanding this prevents misinterpretation of safety.
- **Analytical vigilance:** Environmental labs often report mg/L; small unit errors can produce huge compliance mistakes. This problem mirrors real chains of custody and reporting requirements.

SI Leader Answer Key
Problem 4B - Checking Nitrate Acidity in Runoff

Step-by-step solution
Given / data

- Sample volume (acid): 25.0 mL = 0.0250 L
- Titrant: 0.0200 M NaOH(aq)
- Volume of NaOH at endpoint: 12.50 mL = 0.01250 L
- Reaction (monoprotic case):

$$HNO_3(aq) + NaOH(aq) \rightarrow NaNO_3(aq) + H_2O(l)$$

- Reaction (diprotic variant):

$$H_2SO_4(aq) + 2\ NaOH(aq) \rightarrow Na_2SO_4(aq) + 2\ H_2O(l)$$

(1) Molarity of the acid (HNO_3, strong monoprotic)
Moles NaOH at endpoint

$$n_{NaOH} = M \times V = (0.0200\ mol/L)\ (0.01250\ L) = 2.50 \times 10^{-4}\ mol$$

At equivalence (1:1), moles acid = moles base

$$n_{HNO3} = 2.50 \times 10^{-4}\ mol$$

Molarity in original 25.0 mL sample

$$M_{HNO3} = n / V = 2.50 \times 10^{-4}\ mol / 0.0250\ L = 1.00 \times 10^{-2}\ M = 0.0100\ M$$

(2) Initial pH (strong acid approximation)
For a strong monoprotic acid: $[H^+] \approx M_{acid} = 0.0100\ M$

$$pH = -\log(0.0100\ M) = 2.00$$

(3) If the acid were strong, diprotic (H_2SO_4)
Stoichiometry at equivalence (2:1 NaOH:H_2SO_4):

$$n_{H2SO4} = n_{NaOH} / 2 = 2.50 \times 10^{-4}\ mol / 2 = 1.25 \times 10^{-4}\ mol$$

Molarity in the same 25.0 mL sample:

$$M_{H2SO4} = 1.25 \times 10^{-4}\ mol / 0.0250\ L = 5.00 \times 10^{-3}\ M = 0.00500\ M$$

Key talking points for SI leaders

Likely mistakes / misconceptions

- **mL vs L:** Forgetting to convert 12.50 mL and 25.0 mL to liters before using MxV or n/V.
- **Stoichiometric ratio:** Treating H_2SO_4 as 1:1 with NaOH (it's 2:1) or forgetting that HNO_3 is 1:1.
- **pH from wrong concentration:** Using the titrant concentration to compute pH, or using post-dilution numbers instead of the initial acid molarity.
- **Sig figs:** Reporting "0.01 M" without trailing zeros; here data support 0.0100 M (3 significant figures).

Why each step matters conceptually

- **Equivalence concept:** At the endpoint, moles acid neutralized = moles base added (adjusted by stoichiometric coefficients). This is the bridge from volume × molarity to analyte concentration.
- **Stoichiometry in titration:** Balanced equations determine mole ratios—crucial when comparing mono- vs diprotic acids.
- **pH interpretation:** For strong acids, $[H^+] \approx M_{acid}$, linking titration results to acidity in a meaningful way.

Real-world connection and discussion

Remind students that drinking water is normally pH 6.5–8.5.
Anything outside this range is unusual and can cause corrosion, affect taste, or harm health.

Interpreting low pH:

- A pH near 2 is very acidic, comparable to lemon juice or stomach acid.
- Such acidity would cause severe irritation or harm if consumed, and can also leach toxic metals (like lead or copper) from pipes.

Risks from runoff water:

- Environmental impact: acidic runoff can kill aquatic life, alter soil chemistry, and mobilize heavy metals.
- Human impact: unsafe for drinking, may require costly treatment.

Communication challenge:

- Scientists think in numbers (pH, molarity); communities think in safety and risk.
- SI leaders should guide students to practice rephrasing results in plain language: e.g., "This water is far more acidic than safe drinking water, and could be harmful if not treated."

Encourage multiple perspectives:

- Students may suggest different communication strategies: using analogies (e.g., "this water is as acidic as soda"), citing official standards, or explaining risks in terms of health outcomes.
- All of these are valid approaches and enrich group discussion.

SI Leader Answer Key
Problem 4C - Hemoglobin and Oxygen Transport

Step-by-step solution
Given / data

- Blood volume (adult): 6.0 L
- Hemoglobin content: 16 g per 100 mL blood
- Molar mass of hemoglobin (Hb): 64,000 g/mol (use 6.4×10^4 g/mol)
- O_2 binding: 4 O_2 / 1 Hb
- STP conversion (provided): 1.00 mol gas = 22.4 L gas
- Resting O_2 consumption rate: 250 mL/min = 0.250 L/min

(1) Molar concentration of hemoglobin in blood
Convert 16 g per 100 mL to $g \cdot L^{-1}$:

$$16 \text{ g} / 100 \text{ mL} = 16 \text{ g} / 0.100 \text{ L} = 160 \text{ g} / \text{L}$$

Convert mass concentration to molarity:

$$[\text{Hb}] = 160 \text{ g/L} / 64{,}000 \text{ g/mol} = 2.5 \times 10^{-3} \text{ mol/L} = 2.5 \times 10^{-3} \text{ M}$$

(2) Total moles of hemoglobin in an adult

$$n_{Hb} = [\text{Hb}] \times V_{blood} = (2.5 \times 10^{-3} \text{ mol/}) (6.0 \text{ L}) = 1.5 \times 10^{-2} \text{ mol}$$

(3) Theoretical O_2-carrying capacity (moles of O_2)

$$n_{O2,\,cap} = [\ 4 \text{ mol } O_2 / \text{ mol Hb }] \times n_{Hb} = 4 \times 1.5 \times 10^{-2} \text{ mol } O_2 = 6.0 \times 10^{-2} \text{ mol } O_2$$

(4) Time the Hb-bound O_2 would last (no new O_2 entering)
Convert moles O_2 to volume at STP:

$$V_{O2} = n_{O2,\,cap} \times 22.4 \text{ L/mol} = (6.0 \times 10^{-2} \text{ mol }) \times 22.4 \text{ L/mol} = 1.34 \text{ L}$$

Time at a use-rate of 0.250 L/min:

$$t = 1.34 \text{ L} / 0.250 \text{ L/min} = 5.36 \text{ min}$$

Key talking points for SI leaders
Likely mistakes / misconceptions

- **Volume-to-mass conversion slip:** Not converting "16 g per 100 mL" to g/L (multiply by 10), or confusing per-100 mL with per-liter directly.
- **Concentration vs amount:** Forgetting to multiply molarity by total blood volume to get moles of Hb.
- **Binding stoichiometry:** Omitting the x 4 O_2 per Hb factor.
- **Gas conversion & units:** Using the wrong molar gas volume (e.g., 24.5 L/mol) instead of the provided 22.4 L/mol at STP; forgetting to convert 250 mL/min to 0.250 L/min.
- **Sig figs:** Over- or under-reporting; keep results consistent with inputs (e.g., 6.0 L blood → 2 s.f.).

Why each step matters conceptually

- **From mass ratio to molarity:** Translating a clinical mass-per-volume figure (g per 100 mL) into moles per liter reinforces how chemists quantify species to use balanced relationships.
- **Scaling concentration to the whole system:** Demonstrates the difference between intensive ([Hb]) and extensive (total moles) quantities.
- **Stoichiometry as a bridge to function:** The 4:1 O_2:Hb ratio connects molecular structure to physiological capacity.
- **Inter-domain translation:** Converting moles ⇄ liters of gas and then to a time scale links chemistry to biological rates.

Real-world connection & model limits

- **Physiology is dynamic:** Our estimate assumes 100% Hb O_2 saturation, that all bound O_2 is available, constant consumption at 0.250 L/min, and STP gas behavior. Real blood O_2 content varies with partial pressure (O_2-Hb dissociation curve), tissue demands, and cardio-respiratory regulation.
- **Other reservoirs:** We ignored dissolved O_2, myoglobin in muscle, and CO_2/HCO_3^- buffering. All relevant in real physiology.
- **Takeaway:** The numbers are a useful first approximation that helps students connect molecular counts to bodily function, while appreciating the importance of assumptions and context when interpreting quantitative results.

SI Leader Answer Key
Additional Problem 4D – Dilution and Ionic Concentration

Given / data

- Initial: $V_1 = 35.0$ mL $= 0.0350$ L
- $M_1(CaCl_2) = 0.250$ M
- Final volume: $V_2 = 150.0$ mL $= 0.1500$ L
- Dissociation: $CaCl_2\ (aq) \rightarrow Ca^{2+}(aq) + 2\ Cl^{-1}(aq)$

Step-by-step solution

a) Dilution of $CaCl_2$ (moles conserved):

$$M_1V_1 = M_2V_2 \Rightarrow M_2(CaCl_2) = (0.250)\ (0.0350) / 0.1500 = 0.05833\ M$$

b) Convert salt molarity to ion molarity:

$$[Cl^{-1}] = 2 \times M_2(CaCl_2) = 2\ (0.05833) = 0.11666\ M$$

Key talking points for SI leaders

Likely mistakes:

- Using $M_1V_1 = M_2V_2$ with mL inconsistently; forgetting to divide by V_2.
- Forgetting the factor of 2 for chloride from $CaCl_2$.

Why steps matter:

- Dilution uses mole conservation (moles of solute stay constant; concentration changes with volume).
- Dissociation stoichiometry connects formula units to ion concentrations, crucial for solution chemistry.

Real-world link:

- Correct ion concentrations underpin water hardness, electrolyte balance, and conductivity assessments.

SI Leader Answer Key
Additional Problem 4E – Acid-Base Titration

Given / data

- Acid volume: V_{acid} = 20.0 mL = 0.0200 L
- Base: M_{NaOH} = 0.100 M
- V_{NaOH} = 32.0 mL = 0.0320 L
- Reaction: $H_2SO_4 + 2\ NaOH \rightarrow Na_2SO_4 + 2\ H_2O$

Step-by-step solution

a) Moles of NaOH delivered at endpoint:

$$n_{NaOH} = M \times V = (0.100)(0.0320) = 3.20 \times 10^{-3} \text{ mol}$$

b) Stoichiometry to find moles of H_2SO_4:

$$n_{H2SO4} = n_{NaOH} / 2 = 3.20 \times 10^{-3} / 2 = 1.60 \times 10^{-3} \text{ mol}$$

c) Molarity of the acid:

$$M_{H2SO4} = n / V = 1.60 \times 10^{-3} / 0.0200 = 0.0800 \text{ M}$$

Key talking points for SI leaders

Likely mistakes:

- Not converting mL → L before using M x V.
- Treating H_2SO_4 as monoprotic (forgetting the 2:1 NaOH : H_2SO_4 ratio).
- Swapping analyte and titrant volumes when computing molarity.

Why steps matter:

- Equivalence point logic: moles acid (× acidity) = moles base; coefficients dictate proton count.
- Reinforces reading balanced equations to determine mole ratios in titrations.

Real-world link:

- Accurate titrations are foundational in environmental monitoring, industrial QC, and clinical labs; small ratio/unit errors lead to large reporting errors.

Problem 5A - Catalytic Converters & Heat from CO Oxidation

In a catalytic converter, toxic carbon monoxide (CO) in exhaust is oxidized to carbon dioxide (CO_2):

$$2\,CO(g) + O_2(g) \rightarrow 2\,CO_2(g) \qquad \Delta H^\circ = -482\ kJ$$

A maintenance test shows an engine's exhaust stream contains a "slug" of 89.5 g of CO that gets fully converted in the catalyst (assume excess O_2).

Tasks

1. **Main calculation — Heat released:**
   How much heat is released by the oxidation of 89.5 g CO?

2. **Oxygen demand (quick stoichiometry):**
   How many grams of O_2 does this conversion consume?

3. **Applied concepts:**
   If an engine produced 12.0 g CO per minute (before the converter), estimate the heat released per minute in the converter when that CO is oxidized.

4. **Discussion (safety/design):**
   Catalytic converter bricks can overheat ('thermal runaway'). Based on your answers, why might airflow, heat shields, and temperature sensors be critical in converter design?

Problem 5B - Inflating an Airbag

Airbags in cars inflate rapidly during a crash, producing nitrogen gas that expands against the atmosphere. Imagine an airbag system where the nitrogen gas expands and does 175 kJ of work on the surroundings at a constant external pressure of 5.55 atm. The airbag initially contains only a negligible gas volume.

Tasks

1. **Main calculation:**
   What is the final volume of nitrogen gas in the airbag? Use the conversion 101.3 J = 1 L atm

2. **Critical thinking:**
   Why must we carefully match the energy released to the volume of the airbag? What would happen if the airbag inflated too little or too much?

3. **Extension — connecting to energy:**
   175 kJ is about the energy stored in a chocolate bar. Compare this to the role of energy in the airbag system. Why is "energy balance" a useful way to think about safety devices?

4. **Discussion:**
   Where else in everyday technology do gases expanding against pressure play a critical role, and what challenges might engineers face in those systems?

Problem 5C - Exothermic Dissolution & Safety

A maintenance crew prepares a heavy-duty cleaning solution by dissolving 7.50 g of KOH(s) (molar mass 56.11 g/mol) into 200.0 g of water at room temperature. After stirring, the solution temperature rises by 9.5 °C. Assume the specific heat capacity of the solution is 4.18 J/g °C and that the total mass of the solution is the sum of solute + solvent.

1. Main calculation:
Calculate the molar enthalpy of solution, ΔH_{soln} (kJ/mol), for KOH under these conditions. Be careful with the sign of ΔH_{soln}.

2. Critical thinking:
List two assumptions you made in the calculation. For each, explain briefly how violating that assumption (e.g., heat loss to the beaker/air, using water's c instead of the true solution c) would change the magnitude and/or sign of your ΔH_{soln} estimate.

3. Extension — connecting to energy:
KOH is used in some drain cleaners. Using your result, estimate the heat released when 50.0 g of KOH dissolves (assume the same ΔH_{soln}). In a small, closed space (like a pipe), why might rapid heat release and steam formation be a safety concern?

4. Discussion:
Endothermic salts (e.g., NH_4NO_3) cool their surroundings, while exothermic salts (e.g., KOH, $CaCl_2$) heat them. How could you design a safe classroom demo that lets students feel the difference—without risking burns or spills?

Additional Problem 5D – Dissolving NaOH

When 25.0 g of NaOH(s) dissolves in water, the temperature of the solution increases by 12.0 °C. If the solution has a mass of 250 g and a specific heat capacity of 4.18 J/g·°C, what is the heat released (q) during this process?

Answer: 12.5 kJ

Additional Problem 5E - Burning Methane

The combustion of methane releases –890 kJ/mol **o**f heat. How much heat is released when 5.00 g of methane (CH_4) is burned? (Molar mass = 16.0 g/mol)

Answer: 278 kJ

SI Leader Answer Key
Problem 5A – Catalytic Converters & Heat from CO Oxidation

Step-by-step solution

Data

- Thermochemical equation (as written):
  $2\ CO(g) + O_2(g) \rightarrow 2\ CO_2(g)$ $\Delta H° = -482$ kJ (per 2 mol CO)
- Therefore per 1 mol CO, $\Delta H° = -482$ kJ/mol
- Molar masses ($g \cdot mol^{-1}$): CO = 28.01 g/mol, O_2 = 32.00 g/mol

(1) Heat released for 89.5 g CO

1. Moles CO:
   $N_{CO} = 89.5\ g / 28.01 g/mol = 3.196\ mol$

2. Heat (scale enthalpy per mol CO):
   $q = (3.196\ mol) \times (241\ kJ/mol) = 7.70 \times 10^2\ kJ$

Sign: released; exothermic.

(2) Oxygen demand (grams O_2)

Stoichiometry 2 CO : 1 O_2:

$$n_{O_2} = \frac{3.196}{2} = 1.598\ mol$$

$$m_{O_2} = (1.598\ mol)\ x\ (32.00\ g/mol) = 51.1\ g$$

(3) Heat released per minute for 12.0 g CO/min

1. Moles CO per minute:
   $n_{CO} = (12.0\ g/min)\ /\ (28.01\ g/mol) = 0.428\ mol/min$

2. Heat per minute:
   $q = (0.428\ mol/min) \times (241\ kJ/mol) = 1.03 \times 10^2\ kJ/min$

Key talking points for SI leaders

Likely mistakes / misconceptions

- **Thermochemical scaling:** Treating $\Delta H° = -482$ kJ as per 1 mol CO instead of per reaction as written (2 mol CO).
- **Sign convention:** Dropping the negative sign (exothermic) or calling the heat "absorbed."
- **Mass ↔ moles:** Using grams directly with $\Delta H°$ instead of converting to moles.
- **Stoichiometry for O_2:** Forgetting the 2:1 ratio (CO:O_2), or using 1:1.

- **Rate vs. amount:** For part (3), some students compute total heat for 12.0 g once, not per minute.

Why each step matters conceptually

- **Thermochemical equations are quantitative:** $\Delta H°$ is tied to the balanced coefficients; scaling to actual moles is essential.
- **Proportional reasoning:** Heat change is proportional to extent of reaction—a core idea that generalizes to any reaction progress.
- **Coupling stoichiometry with energy:** Converting mass → moles → heat shows how composition determines thermal load.
- **From rate to power:** Translating kJ/min to kW helps connect chemistry to engineering design constraints.

Real-world context / safety link

- **Converter overheating:** The exothermic oxidation of CO can generate kilojoules of heat at significant flow—risking thermal runaway, catalyst sintering, and substrate damage if heat isn't managed.
- **Design implications:** Adequate airflow, heat shielding, and temperature monitoring/feedback are critical; chemistry informs these engineering controls.
- **Emissions & compliance:** Effective CO conversion reduces toxic exposure and helps meet regulatory limits—a direct chemistry-to-society impact.

SI Leader Answer Key
Problem 5B - Inflating an Airbag

Step-by-step solution

Given / data

- Work done by the gas on surroundings: $w = 175$ kJ(magnitude)
- Constant external pressure: $P_{ext} = 5.55$ atm
- Unit bridge: $101.3\ \text{J} = 1\ \text{L atm}$
- Initial gas volume: negligible ⇒ $V_f \approx \Delta V$

At constant external pressure, $w_{(\text{by gas})} = P_{ext}\Delta V$ in magnitude.
Chem sign convention: $w_{system} = -P_{ext}\Delta V$; we use the magnitude to get ΔV.

a) Convert work to joules

$$175\ \text{kJ} = 175{,}000\ \text{J}$$

b) Solve for ΔV using the L·atm bridge

$\Delta V = (w / P_{ext})$ x (1 L atm / 101.3 J) = (175,000 J / 5.55 atm) x (1 L atm / 101.3 J) = 3.11×10^2 L

c) Final volume (initial volume negligible)

$$V_f \approx \Delta V \approx 3.11 \times 10^2\ \text{L}$$

Key talking points for SI leaders

Likely mistakes / misconceptions

- **Unit bridge omission:** Forgetting 101.3 J = 1 L atm; mixing atm with J directly.
- **Sign confusion:** Plugging a negative w from chem convention into $w = P\Delta V$ and getting a negative volume change. Clarify we used magnitude to find ΔV.
- **Pressure misuse:** Using gauge pressure or variable P instead of the stated constant external pressure.
- **Initial volume:** Trying to subtract a non-zero V_i; here it's negligible, so $V_f \approx \Delta V$.

Why each step matters conceptually

- **Path function:** Expansion work at constant P_{ext} illustrates w as path-dependent, not just state-dependent.
- **Dimensional analysis:** The L·atm ↔ J bridge is essential for connecting mechanical work and pressure–volume changes.
- **Energy–volume link:** Shows how a specified energy release maps to a physical volume required to perform that work against the surroundings.

Real-world connection

- **Safety envelope:** Too little volume → insufficient cushioning; too much volume/pressure → risk of injury or bag rupture.
- **Engineering controls:** Results motivate airflow design, venting, sensor feedback, and material limits in airbags and other systems where gases expand against pressure (e.g., life vests, aerosol cans, engine exhaust after treatment).

SI Leader Answer Key
Problem 5C - Exothermic Dissolution & Safety

Step-by-step solution

Data

- $m_{KOH} = 7.50$ g, $M_{KOH} = 56.11$ g/mol
- $m_{water} = 200.0$ g $\Rightarrow m_{solution} = 200.0 + 7.50 = 207.5$ g
- $\Delta T = +9.5$ °C (causes solution to warm)
- $c_{solution} = 4.18$ J/g °C
- Convention: $q_{solution} = +mc\Delta T$; $q_{rxn} = -q_{solution}$(exothermic dissolution)

(1) Molar enthalpy of solution, ΔH_{soln}(kJ·mol⁻¹)

a) Heat absorbed by solution

$q_{solution}$ = (207.5 g) x (4.18 J/g °C) (9.5 °C) = 8.24 10^3 J = 8.24 kJ

b) Heat released by dissolution (system)

$$q_{rxn} = -8.24 \text{ kJ}$$

c) Moles of KOH dissolved

$$n_{KOH} = 7.50 \text{ g} / 56.11 \text{ g/mol} = 0.13366 \text{ mol}$$

d) Molar enthalpy of solution

$$\Delta H_{soln} = \frac{q_{rxn}}{n_{KOH}} = \frac{-8.24 \text{ kJ}}{0.13366 \text{ mol}} = -6.16 \times 10^1 \text{ kJ/mol}$$

So $\Delta H_{soln}$is negative (exothermic).

(2) Assumptions & their effects (for Critical Thinking task)

- No heat loss to cup/air; all heat warms the solution.
  - If there *is* heat loss, $q_{solution}$is smaller than actual release → | ΔH_{soln} |would be underestimated (less negative).
- Heat capacity equals 4.18 J/g °C for the whole solution.
  - Real c may differ; if actual c is lower, then true $q_{solution}$is lower, making | ΔH_{soln} | smaller (less negative), and vice versa.

(3) Heat released for 50.0 g KOH (Extension)

$$n_{50} = 50.0 \text{ g} / 56.11 \text{ g/mol} = 0.891 \text{ mol}$$
$$q_{50} = n_{50} \text{ x } \Delta H_{soln} = (0.891 \text{ mol}) \text{ x } (-61.6 \text{ kJ/mol}) = -54.9 \text{ kJ}$$

Around 55 kJ released (exothermic). In a confined pipe, this rapid heat release can cause localized boiling/steam and thermal/chemical burns.

Key talking points for SI leaders

Likely mistakes / misconceptions

- **Forgetting solute mass** in total m (using 200.0 g instead of 207.5 g).
- **Wrong sign:** Reporting $+\Delta H_{\text{soln}}$even though solution warmed (exothermic dissolution → negative ΔH_{soln}).
- **Using water's mass-specific heat only** when a different $c$is specified.
- **Not converting to kJ·mol⁻¹:** Stopping at q(kJ) and not scaling per mole of solute.
- **Rounding too early,** which can shift $\Delta H_{\text{soln}}$by a few percent.

Why each step matters conceptually

- **Energy conservation:** $q_{\text{rxn}} = -q_{\text{solution}}$ties the thermal signal (temperature change) to the chemical process.
- **Intensive vs extensive:** $\Delta H_{\text{soln}}$is per mole, enabling fair comparison across amounts and substances.
- **Role of assumptions:** Heat losses and heat capacity choices systematically bias enthalpy estimates—recognizing this is key to sound calorimetry.

Real-world connection / safety

- **Caustic + hot:** KOH solutions are highly caustic and get hot on dissolution; both chemical and thermal burns are risks.
- **Confined spaces:** In drains/pipes, rapid heat release can generate steam, increasing pressure and causing spattering.
- **Safe demonstrations:** Choose small masses, good stirring, and PPE; pre-brief students on exothermic vs endothermic dissolution so they can "feel" the difference safely.

SI Leader Answer Key
Additional Problem 5D – Dissolving NaOH

Given / data

- $m_{\text{solution}} = 250$ g
- $c_{\text{solution}} = 4.18$ J/ g °C
- $\Delta T = +12.0$ °C (temperature increases)

Step-by-step solution

$$q_{\text{solution}} = mc\Delta T = (250)(4.18)(12.0) = 12{,}540 \text{ J} = 12.5 \text{ kJ}$$

Dissolution is exothermic here, so the solution absorbs +12.5 kJ while the process releases -12.5kJ.

Key talking points for SI leaders

- **Likely mistakes:** Forgetting units; using mass of solute instead of total solution mass; dropping the idea that exothermic dissolution ⇒ heat released by process.
- **Why steps matter:** $q = mc\Delta T$links a measured $\Delta T$to energy; clarifying system vs surroundings helps with sign conventions.
- **Real-world context:** NaOH dissolution heats rapidly—relevant to lab safety and handling of caustic solutions.

SI Leader Answer Key
Additional Problem 5E - Burning Methane

Given / data

- $\Delta H_{\text{comb}}(\text{CH}_4) = -890$ kJ/mol
- Mass burned: 5.00 g; $M_{\text{CH}_4} = 16.0$ g/mol

Step-by-step solution
Moles CH_4:

$$n = \frac{5.00}{16.0} = 0.3125 \text{ mol}$$

Heat released:

$$q = n\Delta H = (0.3125)(-890) = -278.1 \text{ kJ}$$

Key talking points for SI leaders

- **Likely mistakes:** Using grams directly with ΔH instead of converting to moles; forgetting the negative sign for exothermic combustion.
- **Why steps matter:** Thermochemical values are per mole; scaling by moles connects mass burned to energy output.
- **Real-world context:** Shows why small amounts of fuel carry substantial energy—relevant to heating value, fuel economy, and safety in combustion systems.

Problem 6A: LED Safety & Hydrogen Emission Lines

Modern LED grow lights for indoor plants often emit wavelengths similar to those produced in atomic emission spectra.

A student calibrating a spectrometer observes a hydrogen emission line produced by the transition from **n = 5 → n = 2** in a low-pressure hydrogen lamp.

Main Calculation

Using the Rydberg formula, determine the wavelength of light emitted in this transition and identify which region of the electromagnetic spectrum it belongs to.

$$\frac{1}{\lambda} = R_H\left(\frac{1}{n_1^2} - \frac{1}{n_2^2}\right) \qquad R_H = 1.097 \times 10^7\ \text{m}^{-1}$$

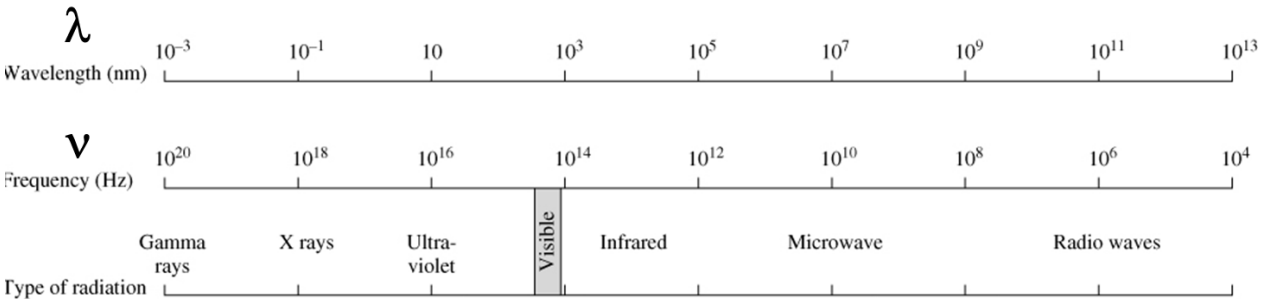
Discussion

Hydrogen emission lines (Balmer series) are used in:

- LED & laser calibration
- Astronomical spectroscopy (identifying stars & galaxies)
- Laboratory plasma diagnostics

Consider these questions:

- Why is the visible region important for both plant growth and observational astronomy?
- Should indoor agricultural technology (like high-intensity grow lights) prioritize energy efficiency, human eye safety, or maximizing plant growth rate?
- How would your answer change in contexts like urban farming, space habitats, or food-insecure regions?



Problem 6B: Solar Cell Materials & Photoelectric Threshold

Modern solar panels use materials chosen so that incoming light can eject electrons and generate electricity.

A research team is evaluating a new semiconductor material that requires 210 kJ/mol of energy to eject electrons from its surface (i.e., the work function).

Main Calculation

Calculate the longest wavelength of light (in nm) capable of ejecting electrons from this material.
Useful constants:

$$E = \frac{hc}{\lambda} \qquad h = 6.626 \times 10^{-34}\ \text{J} \qquad 3.00 \times 10^{8}\ \text{m/s} \qquad 1\ \text{mol} = 6.022 \times 10^{23}\ \text{particles}$$

Critical Thinking

Why does the longest wavelength matter for solar technology?
If sunlight has wavelengths longer than your calculated value, what happens to that portion of the spectrum?

Extension

Solar panel efficiency depends on matching material energy thresholds to sunlight's wavelength distribution.

Reflect:

- Why might a material requiring too much energy to eject electrons be inefficient for real sunlight?
- How could engineers adjust materials (or stack layers) to capture more of the solar spectrum?

Optional context clues students may mention:

- UV vs. visible vs. IR fraction of sunlight
- Multijunction solar cells
- Band gap engineering

Discussion

If you were designing solar technology for a Martian colony, where sunlight is weaker and red-shifted, how would you prioritize:

- Maximum energy capture?
- Durability in dust storms and radiation?
- Manufacturing simplicity from local Martian resources?

Problem 6C: Electron Microscopy & Quantum Wavelengths

Modern transmission electron microscopes (TEMs) rely on the wave-like behavior of electrons to achieve extremely high resolution.

A TEM accelerates electrons so they travel at 6.80×10^6 m/s.

Main Calculation

Calculate the *de Broglie wavelength* (in meters and in picometers) of an electron moving at this speed.

Useful constants:

$$\lambda = \frac{h}{mv} \quad h = 6.626 \times 10^{-34}\ \text{J s} \qquad m_e = 9.11 \times 10^{-31}\ \text{kg}$$

Extension

Electron microscopes can resolve features thousands of times smaller than those visible with light microscopes.
Consider:

- Why does a shorter wavelength allow higher resolution imaging?
- Why don't we simply accelerate electrons to arbitrarily high speeds to get even shorter wavelengths? *(Think energy use, equipment cost, radiation safety, and quantum limits.)*

Discussion

If you were designing a next-generation imaging tool for medical diagnosis, what trade-offs would matter most?

- Resolution vs. patient safety?
- Cost vs. access in low-resource settings?
- Minimizing radiation vs. maximizing detail?

How would your priorities differ for space biology labs, where shielding and power are limited?

Additional Problem 6D – Photon Energy from UV Light

A UV disinfection lamp emits light with a wavelength of 260 nm, which is effective at killing bacteria by damaging DNA.

Calculate the energy of one photon of this light in joules.

Useful information:

$$E = \frac{hc}{\lambda}, \quad h = 6.626 \times 10^{-34} \text{ J s}, \quad c = 3.00 \times 10^{8} \text{ m/s}$$

Answer: $E = 7.64 \times 10^{-19}$ *J/photon*

Additional Problem 6E – Convert Electron Wavelength to Frequency

An electron microscope uses electrons with a wavelength of 0.050 nm. Calculate the frequency of a photon that has the same wavelength (for comparison).

Useful information:

$$c = \lambda\nu, c = 3.00 \times 10^{8} \text{ m/s}$$

Answer: $\nu = 6.00 \times 10^{18} \text{ s}^{-1}$

SI Leader Answer Key
Problem 6A: LED Safety & Hydrogen Emission Lines

Step-by-step solution

Given

- Transition: $n_2 = 5 \rightarrow n_1 = 2$ (Balmer series)
- Rydberg formula: $\frac{1}{\lambda} = R_H \left(\frac{1}{n_1^2} - \frac{1}{n_2^2}\right)$
- $R_H = 1.097 \times 10^7 \text{ m}^{-1}$

Compute the bracket

$$\frac{1}{n_1^2} - \frac{1}{n_2^2} = \frac{1}{2^2} - \frac{1}{5^2} = \frac{21}{100} = 0.21$$

Find $1/\lambda$

$$\frac{1}{\lambda} = (1.097 \times 10^7 \text{ m}^{-1})(0.21) = 2.304 \times 10^6 \text{ m}^{-1}$$

Invert to get λ

$$\lambda = \frac{1}{2.304 \times 10^6} \text{ m} = 4.34 \times 10^{-7} \text{ m} = 434 \text{ nm}$$

Identify the spectral region: 434 nmlies in the visible (violet/blue) region.

Key talking points

Common mistakes / misconceptions

- **Swapping n_1 and n_2:** Always put the lower level as n_1.
- **Forgetting to invert:** Students compute $1/\lambda$ but don't take the reciprocal to find λ.
- **Unit slip:** Final answer should be in meters (or nm); $1 \text{ nm} = 10^{-9}$ m.
- **Series confusion:**
  - Lyman ($n_1 = 1$) → UV
  - Balmer ($n_1 = 2$) → Visible
  - Paschen ($n_1 = 3$) → IR

Why each step matters conceptually

- **Quantized energy levels:** The discrete λarises from specific $n$levels—evidence for quantization.
- **Selection by series:** Fixing $n_1$determines the spectral region; this connects math to real spectra.
- **Model limits:** Bohr model works well for hydrogenic systems; fails for many-electron atoms (electron–electron interactions, shielding).

Real-world connections / discussion cues

- **Instrumentation:** Balmer lines are standards for spectrometer calibration (including LED/laser setups).
- **Astronomy:** Balmer lines reveal elemental composition of astral bodies.
- **Grow lights & safety:** Visible wavelengths (esp. blue) affect plant photoreceptors and can pose eye-safety considerations; design choices balance efficiency, safety, and growth outcomes.

SI Leader Answer Key
Problem 6B: Solar Cell Materials & Photoelectric Threshold

Step-by-step solution

Given

- Work function (per mol of electrons): 210 kJ mol^{-1}
- Constants:
  $h = 6.626 \times 10^{-34}$ J s, $c = 3.00 \times 10^{8}$ m/s, $N_A = 6.022 \times 10^{23}/mol$
  Photoelectric threshold: $E_{\text{photon}} = \frac{hc}{\lambda}$

(a) Convert energy to per-photon

$$E_{\text{photon}} = \frac{210 \times 10^{3}\ \text{J mol}^{-1}}{6.022 \times 10^{23}\ \text{mol}^{-1}} = 3.486 \times 10^{-19}\ \text{J}$$

(b) Solve for the threshold (longest) wavelength

$$\lambda = \frac{hc}{E_{\text{photon}}} = \frac{(6.626 \times 10^{-34})(3.00 \times 10^{8})}{3.486 \times 10^{-19}} = 5.70 \times 10^{-7}\ \text{m} = 570\ \text{nm}$$

Region: 570 nm is in the visible (green–yellow) range.

Key talking points

Common mistakes / misconceptions

- **Per mole vs. per photon:** Students often plug 210 kJ/moldirectly into $E = hc/\lambda$. Emphasize dividing by $N_A$first.
- **"Longest" vs. "shortest" wavelength:** Longer λ = lower E. The threshold wavelength is the longest that still has enough energy.
- **Unit slips:** Finish in nm; 1 nm $= 10^{-9}$ m.
- **Rounding too early:** Keep guard digits through the calculation; round at the end.

Why each step matters conceptually

- **Band-gap/work-function link:** Threshold λ maps directly to a material's energy requirement. This is the core of photoelectric effect and semiconductor band-gap physics.
- **Spectrum matching:** Converting energy to λ lets us compare to the solar spectrum to predict efficiency losses (sub-bandgap photons pass through).

Real-world connection (to support Critical Thinking & Extension)

- **Why "longest" matters:** Photons with λ longer than 570 nm have less energy than the work function → they cannot eject electrons → that part of sunlight is not harvested (spectral loss).

- **Materials choice:** If the work function/band gap is too high, much of the visible/IR spectrum is unused; if too low, higher-energy photons cause thermalization losses.
- **Engineering strategies:** Band-gap engineering, alloying, surface treatments, or multijunction (stacked) cells tune thresholds to capture a broader fraction of sunlight.
- **Mars context:** Weaker, slightly redder sunlight + harsh environment → may favor lower band-gap absorbers, robust encapsulation, and dust-tolerant designs; trade-offs with durability and manufacturability from local resources.

SI Leader Answer Key
Problem 10C - Atmospheric Chemistry and Molecular Motion

Step-by-step solution

Given

- Speed: $v = 6.80 \times 10^6$ m/s
- Electron mass: $m_e = 9.11 \times 10^{-31}$ kg
- Planck's constant: $h = 6.626 \times 10^{-34}$ J s
- Formula (non-relativistic): $\lambda = \frac{h}{mv}$

Check relativity: $v/c \approx 6.80 \times 10^6 / 3.00 \times 10^8 \approx 0.0227 \ll 1 \rightarrow$ non-relativistic $p = mv$ is fine.

(a) Compute momentum

$$p = mv = (9.11 \times 10^{-31}\ \text{kg})(6.80 \times 10^6\ \text{m/s}) = 6.1948 \times 10^{-24}\ \text{kg m/s}$$

(b) de Broglie wavelength

$$\lambda = \frac{6.626 \times 10^{-34}}{6.1948 \times 10^{-24}}\ \text{m} = 1.07 \times 10^{-10}\ \text{m}$$

(c) Convert to picometers

$$1.07 \times 10^{-10}\ \text{m} = 1.07 \times 10^{-10}\ \text{m} \times \frac{10^{12}\ \text{pm}}{1\ \text{m}} = 1.07 \times 10^2\ \text{pm} = 107\ \text{pm}$$

Key talking points

Common mistakes / misconceptions

- **Relativistic worry when not needed:** At $v \approx 0.023c$, $\gamma \approx 1.00026$; using $p = \gamma mv$ barely changes λ. Good chance to discuss when relativity matters.
- **Unit slip in pm conversion:** Students sometimes divide by 10^{12}instead of multiplying; remember 1 pm $= 10^{-12}$ m.
- **Rounding too early:** Keep guard digits in $p$and λ to avoid rounding drift.
- **Using energy route incorrectly:** Some try $E = \frac{1}{2}mv^2$and then $p = E/c$ (not valid for massive particles). Stick with $p = mv$ here.

Why each step matters conceptually

- **Wave–particle duality in action:** de Broglie ties momentum to wavelength; higher momentum → shorter λ→ better resolving power.
- **Resolution link:** Imaging resolution scales with ~ λ; electrons (Å-scale wavelengths) beat visible light (hundreds of nm).

Real-world connection

- **Why shorter λ helps:** Shorter wavelengths diffract less around small features, enabling atomic-scale imaging in TEMs.
- **Why not "infinite" acceleration:** Diminishing returns (relativity increases effective mass), higher cost/energy, radiation safety, lens aberrations, and sample damage. Practical TEM design balances resolution with beam energy, dose limits, and instrument complexity.

SI Leader Answer Key
Additional Problem 6D – Photon Energy from UV Light

Given:

- $\lambda = 260 \text{ nm} = 260 \times 10^{-9} \text{ m} = 2.60 \times 10^{-7} \text{ m}$
- $h = 6.626 \times 10^{-34} \text{ J s}$
- $c = 3.00 \times 10^{8} \text{ m/s}$

Formula:

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

Substitute:

$$E = \frac{(6.626 \times 10^{-34})(3.00 \times 10^{8})}{2.60 \times 10^{-7}}$$

Calculate numerator:

$$6.626 \times 10^{-34} \times 3.00 \times 10^{8} = 1.9878 \times 10^{-25}$$

Divide:

$$E = \frac{1.9878 \times 10^{-25}}{2.60 \times 10^{-7}} = 7.64 \times 10^{-19} \text{ J}$$

Final Answer: 7.64×10^{-19} J/photon

Key SI Teaching Notes

- Common error: forgetting to convert nm → m.
- Emphasize that shorter wavelength → higher energy (UV > visible light), which is why UV kills microbes.
- Real-world note: UV-C (≈260 nm) is used for water purification & hospital sterilization.

SI Leader Answer Key
Additional Problem 6E – Convert Electron Wavelength to Frequency

Given:

- $\lambda = 0.050 \text{ nm} = 0.050 \times 10^{-9} \text{ m} = 5.0 \times 10^{-11} \text{ m}$
- $c = 3.00 \times 10^{8} \text{ m/s}$

Formula:

$$c = \lambda\nu \Rightarrow \nu = \frac{c}{\lambda}$$

Substitute:

$$\nu = \frac{3.00 \times 10^{8}}{5.0 \times 10^{-11}}$$

Divide:

$$\nu = 6.00 \times 10^{18} \text{ s}^{-1}$$

Final Answer: 6.00 x 10^{18} s^{-1}

Key SI Teaching Notes

- Students often confuse units — reinforce nm → m conversion.
- The point is comparison: electron wavelength is like extremely high-frequency photons → high resolution.
- Connect to microscopy: more energy → shorter wavelength → smaller details visible.

Problem 8/9 A — Lewis Structures in Fertilizer & Soil Chemistry

Modern agriculture depends on specific chemical forms of nutrients that plants can absorb from soil. Decades of research have identified key ions that support plant growth and global food production. These ions also play important roles in soil buffering, environmental chemistry, and nutrient cycles.

Below are six common ions found in fertilizers or soil systems:

Ion	Role in agriculture / soil chemistry
NH_4^{+1} (ammonium)	Major nitrogen fertilizer; absorbed by roots and retained in soil
NO_3^{-1} (nitrate)	Highly mobile nitrogen source for plants; excess can contaminate water
NO_2^{-1} (nitrite)	Intermediate in nitrogen cycle; indicator of soil microbial activity
$H_2PO_4^{-1}$ (dihydrogen phosphate)	Primary plant-available phosphorus form; essential for DNA/ATP
SO_4^{-2} (sulfate)	Plant sulfur source for amino acids & proteins
HCO_3^{-1} (bicarbonate)	Key soil buffering ion; affects nutrient availability and pH balance

These ions support the food we eat, the economy that grows it, and the environmental systems that sustain it.

Draw the complete Lewis structures for each species above.
Show:

- All lone pairs and bonds
- Formal charges on all atoms (label or circle)
- Brackets and net charge for each ion
- Resonance structures where applicable (NO_3^-, NO_2^-, SO_4^{2-}, HCO_3^-)

Consider the following:

- Count total valence electrons first (add/subtract for the ionic charge)
- Typical valences: H 1, O 2, N 3, P 5, S 6, C 4 (but use formal charge to choose the best structure)
- Check octets and minimize formal charge magnitude; place negative charge preferentially on more electronegative atoms (usually O).

Problem 8/9 B — VSEPR: Electron-Domain & Molecular Geometries in Fertilizer Ions

You must use your completed Lewis structures from Problem 8A to answer this problem.

For each species below, determine:

1. Central atom
2. Electron Groups: this is the steric number (SN) around the central atom
3. Number of bonding groups and lone pairs on the central atom
4. Electron-domain (electron-group) geometry (VSEPR)
5. Molecular geometry (shape)

Species (same list as 8A):
NH_4^+, NO_3^-, NO_2^-, $H_2PO_4^-$, SO_4^{2-}, HCO_3^-

Species	Central Atom	Electron Groups (SN)	Bonding Groups	Lone Pairs	Electron-Domain Geometry	Molecular Geometry
NH_4^+						
NO_3^-						
NO_2^-						
$H_2PO_4^-$						
SO_4^{2-}						
HCO_3^-						

Notes: Assume typical VSEPR ideal angles; when resonance is present (e.g., NO_3^-), note that equivalent bonds imply equal average bond orders and symmetric shapes. For atoms with X–H bonds (e.g., –OH), remember H does not affect electron-domain count beyond the bond it completes.

Critical Thinking

1. For NO_3^- and SO_4^{2-}, how does resonance influence observed bond lengths and the symmetry of the shape?

2. In $H_2PO_4^-$ and HCO_3^-, which atoms are most likely to carry negative charge in the best Lewis structures, and how does that align with the observed shapes?

Extension

- Mobility & leaching: Planar, symmetric anions like NO_3^- are highly mobile in soils and groundwater. How might geometry + charge distribution contribute to this behavior?
- Hydrogen bonding & acidity: The –OH groups in $H_2PO_4^-$ and HCO_3^- can donate/accept H-bonds. How do their local geometries help explain buffering (pH stabilization) in soils?
- Plant transport/recognition: Ion channels and transporters often "recognize" ions by size, charge, and shape. Based on your geometries, which of the six would you expect to interact strongly with H-bonding sites on proteins?

Discussion

Fertilizer design balances plant uptake, soil health, and water quality:

- Should policy incentivize nutrient forms that are less mobile (reduced leaching) even if they are harder to absorb quickly?
- How would your recommendations change for arid soils vs humid regions, or for small farms vs industrial-scale agriculture?

Justify your reasoning using molecular geometry, charge distribution, and resonance from your analyses.

Useful Information

Electron Groups*	Bonding Groups	Lone Pairs	Electron Geometry	Molecular Geometry
2	2	0	Linear	Linear
3	3	0	Trigonal planar	Trigonal planar
3	2	1	Trigonal planar	Bent
4	4	0	Tetrahedral	Tetrahedral
4	3	1	Tetrahedral	Trigonal pyramidal
4	2	2	Tetrahedral	Bent

Electron Groups*	Bonding Groups	Lone Pairs	Electron Geometry	Molecular Geometry
5	4	1	Trigonal bipyramidal	Seesaw
5	3	2	Trigonal bipyramidal	T-shaped
5	2	3	Trigonal bipyramidal	Linear
6	6	0	Octahedral	Octahedral
6	5	1	Octahedral	Square pyramidal
6	4	2	Octahedral	Square planar

Additional Problem 8/9 C – Lewis, Electron-Domain, and Molecular Geometries

For each species below, do the following:

1. Draw the best Lewis structure (show all bonds and lone pairs).
2. Identify the electron-domain (electron-group) geometry using VSEPR.
3. Identify the molecular geometry (shape).

Species: SF_4 and IF_5

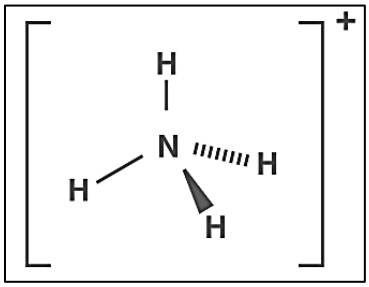
Consider the following:

- Count total valence electrons first.
- Steric number (SN) = number of σ-bonding regions + lone pairs on the central atom.
- Lone pairs occupy positions that minimize repulsions (in trigonal bipyramidal domains, lone pairs prefer equatorial sites).

Answers
SF_4 → Seesaw
IF_5 → Square pyramidal

SI Leader Answer Key
Problem 8/9 A — Lewis Structures in Fertilizer & Soil Chemistry

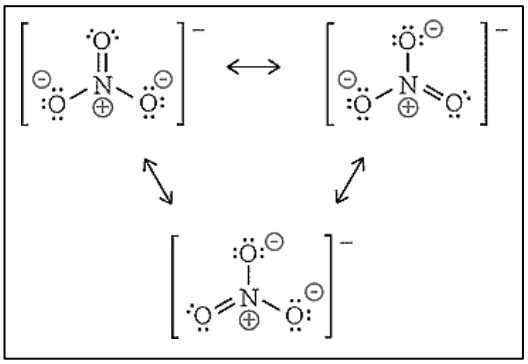
Step-by-step solution



1) Ammonium, NH_4^+
Valence electrons: N(5) + 4×H(1) − 1 (charge) = 8 e⁻
Structure: Central N with four N–H single bonds; no lone pairs on N. Brackets with +1.
Formal charges:

- N: $5 - (0 + 8/2) = +1$
- Each H: $1 - (0 + 2/2) = 0$

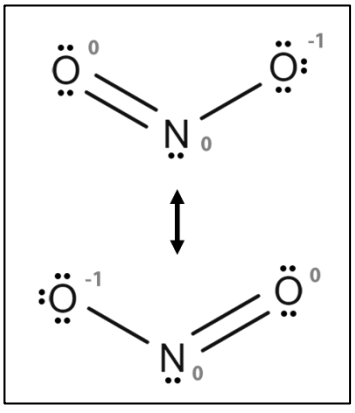
Notes: Tetrahedral electron/molecular geometry (VSEPR). No resonance.



2) Nitrate, NO_3^-
Valence electrons: N(5) + 3×O(6) + 1 (charge) = 24 e⁻
Structure (one resonance form): with one N=O and two N–O single (each O has octet). Brackets with –1.
Formal charges (in any single resonance form):

- N: $5 - (0 + 8/2) = +1$
- Double-bond O: $6 - (4 + 4/2) = 0$
- Single-bond O^-: $6 - (6 + 2/2) = -1$

Resonance: 3 equivalent forms (double bond rotates among O's). Average N–O bond order ≈ 1.33.
Notes: Trigonal planar around N (VSEPR).

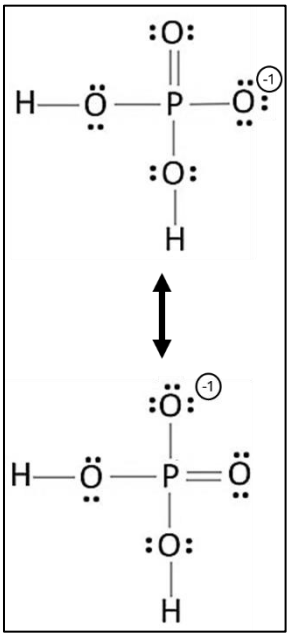


3) Nitrite, NO_2^-
Valence electrons: N(5) + 2×O(6) + 1 (charge) = 18 e⁻
Structure (one resonance form):
$[\, O = N - O^- \,]$ with one N=O, one N–O single, and one lone pair on N. Brackets with –1.
Formal charges (in one form):

- N: $5 - (2 + 6/2) = 0$
- Double-bond O: $6 - (4 + 4/2) = 0$
- Single-bond O^-: $6 - (6 + 2/2) = -1$

Resonance: 2 equivalent forms (double/single swap between O's). Average N–O bond order ≈ 1.5.
Notes: Electron-domain geometry trigonal planar; molecular geometry bent (lone pair on N).

4) Dihydrogen phosphate, $H_2PO_4^-$



Valence electrons: P(5) + 4×O(6)=24 + 2×H(1)=2 + 1 (charge) → 32 e⁻
Structure (one low-charge form):
Central P with four P–O bonds: two P–OH singles, one P=O, one P–O⁻.
Brackets with –1. All atoms with octets.
Formal charges (this resonance form):

- P: $5 - (0 + (4 + 2 + 2 + 2)/2) = 5 - 5 = 0$
- O = P(double-bond O): $6 - (4 + 4/2) = 0$
- Each OH oxygen: $6 - (4 + 4/2) = 0$(bonded to P and H)
- O^- (single to P, no H): $6 - (6 + 2/2) = -1$

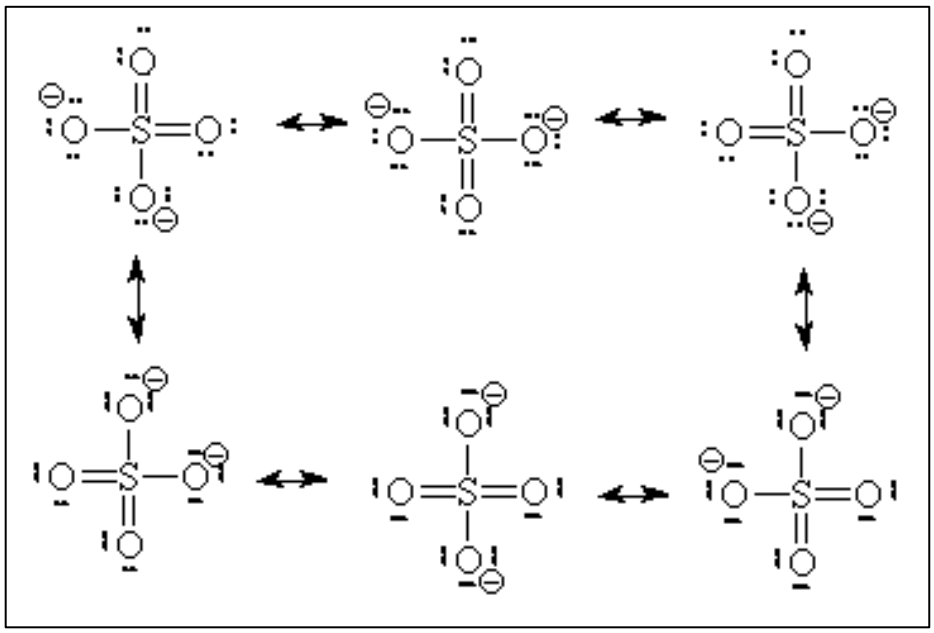
Resonance: The P=O and P–O⁻ positions can swap among the two non-protonated O atoms, giving 2 major resonance forms (more if you allow different P=O placements while keeping both OH groups present).
Notes: Around P, 4 electron domains → tetrahedral e⁻-domain geometry.

5) Sulfate, SO_4^{2-}

Valence electrons: S(6) + 4×O(6)=24 + 2 (charge) → 32 e⁻
Common textbook representations:

- All single bonds: $[\ O^- - S - O^-\ ]$with two additional S–O single bonds (all O octets).
  Formal charges: S: $+2$; each O: -1. Sum = −2.
- Lower formal-charge set (expanded octet on S): two S=O and two S–O⁻ (resonance among which O's are double/single).
  Formal charges (per one form): S: 0; each S=O O: 0; each S–O⁻ O: −1. Sum = −2.

Best description: Resonance hybrid with electron density delocalized; experimentally all S–O bonds are similar (avg. bond order between 1 and 2). Brackets with –2.
Notes: 4 electron domains around S → tetrahedral e⁻-domain geometry.



6) Bicarbonate, HCO_3^-

Valence electrons: H(1) + C(4) + 3×O(6)=18 + 1 (charge) → 24 e^-

Structure (one resonance form):

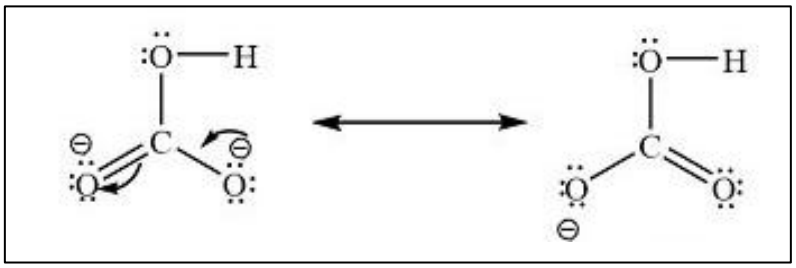
$[\ HO-C(=O)-O^-\ ]$: carbonyl C=O, one C–OH, one C–O^-. Brackets with –1.

Formal charges (this form):

- C: $4-(0+8/2)=0$
- C=O oxygen: $6-(4+4/2)=0$
- OH oxygen: $6-(4+4/2)=0$(bonded to C and H)
- O^- oxygen: $6-(6+2/2)=-1$

Resonance: The double bond/negative charge delocalizes between the two non-hydrogen-bonded oxygens → 2 major resonance forms (the OH remains single to C).

Notes: Around C, 3 electron domains in the carboxylate part → trigonal planar local geometry.



Teaching & Guiding Notes

- Workflow to model:
    1. Count valence e^- (adjust for ion charge).
    2. Choose a sensible skeleton (central atom least electronegative except H).
    3. Complete octets on terminal atoms, then central atom.
    4. Add multiple bonds if needed to satisfy octets/minimize formal charge (FC).
    5. Compute formal charges; prefer structures with lowest FC and negative charge on more electronegative atoms (O).
    6. Draw resonance where applicable.

- Frequent student errors:
    - Forgetting to add/subtract electrons for the ionic charge.
    - Placing H as a central atom or giving H more than 2 e^-.
    - Missing lone pairs, then miscounting octets.
    - Not bracketing ions / omitting net charge.
    - Confusing “moves atoms” with resonance; only electrons move in resonance.

- Concept bridges to next topics:
    - These correct Lewis structures set up VSEPR geometries, bond polarity, formal-charge reasoning, resonance stabilization, and later acid–base behavior (e.g., $H_2PO_4^-/HPO_4^{2-}$ pair, HCO_3^- buffering).

SI Leader Answer Key
Problem 8/9 B — VSEPR: Electron-Domain & Molecular Geometries in Fertilizer Ions

Species	Central Atom	Electron Groups (SN)	Bonding Groups	Lone Pairs	Electron-Domain Geometry	Molecular Geometry (Shape)
NH_4^+	N	4	4	0	Tetrahedral	Tetrahedral *(~109.5°)*
NO_3^-	N	3	3	0	Trigonal planar	Trigonal planar *(120°; all N–O equivalent by resonance)*
NO_2^-	N	3	2	1	Trigonal planar	**Bent** *(<120° due to LP compression; two equivalent N–O by resonance)*
$H_2PO_4^-$	P	4	4	0	Tetrahedral	Tetrahedral *(~109.5°; mix of P=O, P–OH, P–O^- across resonance)*
SO_4^{2-}	S	4	4	0	Tetrahedral	Tetrahedral *(~109.5°; all S–O similar in resonance hybrid)*
HCO_3^-	C	3	3	0	Trigonal planar	Trigonal planar *(~120°; C=O / C–O^- resonance over non-OH oxygens)*

Notes & reminders
- "Electron Groups (SN)" = total regions around central atom (σ bonds + lone pairs).
- X–H bonds count as one bonding group each; H never contributes a lone pair.
- Resonance makes bond orders fractional and often equalizes bond lengths, but does not change the local VSEPR count.

Brief Justifications (by species)
- NH_4^+: N has 4 σ bonds, 0 LP → AX_4E_0 → tetrahedral / tetrahedral.
- NO_3^-: N has 3 σ to O, 0 LP → AX_3E_0 → trigonal planar / trigonal planar; resonance among three N=O placements gives equal N–O lengths.
- NO_2^-: N has 2 σ to O + 1 LP → AX_2E_1 → trigonal planar e-domain, bent molecular; resonance between two O's.
- $H_2PO_4^-$: P bonded to four O (two –OH, one =O, one –O^- in a given form) → AX_4E_0 → tetrahedral / tetrahedral; resonance swaps which non-protonated O is =O vs O^-.

- SO_4^{2-}: S with four S–O bonds → AX_4E_0 → tetrahedral / tetrahedral; resonance yields similar S–O bonds overall.
- HCO_3^-: Central C with three σ to O (one –OH, one =O, one $–O^-$ in a given form) → AX_3E_0 → trigonal planar / trigonal planar; resonance between the two non-OH O atoms.

Answers to Critical Thinking

1. **Resonance in NO_3^- and SO_4^{2-}:**
   Resonance delocalizes π electrons over multiple X–O bonds, making the bonds equivalent (same average bond order) and preserving high symmetry (NO_3^- trigonal planar; SO_4^{2-} tetrahedral). Observed bond lengths are intermediate between single and double bonds and nearly equal.
2. **Negative charge placement in $H_2PO_4^-$ and HCO_3^-:**
   Negative charge resides primarily on oxygen atoms (most electronegative). In $H_2PO_4^-$, the non-protonated O often bears the −1 FC in a given resonance form; in HCO_3^-, the −1 is on the non-OH oxygen. These placements align with the drawn shapes (tetrahedral around P; trigonal planar around C) and with acidity (–OH groups as Brønsted donors).

Extension Talking Points

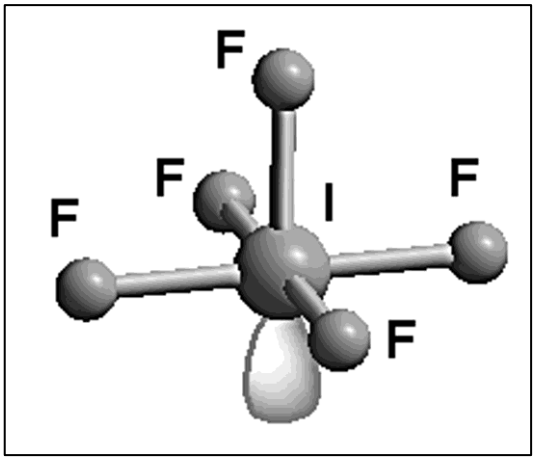
- **Mobility & leaching (NO_3^-):** Planar, symmetric shape and charge delocalization reduce strong site-specific binding in soils → greater solubility and mobility, hence leaching risk.
- **Hydrogen bonding & buffering ($H_2PO_4^-$, HCO_3^-):** Local geometries place –OH and lone pairs in orientations favorable for H-bonding and proton transfer, supporting buffer systems in soils and waters.
- **Protein/transport recognition:** Ions with exposed negative O sites (NO_3^-, $H_2PO_4^-$, SO_4^{2-}, HCO_3^-) can engage in H-bonding / electrostatic interactions with transporter residues; geometry dictates complementarity.

Tips

- Emphasize the difference between electron-domain vs molecular geometry (Lone pairs remove from the "visible" shape).
- When students get stuck, have them recount electron groups on the central atom directly from their Problem 8/9A structures.
- For resonance cases, make students state: "VSEPR is local and counts σ-framework only; resonance ≠ more domains."

SI Leader Answer Key
Additional Problem 6D – Photon Energy from UV Light

SF_4

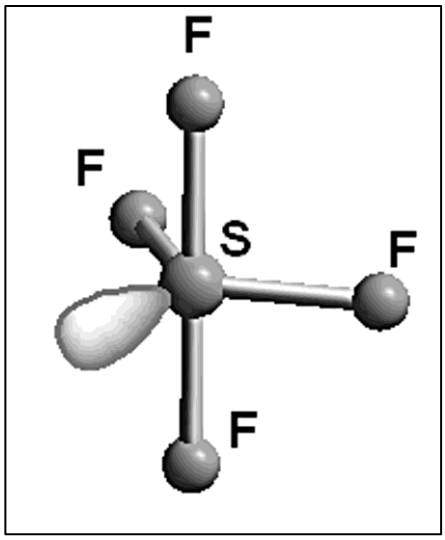


Valence e^- count: S (6) + 4×F (4×7 = 28) = 34 e^-
Lewis structure:

- Central S with four S–F single bonds (8 e^-).
- Place remaining electrons to complete F octets (4 F × 6 nonbonding e^- = 24 e^-).
- Used so far: 8 + 24 = 32 e^- → 2 e^- left → one lone pair on S.
- Total around S: 4 σ bonds + 1 LP → SN = 5.

VSEPR (electron-domain geometry): Trigonal bipyramidal (TBP).
Molecular geometry: Seesaw (AX_4E).
Lone pair placement: Equatorial (minimizes 90° LP–bond repulsions).
Angle notes: ~120° and 90° in the TBP framework, with deviations (axial S–F bonds slightly longer, smaller than 180° due to LP compression).

IF_5



Valence e^- count: I (7) + 5×F (5×7 = 35) = 42 e^-
Lewis structure:

- Central I with five I–F single bonds (10 e^-).
- Complete F octets (5 F × 6 nonbonding e^- = 30 e^-).
- Used so far: 10 + 30 = 40 e^- → 2 e^- left → one lone pair on I.
- Total around I: 5 σ bonds + 1 LP → SN = 6.

VSEPR (electron-domain geometry): Octahedral.
Molecular geometry: Square pyramidal (AX_5E).
Lone pair placement: On one octahedral vertex opposite the apex (any one site; square base remains).
Angle notes: Base angles ~90°; apex-base ~90°; slight deviations due to LP compression.

Quick Summary Table

Species	Valence e^-	SN (Central)	Electron-Domain Geometry	Molecular Geometry
SF_4	34	5	Trigonal bipyramidal	Seesaw
IF_5	42	6	Octahedral	Square pyramidal

Tips:

- Emphasize SN counting and lone pair placement (equatorial in TBP).
- Remind that Lewis ≠ geometry: you must go from Lewis → SN → VSEPR to get the correct 3D shape.
- Ask students to sketch the 3D shapes and mark the lone pair position explicitly.

Problem 10A - Emergency O_2 "Candle" for a Submersible

In some submersibles, oxygen candles are used as backups to supply breathable O_2 during emergencies. One common candle relies on the thermal decomposition of potassium chlorate:

$$2\ KClO_3(s) \rightarrow 2\ KCl(s) + 3\ O_2(g)$$

A maintenance team tests a candle by heating 475 g of $KClO_3(s)$ in a sealed steel tank until it fully decomposes. The tank pressure is held at 1.80 atm, and the gas temperature stabilizes at 650 °C.

1. Main calculation
How many liters of $O_2(g)$are produced in the tank under these conditions?
Assume ideal-gas behavior; use R = 0.08206 L atm / mol K show the stoichiometry from $KClO_3$ to O_2 before applying $PV = nRT$.

2. Critical thinking
- Identify two assumptions in your calculation (about purity, reaction completeness, gas behavior, leaks, etc.).
- For one assumption, explain how violating it would make your calculated volume an overestimate or an underestimate and why.

3. Extension
A resting adult uses about 250 mL O_2 per minute. If this candle run produced the O_2 volume you calculated, about how many person-minutes of breathable O_2 would it supply (ignoring mixing/CO_2 removal)? What other systems (beyond O_2 generation) are needed to keep the air safe in an enclosed space?

4. Discussion prompt
Oxygen candles are compact and reliable, but they run hot and can't be switched "off" once started. In designing life-support backups for submarines, spacecraft, or mines, how would you balance simplicity, controllability, heat management, and storage mass/volume? What trade-offs would you accept?

Problem 10B - Hospital Oxygen Cylinder Leak Detection

A hospital keeps backup oxygen cylinders for patient care. A steel cylinder has an internal volume of 10.0 L. At the start of a shift, a cylinder reads 145 atm at 22 °C. Later, staff notice the pressure has fallen to 130 atm (the valve was slightly open). Assume the cylinder volume stays constant and the gas behaves ideally at a constant temperature.

1. Main calculation
How many moles of O_2 were lost from the cylinder during the leak?
Use R = 0.08206 L·atm / mol·K; treat Tas constant at 22 °C.

2. Critical thinking
- Identify two assumptions you made (e.g., constant temperature, ideal behavior, gauge accuracy).
- If the cylinder actually cooled during the leak (Joule–Thomson effect), would your constant-*T* calculation overestimate or underestimate the moles lost? Explain briefly.

3. Extension
A patient on low-flow oxygen uses about 2.0 L O_2 per minute at STP. Convert the moles of O_2 lost into liters at STP (use 1.00 mol = 22.4 L at STP) and estimate how many minutes of therapy were lost because of the leak. What practical steps could staff take to prevent or detect such losses?

4. Discussion prompt
Hospitals balance safety, cost, and reliability in oxygen delivery (high-pressure cylinders vs. liquid O_2 vs. on-site concentrators). Given your analysis, how would you prioritize monitoring, leak prevention, and system redundancy? What trade-offs would you accept for critical care units versus general wards?

Problem 10C - Atmospheric Chemistry and Molecular Motion

Ozone (O_3) in the upper atmosphere absorbs harmful ultraviolet radiation, converting solar energy into molecular motion (heat). Scientists sometimes estimate temperature from the average kinetic energy of molecules using the relationship:

$$E_{\text{avg}} = \frac{3}{2}RT$$

where E_{avg} is the average kinetic energy per mole of gas molecules and R = 8.314 J / mol K.

1. Main calculation
A stratospheric sample of ozone gas has an average kinetic energy of 2.49 kJ/mol. What is the temperature (in K) of this ozone sample?

2. Critical thinking
- If the same average kinetic energy were measured for nitrogen (N_2) instead of ozone, would the temperature be higher, lower, or the same? Why?
- What does this imply about the relationship between temperature and molecular mass?

3. Extension
The ozone layer absorbs solar UV energy and warms the upper atmosphere. Suppose a section of air (1.0 mol of O_3 and 9.0 mol of N_2) has its average kinetic energy increase by 0.50 kJ/mol due to UV absorption.
- How much total energy (in kJ) is absorbed by this 10.0 mol mixture?
- What could this mean for atmospheric circulation and temperature gradients near the ozone layer?

4. Discussion prompt
- If global ozone concentrations were to increase or decrease significantly, how might that affect stratospheric temperature and weather patterns on Earth's surface?
- How should scientists communicate such indirect but important links between chemistry and climate to the public?

Additional Problem 10D – Boyle's Law and Diving Pressure

A scuba diver's air bubble has a volume of 2.50 L at the surface (1.00 atm). What will be the volume of the bubble when it rises to a region where the pressure is 0.850 atm, assuming constant temperature?

Answer: 2.94 L

Additional Problem 10E - Gas Density and Molar Mass

A gas sample has a density of 1.43 g/L at STP. What is the molar mass of the gas?
Hint: Use density $= \frac{PM}{RT}$

Answer: 31.9 g/mol (approximately O_2)

SI Leader Answer Key
Problem 10A – Emergency O_2 "Candle" for a Submersible

Step-by-step solution

Data

- Balanced reaction:

$$2\ KClO_3(s) \rightarrow 2\ KCl(s) + 3\ O_2(g)$$

  Mass $KClO_3$= 475 g
- Pressure P= 1.80 atm
- Temperature T= 650 °C = 923 K (use 650 + 273)
- Gas constant R = 0.08206 L·atm /mol·K
- Molar mass $KClO_3$ = 122.55 g

(1) Moles of $KClO_3$

$$n_{KClO3} = 475\ g / 122.55\ g/mol = 3.88\ mol$$

(2) Moles of O_2 from stoichiometry

From 2 KCl → 3 O_2:

$$n_{O2} = 3/2\ n_{KClO3} = 1.5 \times 3.88\ mol = 5.82\ mol$$

(3) Volume of O_2 at 1.80 atm and 923 K (ideal gas)

$$PV = nRT \ \Rightarrow\ V = \frac{nRT}{P}$$

$$V = [\ (5.82\ mol) \times (0.08206\ L{\cdot}atm\ /mol{\cdot}K) \times (973\ K)] / (1.80\ atm) = 2.45 \times 10^2\ L$$

(4) Extension — person-minutes of O_2 at 0.250 L·min^{-1}

$$person\text{-}minutes = V / (0.250\ L / min) = 2.45 \times 10^2\ L / (0.250\ L / min) = 9.8 \times 10^2\ min$$

980 min per person (~16.3 hours), ignoring mixing, CO_2 removal, pressure effects, other.

Key talking points for SI leaders

Likely mistakes / misconceptions

- Stoichiometry slip: Using 1:1 instead of the correct 2:3 ratio ($KClO_3$:O_2).
- Temperature units: Forgetting to convert °C → K for $PV = nRT$.
- Unit consistency for R: Mixing units (e.g., using R in kPa or J with atm and L).
- Significant figures: Over-precision or under-precision; inputs support about 3 s.f.
- Assumptions ignored: Treating non-ideal gas at 923 K / 1.80 atm or incomplete decomposition as negligible without comment.

Why each step matters conceptually

- Mass → moles → moles (stoich): Reinforces that balanced equations are the bridge from a solid's mass to gaseous product amount.
- State equation application: Shows how amount, temperature, and pressure set gas volume. A direct PV-work and containment consideration.
- Scaling to use-rate: Translating liters of O_2 to person-minutes connects chemical output to operational capacity.

Real-world context / limitations to note

- Safety & systems integration: Oxygen generation alone is insufficient; real life-support also requires CO_2 scrubbing, humidity/temperature control, and heat management (O_2 candles run hot).
- Assumptions to surface: Purity of $KClO_3$, complete reaction, no leaks, ideal-gas behavior, and accurate pressure/temperature control.
    - If incomplete reaction or leaks occur: n_{O_2} is lower → volume is overestimated.
    - If gas deviates from ideality or P/T misread: computed V may be biased (at these conditions, ideal assumption is usually acceptable for first pass).

SI Leader Answer Key
Problem 10B - Hospital Oxygen Cylinder Leak Detection

Step-by-step solution

Given / data

- Cylinder volume V = 10.0 L (constant)
- Initial pressure $P_1 = 145$ atm
- Final pressure $P_2 = 130$ atm
- Temperature T = 22 °C = 295 K (constant)
- Gas constant R = 0.08206 L·atm /mol·K
- Ideal-gas behavior, isothermal

At constant V, T: $n \propto P$. Moles lost correspond to the pressure drop $\Delta P = P_1 - P_2$.

a) Moles lost (isothermal ideal gas):

$$\Delta n = \frac{\Delta P\, V}{R\, T}$$

$$\Delta n = (\Delta P \text{ x } V) / (R \text{ x } T)$$

$$\Delta n = [\ (145\text{-}130) \text{ atm x } 10.0 \text{ L }] / (0.08206 \text{ x } 295 \text{ K}) = 150 / 24.21 = 6.20 \text{ mol } O_2$$

b) Convert to liters at STP (Extension part foundation):
Use 1 mol = 22.4 L at STP.

$$V = 6.20 \text{ mol x } (22.4 \text{ L} / 1 \text{ mol}) = 1.389 \text{ x } 10^2 \text{ L} = 139 \text{ L}$$

c) Minutes of therapy at 2.0 L·min⁻¹ (Extension):

$$t = 139 \text{ L} / (2.0 \text{ L} / \text{min}) = 69.5 \text{ min}$$

Key talking points for SI leaders

Likely mistakes / misconceptions

- Forgetting isothermal proportionality: Students may try to use full $PV = nRT$ twice; emphasize that at constant V, T, n is proportional to P, and using ΔPdirectly is efficient and correct.
- Unit handling: Not converting $T$to Kelvin; mixing units inconsistent with R.
- Gauge vs. absolute pressure: Ideal gas law uses absolute pressure. With the same gauge for both readings, the difference $\Delta P$is unaffected by the +1 atm offset, but this is worth noting.
- Rounding too early: Keep guard digits through intermediate steps.

Why each step matters conceptually

- Proportional reasoning at constant V, T: This is a classic application of the ideal gas law showing how mole count tracks pressure when volume and temperature don't change.
- Link to patient care: Converting moles → liters at STP then to a flow-rate time grounds abstract quantities in a clinically meaningful metric (minutes of therapy).
- Assumption awareness: The calculation's validity hinges on isothermal, ideal behavior, and accurate gauges. Surfacing these helps students reason about uncertainty and bias.

Addressing the Critical Thinking prompt

- Two assumptions: (i) Temperature truly constant; (ii) Ideal-gas behavior holds at high P; (iii) No leaks other than the valve and no regulator/line volume effects; (iv) Gauge is accurate and stable.
- Joule–Thomson cooling effect: If the cylinder cooled during the leak, the observed pressure would be lower than it would be at the same $n$and constant T. Using a constant-T model would then overestimate moles lost (since part of ΔPcomes from ΔT, not Δn).

Real-world connection

- Operations & safety: Even a small leak can cost ~70 minutes of low-flow oxygen. Significant in clinical settings. This supports policies for valve checks, leak detection (soapy water, ultrasonic, pressure decay tests), and inventory monitoring.
- System choice trade-offs: The result informs discussions comparing high-pressure cylinders, liquid O_2 dewars, and on-site concentrators in terms of reliability, maintenance, and redundancy for critical care.

SI Leader Answer Key
Problem 10C - Atmospheric Chemistry and Molecular Motion

Step-by-step solution

Given / data

- Relationship: $E_{\text{avg}} = \frac{3}{2}RT$
- R = 8.314 J/ mol·K
- For part (1): E_{avg} = 2.49 kJ/ mol = 2490 J / mol

(1) Temperature from average kinetic energy

Solve $E_{\text{avg}} = \frac{3}{2}RT$ for T:

$$T = 2/3 \times E_{\text{avg}} / R = 2/3 \times (2490 \text{ J/mol})/ (8.314 \text{ J/mol·K}) = 200. \text{ K}$$

(2) Critical thinking

- Same E_{avg} for N_2 → same T.
  Because $E_{\text{avg}} = \frac{3}{2}RT$ depends only on T (ideal gas), not on molecular identity.
- Implication: Temperature is a measure of average translational kinetic energy; molecular mass does not change E_{avg} at a given T. Heavier molecules move slower on average at the same T, but have the same average KE as lighter ones.

(3) Extension — energy absorbed by a 10.0-mol air parcel

Mixture: 1 mol O_3 + 9 mol N_2 = 10.0 mol
Increase in average kinetic energy: 0.50 kJ/mol
Total absorbed energy:

$$q_{\text{total}} = 10.0 \text{ mol} \times 0.50 \text{ kJ/mol} = 5.0 \text{ kJ}$$

Interpretation: extra KE in the stratosphere warms that layer, affecting temperature gradients and thus circulation/stability (e.g., stronger stratification, altered vertical mixing).

Key talking points for SI leaders

Likely mistakes / misconceptions

- Unit slip (kJ → J): Forgetting to convert 2.49 kJ/mol to J/mol before plugging into $E_{\text{avg}} = \frac{3}{2}RT$.
- Wrong formula: Using $E = \frac{1}{2}mv^2$ per molecule without Avogadro's number or mixing per-molecule and per-mole quantities.
- Mass dependence confusion: Assuming heavier gases must have higher T for the same KE. (At a given T, all ideal gases have the same E_{avg}; masses affect speeds, not KE at fixed T.)
- Sig figs: Over-reporting precision; 2.49 kJ/mol → 200 K

Why each step matters conceptually

- Thermal definition: $E_{\text{avg}} = \frac{3}{2}RT$ ties temperature directly to molecular motion, a core idea of kinetic theory.
- Identity independence: Reinforces that temperature is an average energy property of the ensemble, not of a particular molecule's mass.
- Scaling energy to systems: Multiplying per-mole energy by moles connects microscopic kinetic theory to macroscopic heat content.

Real-world context / climate link

- Stratospheric temperatures: A value near 200 K is realistic for the lower stratosphere, anchoring the calculation in reality.
- Ozone & heating: UV absorption by ozone adds kinetic energy aloft, shaping temperature profiles and stability (e.g., temperature inversions that limit vertical mixing).
- Communication angle: Encourage students to articulate how a molecular-level energy change scales up to layer heating, influencing circulation and potentially surface weather patterns—and why that's important for public understanding of atmospheric chemistry.

SI Leader Answer Key
Additional Problem 10D – Boyle's Law and Diving Pressure

Given / data

- $P_1 = 1.00$ atm
- $V_1 = 2.50$ L
- $P_2 = 0.850$ atm
- $T =$ constant

$$P_1V_1 = P_2V_2$$

Solve for V_2:

$V_2 = P_1 \times V_1 / P_2 = (1.00 \text{ atm}) \times (2.50 \text{ L}) / (0.850 \text{ atm}) = 2.94 \text{ L}$

Key talking points for SI leaders

- **Common student errors:**
  - Forgetting the inverse relationship between $P$and V.
  - Incorrect setup of the equation (e.g., using $P_2V_2 = P_1V_1$but mixing which side has higher pressure).
  - Forgetting to keep units consistent (both pressures must be in atm, volume in liters).
- **Conceptual emphasis:**
  - As the bubble rises (pressure decreases), volume increases — a direct application of Boyle's Law.
  - The law holds only when temperature and amount of gas are constant (isothermal conditions).
- **Real-world context:**
  - In scuba diving, expanding bubbles can be dangerous — gas expansion in lungs or equipment can cause injury.
  - This reinforces why divers must ascend slowly and release air during ascent.

SI Leader Answer Key
Additional Problem 10E - Gas Density and Molar Mass

Given / data

- d = 1.43 g/L
- $P = 1.00$ atm
- $T = 273$ K
- R = 0.0821 L atm / mol K

Relationship:

$$d = \frac{PM}{RT} \Rightarrow M = \frac{dRT}{P}$$

Substitute values:

M = 1.43 g/L x 0.0821 L atm / mol K x 273 K / 1.00 atm = 31.9 g ➔ O_2,molecular oxygen

Key talking points for SI leaders

- **Common student errors:**
  - Mixing up formulas (e.g., using $PV = nRT$instead of rearranging to use density).
  - Forgetting that density in this context must be in g/L, not g/mL or kg/m³.
  - Leaving Runits inconsistent with $P$and T.
- **Conceptual emphasis:**
  - Reinforces that gas density depends on molar mass and the conditions of $P$and T.
  - At STP, lighter gases (like H_2, He) have smaller molar masses → lower densities; heavier gases → higher densities.
  - The derived M = 32 g/mol confirms O_2, showing how density can identify gases.
- **Real-world context:**
  - Used in gas identification and leak detection (e.g., industrial gas analysis).
  - Demonstrates how macroscopic properties like density connect directly to molecular-scale quantities through the ideal gas law.

Faculty Interview Protocol

Project: Cultivating Resilience in General Chemistry I
Purpose: Capture faculty experiences that model resilience, growth mindset, and the development of confidence in chemistry for General Chemistry I students.

Opening (Warm-Up and Context)

Begin with brief context-setting:

- Please introduce yourself (name, role, area of chemistry, years teaching General Chemistry I).
- What do you enjoy most about teaching Chemistry?

Section I: Personal Experience as a Learner

1. When you were a student learning chemistry, did you ever struggle?
    - Can you describe a specific challenge you faced?
    - How did you feel at the time?
    - What helped you push through it?

2. Did your confidence in chemistry change over time as you progressed through your studies?
    - What experiences helped build that confidence?

3. Looking back, how did your beliefs about your ability to learn chemistry evolve?
    - Was there a moment when you realized that ability in chemistry can grow with effort and effective strategies?

Section II: Influence on Teaching Practice

- 4. How have your own experiences with struggle influenced the way you teach chemistry today?
- 5. What specific strategies do you use in your classes to reinforce the idea that students' abilities in chemistry can develop over time?

Section III: Observations of Student Mindset and Growth

6. Have you observed differences between students who approach chemistry challenges with a growth mindset versus those who believe ability is fixed? What differences stand out to you?

7. Can you share an example of a student who transformed their performance or confidence over time? What do you think contributed to that change?

Section IV: Closing Message to Students

8. What message would you like new General Chemistry students to hear when they begin this course?

Interviewer Guidance (Optional Prompts)

Use as needed to deepen responses:

- Can you give a specific example?
- What did you say to that student?
- How did the student respond?
- How did that experience shape your view of learning?
- How do you communicate that message early in the semester?

Student Interview Protocol

Project: Cultivating Resilience in General Chemistry I
Purpose: Capture authentic student experiences that model growth mindset, resilience, and chemistry self-efficacy for General Chemistry I students.

Opening (Warm-Up – Build Comfort)

1. To begin, can you briefly introduce yourself?
   - What is your major?
   - When did you take General Chemistry I?
   - What were your initial expectations about the course?

Section I: Initial Confidence and Expectations

2. At the beginning of your chemistry coursework, how confident did you feel about succeeding?
   - What influenced that level of confidence?
   - Did that confidence change over time?

Section II: Experiencing Challenge

3. Can you describe a challenging moment you faced while learning chemistry?
   - What made it difficult?
   - How did you feel at that moment?
   - What thoughts were going through your mind?

4. Were there topics that initially felt impossible to understand?
   - What made them feel that way?
   - What helped you eventually make progress?

5. Did you ever feel overwhelmed or anxious in lecture or lab?
   - What contributed to that feeling?
   - How did you work through it?

Section III: Mindset Shift and Strategy Development

6. After experiencing a setback, did your beliefs about your ability to learn chemistry change?
   - If so, how?
   - What helped you shift your perspective?

7. Can you share a specific example where effort, strategy changes, or persistence led to improvement?
   - What did you do differently?
   - When did you realize you were improving?

8. General chemistry often requires independent practice outside of class. How did you learn to take ownership of your studying?
- What habits or systems worked best for you?

Section IV: Role of Others and Support Systems

9. What role did support from peers, instructors, SI leaders, or mentors play in your progress?
- Did seeing others struggle or succeed change how you viewed your own challenges?
- How important was having a support system?

Section V: Reflection and Advice

10. Looking back, how would you describe your growth in chemistry?
- What changed most: your skills, your mindset, your confidence?

11. What specific habits or study strategies helped you the most?

12. If you could give advice to new General Chemistry students about handling challenges, what would you tell them?
- What should they remember when things feel difficult?

Optional Follow-Up Prompts for Interviewer

Use as needed to deepen responses:
- Can you give a specific example?
- What did that experience teach you?
- How did that affect your confidence?
- What would you say to a student feeling that way right now?
- What did you do differently after that experience?

45 minutes

Supplemental Instruction Sessions - Week of Feb 16th to Feb 20th

Thank you for participating in this week's SI session!
This short survey helps us confirm your attendance, understand how you engaged with the problem set, and learn about your experiences in chemistry.
Your responses will:

- Verify your participation for extra credit,
- Provide feedback to improve future SI sessions, and
- Help us better understand how students think and learn chemistry.

There are no right or wrong answers. Please respond honestly.

Your individual responses will remain confidential and will only be used for instructional purposes to improve our services.

* Required

* This form will record your name, please fill your name.

1. Which SI session did you attend this week? *

- ◯ Monday
- ◯ Tuesday
- ◯ Wednesday
- ◯ Thursday
- ◯ Friday

2. Which chapter problems did you work on? *

- ◯ Chapter 1
- ◯ Chapter 2

3. Please read each statement below and then indicate how much you agree with each sentence from strongly agree to strongly disagree. There are no right or wrong answers. *

	strongly agree	agree	neutral	disagree	strongly disagree
You have a certain amount of intelligence in chemistry, and you really can't do much to change it.	◯	◯	◯	◯	◯
Your intelligence in chemistry is something about you that you can't change very much.	◯	◯	◯	◯	◯
No matter who you are, you can significantly change your intelligence level in chemistry.	◯	◯	◯	◯	◯
To be honest, you can't really change how intelligent you are in chemistry.	◯	◯	◯	◯	◯
You can always substantially change how intelligent you are in chemistry.	◯	◯	◯	◯	◯
You can learn new things, but you can't really change your basic intelligence in chemistry.	◯	◯	◯	◯	◯
This is a control question. Please select "Disagree" as your response to this question.	◯	◯	◯	◯	◯
No matter how much intelligence you have in chemistry, you can always change it quite a bit.	◯	◯	◯	◯	◯
You can change even your basic intelligence level in chemistry considerably.	◯	◯	◯	◯	◯

4. Please read each statement below and indicate how well you understand different areas of chemistry on a scale from very poorly to very well.

	very poorly	poorly	neutral	well	very well
To what extent can you explain chemical laws and theories?	◯	◯	◯	◯	◯
How well can you choose an appropriate formula to solve a chemistry problem?	◯	◯	◯	◯	◯
How well can you describe the properties of elements by using the periodic table?	◯	◯	◯	◯	◯
How well can you read the formulas of elements and compounds?	◯	◯	◯	◯	◯
How well can you interpret chemical equations?	◯	◯	◯	◯	◯
How well can you interpret graphs/charts related to chemistry?	◯	◯	◯	◯	◯

5. What's one strategy, concept, or insight from the SI session you're taking away that will help on exams? *