

Grade 11 Chemistry

Study Guide

Formulas · worked examples · practice problems with the answer key

Built to Ontario's SCH3U. Also covers most of
Alberta Chemistry 20 and parts of BC Chemistry 11.



6 modules · 26 formulas · 80 practice problems



The Exam Room · formula.expert · metric edition 1
October 2026 · Mathieu Leclerc, with Claude

About this study guide

Who it's for. Students working through Grade 11 Chemistry, and the teachers, tutors and parents printing handouts for them. Print it, copy it, staple it: it's free for classroom use, whole or by the page.

How to use it. Each module teaches its formulas in order (equation, picture, and what every symbol means), then closes with a practice set. Work the problems with a pencil; the answer key waits at the back so honest work gets an honest check. The two-page formula sheet up front is the exam-day companion, and every formula carries its number (Chemistry 11-1 onward) so your written work can cite exactly which one you used.

Do it live. Every problem in here has a living version in the Exam Room at formula.expert/exams/grade-11-chemistry: new numbers every attempt, hints that pop down when you want them, stars and streaks when you don't need them. When a printed problem stumps you, the matching lesson will coach you through it, step by step, free and with no sign-up. Other courses' study guides live at formula.expert/study-guides.

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The formula sheet

Module 1 · The Mole

Chemistry 11-1 Moles from Mass ($n = m/M$)

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

Converts a measured mass into an amount of substance by dividing by the molar mass.

Chemistry 11-2 Particles from Moles (Avogadro's Number)

$$N = n N_A$$

Converts an amount in moles into an actual particle count using the Avogadro constant $N_A \approx \dots$

Chemistry 11-3 Percent Composition of an Element

$$\%X = \frac{a M_X}{M_{\text{compound}}} \times 100\%$$

Gives the mass percent an element contributes to a compound from the formula subscript and molar...

Module 2 · Reactions by the Numbers

Chemistry 11-4 Percent Yield

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

Compares the mass actually isolated from a reaction to the maximum mass stoichiometry predicts.

Chemistry 11-5 Heat of Reaction

$$q = n\Delta H$$

Scales a reaction's molar enthalpy change by the amount reacted to give the total heat released or...

Module 3 · Solutions & Dilution

Chemistry 11-6 Molarity ($C = n/V$)

$$C = \frac{n}{V}$$

Defines molar concentration as moles of solute per volume of solution.

Chemistry 11-7 Mass Percent of a Solution

$$c = \frac{m_{\text{solute}}}{m_{\text{solution}}} \times 100\%$$

Expresses a solution's concentration as the solute's share of the total solution mass in percent.

Chemistry 11-8 Dilution Equation ($C_1V_1 = C_2V_2$)

$$C_1V_1 = C_2V_2$$

States that concentration times volume is conserved when a solution is diluted, since the moles of...

Module 4 · Gases

Chemistry 11-9 Ideal Gas Law

$$PV = nRT$$

Relates pressure, volume, amount, and temperature of a gas.

Chemistry 11-10 Boyle's Law

$$P_1V_1 = P_2V_2$$

At constant temperature, pressure times volume stays constant for a fixed amount of gas.

Chemistry 11-11 Charles's Law

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

At constant pressure, gas volume is directly proportional to absolute temperature.

Chemistry 11-12 Gay-Lussac's Law

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

At constant volume, gas pressure is directly proportional to absolute temperature.

Chemistry 11-13 Combined Gas Law

$$\frac{P_1V_1}{T_1} = \frac{P_2V_2}{T_2}$$

For a fixed amount of gas, pressure times volume over absolute temperature stays constant between...

Chemistry 11-14 Gas Volume at STP

$$V = n V_m$$

Converts between moles of an ideal gas and its volume at STP using the molar volume $V_m = 22.414 \dots$

Chemistry 11-15 Gas Density from Molar Mass

$$\rho = \frac{PM}{RT}$$

Gives an ideal gas's density from its molar mass, pressure, and absolute temperature using $R = \dots$

Module 5 · Water by the Numbers

Chemistry 11-16 Nutrient ppm from Fertiliser Weight

$$c = \frac{mf}{V}$$

The concentration of a single nutrient element delivered to a reservoir, from the weight of...

Chemistry 11-17 TDS Estimated from Conductivity ($\text{TDS} = k \times \text{EC}$)

$$\text{TDS} = k \times \text{EC}$$

Estimates total dissolved solids in mg/L from a conductivity reading in $\mu\text{S/cm}$ using an empirical...

Chemistry 11-18 Total Hardness as CaCO_3

$$\text{TH} = 2.497 \text{ Ca} + 4.118 \text{ Mg}$$

Converts a lab report's calcium and magnesium ion results into a single total hardness expressed as...

Chemistry 11-19 Grains per Gallon $\leftrightarrow$ ppm Hardness

$$H = 17.118 G$$

Converts water hardness between grains per US gallon and mg/L as CaCO_3 using the exact 17.118 mg/L...

Chemistry 11-20 Ion Concentration as CaCO_3 Equivalent

$$C_{\text{CaCO}_3} = C_{\text{ion}} \times \frac{50.04}{\text{EW}}$$

Restates any ion's concentration in the trade's common currency of CaCO_3 by scaling with the ratio...

Chemistry 11-21 Equivalent Weight from Molar Mass and Valence

$$\text{EW} = \frac{M}{z}$$

Gives the gram-equivalent weight of an ion or compound as its molar mass divided by the number of...

Chemistry 11-22 Chlorine Dose, Demand and Residual

$$D = C_{\text{demand}} + C_{\text{res}}$$

The fundamental chlorination balance: the dose applied equals the chlorine consumed by the water's...

Chemistry 11-23 Chemical Feed Rate ($\text{lb/day} = \text{mg/L} \times \text{MGD} \times 8.34$)

$$\dot{m} = C Q$$

Converts a target chemical dose and a plant flow into a mass feed rate, the classic pounds-per-day...

Module 6 · Acids, Bases & pH

Chemistry 11-24 pH from Hydrogen Ion Concentration

$$\text{pH} = -\log_{10} [\text{H}^+]$$

Expresses acidity as the negative base-10 logarithm of the hydrogen ion concentration in mol/L.

Chemistry 11-25 pH and pOH Relation

$$\text{pH} + \text{pOH} = 14$$

Links the acidity and basicity scales of any aqueous solution at 25 °C, where $\text{pK}_w = 14$.

Chemistry 11-26 Titration: Concentration of an Unknown

$$C_a = \frac{n C_b V_b}{V_a}$$

Finds an analyte's concentration from the titre volume, titrant concentration, and the reaction's...

MODULE 1

The Mole

Counting by weighing · Moles from mass · Mass and molar mass · Counting particles · Percent composition · The mole two-step · The Mole Inspection

1.1 Moles from Mass ($n = m/M$)

(Chemistry 11-1)
$$n = \frac{m}{M}$$

Where,

n = Amount of substance (mol)

m = Mass (kg)

M = Molar mass (g/mol)

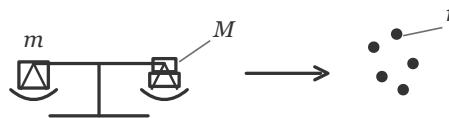


Figure 1.1: moles from mass ($n = m/M$).

The mole is chemistry's counting unit: a fixed number of particles large enough to weigh on a bench balance. Dividing a measured mass by the molar mass is the bridge between the two worlds: weigh out 36.0 g of water, divide by its molar mass of 18.02 g/mol, and you know you have 2.00 mol, about 1.2×10^{24} molecules. Every stoichiometry problem starts or ends with this conversion, because balanced equations speak in moles while balances speak in grams.

The word *mole* was coined by Wilhelm Ostwald in the 1890s, but the idea goes back to Avogadro's 1811 hypothesis that equal gas volumes hold equal numbers of molecules. Since the 2019 SI redefinition, the mole is defined by an exact count, $6.02214076 \times 10^{23}$ particles, so molar masses in g/mol are now measured quantities rather than definitions.

Worked example 1.1. Given $m = 116.88$ g, $M = 58.44$ g/mol; find n (amount of substance). *116.88 g NaCl ($M = 58.44$ g/mol) $\rightarrow$ exactly 2 mol.*

1.2 Particles from Moles (Avogadro's Number)

(Chemistry 11-2)
$$N = n N_A$$

Where,

N = Number of particles

n = Amount of substance (mol)

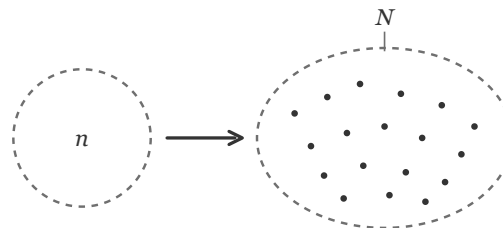


Figure 1.2: particles from moles (avogadro's number).

A mole is a count and nothing more mysterious: the chemist's dozen, scaled up to something useful. Multiply an amount in moles by the Avogadro constant and you have the literal number of particles in front of you. The only real question is why the constant is that particular size, and the answer is that it was chosen to make the bridge between two worlds land cleanly: N_A is the number of atoms

that makes a mole of carbon-12 weigh exactly 12 grams. That choice is what lets a mass in grams read off a balance be converted into a count of atoms, which is the single most useful trick in chemistry.

Scale is the thing worth feeling here. A 250 g glass of water is $250/18.015 = 13.9$ mol, so it holds $13.9 \times 6.022 \times 10^{23} = 8.4 \times 10^{24}$ molecules. Now count the other way: all the water on Earth, about 1.4×10^{21} litres, divided into 250 g glasses, comes to roughly 5.4×10^{21} glasses. **There are about fifteen hundred times more molecules in one glass of water than there are glasses of water in every ocean on the planet.** That ratio is why chemists never think about individual molecules and why statistical behaviour is so reliable at this scale.

The constant carries Avogadro's name but not his arithmetic: he proposed in 1811 that equal gas volumes hold equal numbers of particles, and never estimated the number. Jean Perrin did, from Brownian motion, and named it for Avogadro in 1909, work that took him the 1926 Nobel Prize. The 2019 SI redefinition then reversed the logic entirely. N_A is now *exact by decree* at $6.02214076 \times 10^{23}$ per mole, and the mole is defined as that many entities. Carbon-12's role is retired: a mole of it now weighs 12 grams only to within experimental uncertainty, rather than by definition. So 2.00 mol contains exactly $1.204428152 \times 10^{24}$ particles, with no uncertainty in the constant at all, though your measured 2.00 mol still has its own.

The error that swallows this page whole is leaving "particles of what" unstated. A mole of O_2 is 6.022×10^{23} molecules but 1.204×10^{24} atoms. A mole of NaCl is 6.022×10^{23} formula units, which is

6.022×10^{23} sodium ions plus the same number of chloride ions: 1.204×10^{24} ions in total. A mole of $Al_2(SO_4)_3$ contains three moles of sulfate. Almost every wrong answer here is a correct calculation attached to the wrong noun, so write the noun down before you multiply: molecules, atoms, ions, or formula units.

Two smaller slips. Dividing a molar mass by N_A gives the mass of one particle in grams, water comes out at 2.99×10^{-23} g, and people routinely lose a factor of 1000 by mixing kilograms into that step, or confuse it with the mass in unified atomic mass units, which is just the molar mass number again with different units. And because N_A is exact, it never limits your significant figures; only the amount you measured does.

Worked example 1.2. Given $n = 2$ mol; find N (number of particles). $2 \text{ mol} \rightarrow 1.204428152e24 \text{ particles}$.

1.3 Percent Composition of an Element

(Chemistry 11-3)
$$\%X = \frac{a M_X}{M_{\text{compound}}} \times 100\%$$

Where,

$\%X$ = Mass percent of element X (%)

a = Atoms of X per formula unit

M_X = Molar mass of element X (g/mol)

M_{compound} = Molar mass of compound (g/mol)

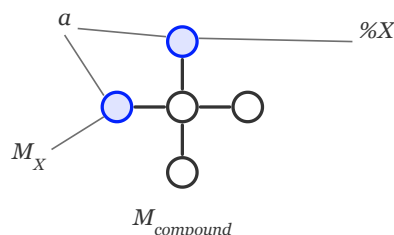


Figure 1.3: percent composition of an element.

A chemical formula is a recipe by count, and this equation translates it into a recipe by mass. Molar mass is additive (the compound weighs exactly what its atoms weigh, summed) so an element's share of the total is its atoms' contribution divided by the whole. In water, hydrogen brings $2 \times 1.008 = 2.016$ g/mol out of 18.02, which is 11.2% by mass; oxygen carries the other 88.8%. Two atoms out of three, and barely a ninth of the weight.

A case with real stakes. Ammonium sulfate, $(NH_4)_2SO_4$, has a molar mass of 132.14 g/mol. Nitrogen appears twice, so $\%N = (2 \times 14.007)/132.14 \times 100 = 21.2\%$. That is not a coincidence of the

classroom: it is why the fertilizer is sold as **21-0-0**, and why a 25 kg bag delivers 5.3 kg of actual nitrogen to a field.

Run the arithmetic backwards and it becomes analysis rather than description. Burn an unknown compound, weigh the CO_2 and H_2O produced, convert those to masses of carbon and hydrogen, take oxygen by difference, then divide each element's mass percent by its atomic mass and normalise to the smallest result. What comes out is the empirical formula. This is combustion analysis, and it was the central technique of organic chemistry from Liebig's refinement of it in the 1830s until spectroscopy arrived. Behind all of it sits Proust's

law of definite proportions: a given compound always holds the same elements in the same mass ratio, which is exactly the claim this equation encodes.

The subscript a is the number of atoms of that element per formula unit, and parentheses are where people lose it. In

$(\text{NH}_4)_2\text{SO}_4$ the hydrogen count is 8, not 4: the subscript outside the bracket multiplies everything inside. Hydrates catch people the same way: copper(II) sulfate pentahydrate, $\text{CuSO}_4 \cdot 5\text{H}_2\text{O}$, has a molar mass of 249.7 g/mol including its water, so the copper is 25.4%, not the 39.8% you would get from the anhydrous salt. If the bottle says pentahydrate, the water is part of what you weighed.

Two further traps. Percent composition can only ever give you an *empirical* formula, never a molecular one: formaldehyde CH_2O , acetic acid $\text{C}_2\text{H}_4\text{O}_2$ and glucose $\text{C}_6\text{H}_{12}\text{O}_6$ all analyse to identical percentages, and separating them needs an independent molar mass. And on fertilizer bags, only the first number is what it appears to be. The N in N-P-K is genuinely percent nitrogen, but P is reported as P_2O_5 equivalent and K as K_2O : a nineteenth-century convention that never died. A bag marked 0-46-0 is 46% P_2O_5 , which works out to about 20% actual phosphorus. Reading it as elemental phosphorus overstates the dose by well over a factor of two.

Worked example 1.3. Given $a = 2$, $M_X = 1.008 \text{ g/mol}$, $M_{\text{compound}} = 18.015 \text{ g/mol}$; find %X (mass percent of element x).
Hydrogen in water: $2 \times 1.008 / 18.015 \rightarrow 11.1907\%$ by mass.

Module 1 practice

1. Moles from mass: A student weighs out 175.5 g of sodium chloride (NaCl, $M = 58.5 \text{ g/mol}$) into a clean beaker. *Calculate the amount of substance in the sample.*

2. Moles from mass: A student weighs out 88.0 g of carbon dioxide (CO_2 , $M = 44.0 \text{ g/mol}$) into a clean beaker. *Calculate the amount of substance in the sample.*

3. Mass and molar mass: A synthesis procedure calls for 5.00 mol of methane (CH_4 , $M = 16.0 \text{ g/mol}$) to be weighed out on the analytical balance. *Determine the mass the balance should read.*

4. Mass and molar mass: An unlabelled bottle in the stockroom holds a white solid. The technician's note says only that the bottle came from a set of four reagents: H_2O (ice), CO_2 (dry ice), NaCl and $\text{C}_6\text{H}_{12}\text{O}_6$. A 292.5 g portion of the solid is analysed and found to contain exactly 5.00 mol. *Identify the compound in the bottle.*

5. Counting particles: A sealed flask of neon gas holds 4 mol of the substance. *Calculate the number of atoms present.*

6. Counting particles: A mass spectrometer reports 1.204×10^{24} molecules in a captured sample of a pure gas. *Determine the amount of substance the sample contains.*

7. Percent composition: A lab manual asks for the mass percent of carbon in glucose, $\text{C}_6\text{H}_{12}\text{O}_6$. The data table gives $M(\text{C}) = 12.0 \text{ g/mol}$ and $M(\text{C}_6\text{H}_{12}\text{O}_6) = 180.0 \text{ g/mol}$. *Calculate the mass percent of carbon in $\text{C}_6\text{H}_{12}\text{O}_6$.*

8. Percent composition: A lab manual asks for the mass percent of oxygen in carbon dioxide, CO_2 . The data table gives $M(\text{O}) = 16.0 \text{ g/mol}$ and $M(\text{CO}_2) = 44.0 \text{ g/mol}$. *Calculate the mass percent of oxygen in CO_2 .*

9. The mole two-step: An air-quality instrument captures a sample of sodium chloride (NaCl, $M = 58.5 \text{ g/mol}$) and counts 1.204×10^{24} molecules in it. The report has to state the sample's mass. *Determine the mass of the captured sample.*

10. The mole two-step: An air-quality instrument captures a sample of sodium carbonate (Na_2CO_3 , $M = 106.0 \text{ g/mol}$) and counts 3.010×10^{23} molecules in it. The report has to state the sample's mass. *Determine the mass of the captured sample.*

11. The Mole Inspection: Inspection day. On the bench sits one weighed sample of distilled water, H_2O , at 72.0 g. The atomic masses posted on the wall are $\text{H} = 1.0$, $\text{O} = 16.0$, and Avogadro's number is $6.02 \times 10^{23} / \text{mol}$. Work each line — every answer feeds the next. *Determine the molar mass, the amount, the particle count and the mass percent of hydrogen, one line at a time.*

12. The Mole Inspection: The last jar on the inspection bench is labelled in a previous technician's handwriting: " H_2O — contains 5.00 mol". You put the jar's contents on the balance and read 72.0 g. The wall still posts $\text{H} = 1.0$, $\text{O} = 16.0$. *Decide whether the label can stay on the jar.*

MODULE 2

Reactions by the Numbers

Paper versus flask · Percent yield · Working back from yield · Heat of reaction · Moles to heat · Estimate, then compute · The Lab Report

2.1 Percent Yield

(Chemistry 11-4)
$$\% \text{ yield} = \frac{m_{\text{actual}}}{m_{\text{theoretical}}} \times 100\%$$

Where,

$\% \text{ yield}$ = Percent yield (%)

m_{actual} = Actual yield (mass obtained) (kg)

$m_{\text{theoretical}}$ = Theoretical yield (mass predicted) (kg)

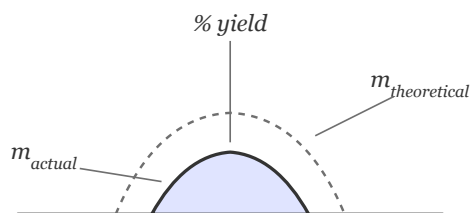


Figure 2.1: percent yield.

A balanced equation promises a certain mass of product; the flask rarely delivers all of it. Side reactions consume reactant, some product stays dissolved in the mother liquor, and a little is always lost on filter paper and glassware. Percent yield is the honest scorecard: the mass you actually isolated divided by the stoichiometric maximum, times 100. In a classic teaching lab, synthesizing aspirin from salicylic acid might predict 5.00 g of product; if 4.21 g of dry crystals come off the funnel, the yield is 84.2%.

The number matters far beyond the classroom. Process chemists judge manufacturing routes largely by yield, because in a multi-step synthesis losses multiply: five steps at 80% each deliver only 33% overall. A reported yield above 100% is a red flag, not a triumph: it usually means the product is still wet or carries impurities.

Worked example 2.1. Given $m_{\text{actual}} = 4.10 \text{ g}$, $m_{\text{theoretical}} = 5.00 \text{ g}$; find $\% \text{ yield}$ (percent yield). *4.10 g isolated of 5.00 g theoretical $\rightarrow$ 82% yield.*

2.2 Heat of Reaction

(Chemistry 11-5)
$$q = n\Delta H$$

Where,

q = Heat released or absorbed (J)

n = Amount of substance (mol)

ΔH = Molar enthalpy change (kJ/mol)

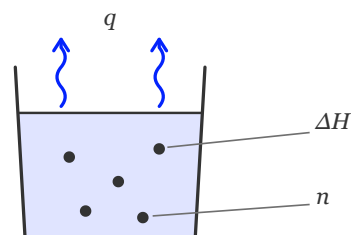


Figure 2.2: heat of reaction.

Enthalpy is an extensive quantity: run a reaction twice and you get twice the heat. That is the entire justification for $q = n\Delta H$, and it is why tabulating a single per-mole figure is enough to describe a reaction at any scale from a test tube to a boiler. The sign convention runs from the system's point of view: ΔH negative for an exothermic reaction, because the system's enthalpy falls as heat leaves it. A negative q is not an error message; it is the answer telling you the heat came out.

Something concrete. Heating 200 L of water from 10 °C to 60 °C takes $200 \times 4.186 \times 50 = 41\,860 \text{ kJ}$. Methane burns at $\Delta H = -890.3 \text{ kJ/mol}$, so the amount required is $n = 41\,860/890.3 = 47.0 \text{ mol}$: about **754 g of methane, or 1.05 m³ at STP**, before any consideration of how much of that heat actually reaches the water. Chain those three pages together and you have most of a combustion calculation.

The relation's real power is that enthalpy is a state function, which is Hess's law: the heat of a reaction depends only on where it starts and ends, not on the route. So you can add reactions like algebra, and you can build any ΔH you need from tabulated standard enthalpies of formation, products minus reactants, without ever running the reaction. One caution on what is being measured: ΔH is the heat at constant pressure, which is what an open vessel or a flowing burner delivers. A bomb calorimeter holds volume constant and measures ΔU instead, and the two differ by the work done pushing the atmosphere aside, $\Delta H = \Delta U + \Delta n_{\text{gas}}RT$.

The question this page most needs you to ask is "per mole of what?" ΔH belongs to the balanced equation as written, not to any one substance in it. $2\text{H}_2 + \text{O}_2 \rightarrow 2\text{H}_2\text{O}$ has $\Delta H = -571.6 \text{ kJ}$, but $\text{H}_2 + \frac{1}{2}\text{O}_2 \rightarrow \text{H}_2\text{O}$ has $\Delta H = -285.8 \text{ kJ}$. Same chemistry, same physical world, different bookkeeping, and a factor of two waiting for anyone who reads a table without reading the equation above it. Write the equation first, decide which species your n counts, and make the two agree.

Then the trap that costs more marks than any other in stoichiometry. **The limiting reagent is found by moles, never by mass.** Given 100 g of hydrogen and 100 g of oxygen, the masses are equal and the

amounts are not remotely: 49.6 mol of H_2 against 3.13 mol of O_2 , and the reaction needs two hydrogens per oxygen. Oxygen limits by a factor of eight, and 100 g of hydrogen looks generous only because hydrogen is light. Convert everything to moles, divide each by its coefficient in the balanced equation, and the smallest quotient is the limiter. The n that goes into $q = n\Delta H$ is then the extent of reaction that limiter permits, not the amount of whatever you happened to weigh out.

One last note for anyone comparing fuel figures. Combustion enthalpies come in two flavours depending on whether the product water is counted as liquid or as vapour, differing by 44 kJ per mole of water. That is the higher-heating-value and lower-heating-value distinction, and for methane it is about a 10% gap. Two sources can disagree by that much while both being correct.

Worked example 2.2. Given $n = 2 \text{ mol}$, $\Delta H = -890 \text{ kJ/mol}$; find q (heat released or absorbed). *2 mol CH₄ at $dH = -890 \text{ kJ/mol} \rightarrow q = -1780 \text{ kJ}$.*

Module 2 practice

13. Percent yield: A copper-recovery lab predicts 20 g of copper at best; the dried filter paper holds 12 g of it. *Determine the percent yield.*

14. Percent yield: On paper, an esterification can deliver 20 g of ester. The separated, dried ester comes to 12 g. *Calculate the percent yield of the reaction.*

15. Working back from yield: A run isolates 12 g of dry product, and the procedure's notes call that a 80% yield. *Determine the theoretical yield the stoichiometry must have promised.*

16. Working back from yield: A run isolates 36 g of dry product, and the procedure's notes call that a 90% yield. *Determine the theoretical yield the stoichiometry must have promised.*

17. Heat of reaction: A camping stove burns 3.5 mol of fuel; the combustion's enthalpy change is $\Delta H = -300 \text{ kJ/mol}$. *Calculate the heat released.*

18. Heat of reaction: A cold pack dissolves 2.5 mol of salt; the process absorbs heat with $\Delta H = +250 \text{ kJ/mol}$. *Calculate the heat absorbed.*

19. Moles to heat: A fuel pellet weighing 54 g (molar mass 18 g/mol) burns completely; the combustion's enthalpy change is $\Delta H = -250 \text{ kJ/mol}$. *Determine the heat released, moles first.*

20. Moles to heat: A chemical hand warmer must deliver 400 kJ over an evening. Its reaction releases 100 kJ per mole of reactant ($\Delta H = -100 \text{ kJ/mol}$), and the reactant's molar mass is 58.5 g/mol. *Determine the mass of reactant to load, moles first.*

21. Estimate, then compute: A distillation of banana ester is predicted to give 9.2 g at best; the flask's dried product weighs 2.3 g. *Determine the percent yield — estimate first.*

22. Estimate, then compute: An industrial batch reacts 3 mol of feedstock; the process's enthalpy change is $\Delta H = -2000 \text{ kJ/mol}$. *Determine the heat released — estimate the power of ten first.*

23. The Lab Report: Last page of the term. A fuel sample weighing 110 g (molar mass 44 g/mol) is burned in a calorimeter; the data table lists $\Delta H = -200 \text{ kJ/mol}$, and the calorimeter logs 425 kJ actually captured. Work each line — every answer feeds the next. *Determine what percent of the predicted heat the calorimeter captured, one line at a time.*

24. The Lab Report: Bonus mark, worked backwards: a classmate's report shows 75% of the predicted heat captured, with 150 kJ observed. The fuel's table values: $\Delta H = -200$ kJ/mol, molar mass 44 g/mol. Determine the mass of fuel the classmate weighed out.

MODULE 3

Solutions & Dilution

Reading the label · Molarity · Molarity, every way · Mass percent · Dilution · Dilution, every way · Stockroom chains · The Stockroom

3.1 Molarity ($C = n/V$)

(Chemistry 11-6)
$$C = \frac{n}{V}$$

Where,

C = Molar concentration (M)

n = Amount of solute (mol)

V = Volume of solution (L)

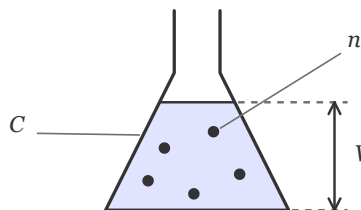


Figure 3.1: molarity ($c = n/v$).

Molarity answers the practical question "how much stuff is in this bottle?" by counting moles of solute per liter of solution. Dissolve 58.44 g of table salt, exactly one mole of NaCl, in water and top up to the 1.00 L mark of a volumetric flask, and you have a 1.00 M solution. Note the fine print: it is per liter of *solution*, not per liter of water added, which is why chemists fill to a calibrated mark instead of adding a measured liter of solvent.

Molarity is the workhorse concentration unit because reactions are mole-to-mole affairs: multiplying C by a dispensed volume immediately gives the moles delivered, which is exactly what a titration calculation needs. Physiological saline is about 0.154 M NaCl, ocean water roughly 0.5 M, and concentrated hydrochloric acid around 12 M: a span that dilution calculations cross daily in every lab.

Worked example 3.1. Given $n = 2$ mol, $V = 4$ L; find C (molar concentration). $2 \text{ mol in } 4 \text{ L} \rightarrow 0.5 \text{ mol/L}$.

3.2 Mass Percent of a Solution

(Chemistry 11-7)
$$c = \frac{m_{\text{solute}}}{m_{\text{solution}}} \times 100\%$$

Where,

c = Mass percent (%)

m_{solute} = Mass of solute (kg)

m_{solution} = Mass of solution (kg)

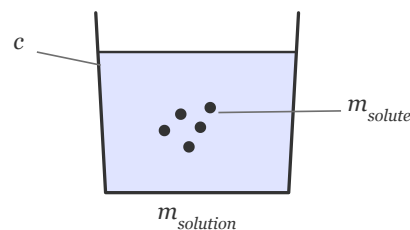


Figure 3.2: mass percent of a solution.

Mass percent is the concentration measure that needs nothing but a balance: the solute's mass divided by the mass of the *whole solution*, solute plus solvent, times 100. Dissolve 25 g of salt in 225 g of water

and the solution is $25/250 \times 100 = 10\%$ salt by mass. The classic mistake is dividing by the solvent mass alone, which would wrongly give 11.1%.

Because it is temperature-proof and instrument-free, mass percent dominates industrial and household labels: household vinegar is about 5% acetic acid, physiological saline 0.9% NaCl, seawater

roughly 3.5% dissolved salts, and concentrated sulfuric acid ships at 98%. Converting to molarity requires the solution's density, which is why bottle labels for concentrated acids list both figures side by side.

Worked example 3.2. Given $m_{\text{solute}} = 25 \text{ g}$, $m_{\text{solution}} = 250 \text{ g}$; find c (mass percent). $25 \text{ g salt in } 250 \text{ g solution} \rightarrow c = 10\%$.

3.3 Dilution Equation ($C_1V_1 = C_2V_2$)

(Chemistry 11-8) $C_1V_1 = C_2V_2$

Where,

C_1 = Initial concentration (M)

V_1 = Initial volume (L)

C_2 = Final concentration (M)

V_2 = Final volume (L)

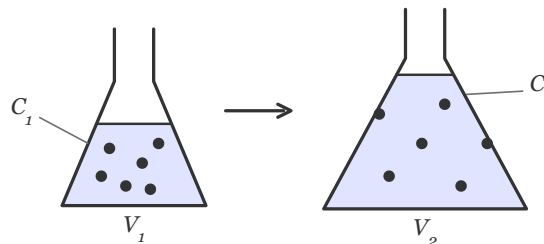


Figure 3.3: dilution equation ($c_1v_1 = c_2v_2$).

Adding solvent to a solution spreads the same solute through a larger volume: the moles do not change, only their crowding. Since moles equal concentration times volume, C_1V_1 must equal C_2V_2 before and after any dilution. To prepare 250 mL of 1.0 M hydrochloric acid from a 12.1 M concentrated stock, solve for V_1 : $(1.0 \times 250)/12.1 \approx 20.7 \text{ mL}$ of stock, made up to the 250 mL mark with water. (And always add acid to water, never the reverse.)

The law is not limited to molarity: any concentration measure proportional to moles per volume works, as long as both sides use the same one. Water-treatment operators lean on it daily when dosing inhibitor or biocide from concentrated drums into recirculating loops, and biologists use the identical arithmetic for serial dilutions, where each step divides concentration by a fixed factor.

Worked example 3.3. Given $C_1 = 6.00 \text{ M}$, $V_1 = 50.0 \text{ mL}$, $V_2 = 300.0 \text{ mL}$; find C_2 (final concentration). $50 \text{ mL of } 6.0 \text{ M diluted to } 300 \text{ mL} \rightarrow 1.0 \text{ mol/L}$.

Module 3 practice

25. Molarity: A technician dissolves 3 mol of sodium chloride in enough water to make 2 L of solution. *Calculate the molar concentration of the solution.*

26. Molarity: A student dissolves 0.2 mol of glucose in enough water to make 500 mL of solution. *Calculate the molar concentration, in mol/L.*

27. Molarity, every way: A procedure calls for 600 mL of a 2.5 mol/L silver nitrate solution. *Determine the amount of solute dissolved in that volume.*

28. Molarity, every way: A procedure calls for 400 mL of a 0.5 mol/L silver nitrate solution. *Determine the amount of solute dissolved in that volume.*

29. Mass percent: A student stirs 45 g of table salt into 255 g of water until it fully dissolves. *Calculate the mass percent of salt in the solution.*

30. Mass percent: A student stirs 20 g of table salt into 380 g of water until it fully dissolves. *Calculate the mass percent of salt in the solution.*

31. Dilution: A technician measures 250 mL of a 2 mol/L stock solution into a flask and adds water up to the 1000 mL mark. *Calculate the concentration of the diluted solution.*

32. Dilution: A technician measures 50 mL of a 5 mol/L stock solution into a flask and adds water up to the 1000 mL mark. *Calculate the concentration of the diluted solution.*

33. Dilution, every way: The procedure calls for 500 mL of 0.9 mol/L acid. The stockroom shelf offers a 6 mol/L stock bottle. *Determine the volume of stock to measure out.*

34. Dilution, every way: A label has worn off a stock bottle. To identify it, a technician dilutes 80 mL of the stock to 400 mL total and measures the result at 0.7 mol/L. *Determine the concentration of the mystery stock.*

35. Stockroom chains: Monday morning in the stockroom. The prep list asks for a fresh bottle of sodium hydroxide solution, and the recipe card reads: weigh 80 g of NaOH pellets (molar mass 40 g/mol), dissolve in distilled water, and top up to exactly 500 mL of solution. The finished bottle needs a concentration for its label. *Determine the molar concentration to write on the label.*

36. Stockroom chains: The afternoon job order: the biology lab wants dilute hydrochloric acid for tomorrow, and all the stockroom holds is a 6 mol/L stock bottle. The card says: measure 100 mL of stock into a volumetric flask, add distilled water to the 800 mL line, cap and invert. The new bottle needs its own label. *Determine the concentration of the diluted acid.*

37. The Stockroom: Last job before the bell. The order: make 2 L of 0.5 mol/L sodium hydroxide for tomorrow's titration lab. On the bench: NaOH pellets ($M = 40 \text{ g/mol}$ — a round number, chosen kindly), a balance, a 0.5 L volumetric flask for the stock, and distilled water. The card says: weigh 40 g, make 0.5 L of stock, then dilute. Work each line — every answer feeds the next. *Determine how much stock and how much water build the lab's bottle, one line at a time.*

38. The Stockroom: Bonus mark, worked backwards: a bottle on the shelf reads "0.5 mol/L — made by diluting 0.5 L of stock up to 2 L". The stock bottle it came from has lost its own label. *Determine the concentration of the unlabelled stock.*

MODULE 4

Gases

The gas variables · Boyle's law · Hot gases · The combined gas law · The ideal gas law · $PV = nRT$, every way · Molar volume at STP · Weighing a gas · The Pressure Test

4.1 Ideal Gas Law

(Chemistry 11-9) $PV = nRT$

Where,

P = Pressure (kPa)

V = Volume (L)

n = Amount (mol)

T = Temperature ($^{\circ}\text{C}$)

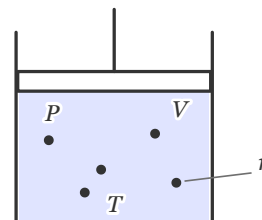


Figure 4.1: ideal gas law.

$PV = nRT$ says that for a gas, four quantities are not independent: fix any three and the fourth is decided. Squeeze it and the pressure rises; warm it and it pushes harder or swells; add more of it and both go up. What makes the equation remarkable is not that those things are true, anyone with a bicycle pump knows them, but that **one constant serves every gas**. Helium, nitrogen and steam all obey it with the same $R = 8.314 \text{ J/(mol}\cdot\text{K)}$, which is a strong hint that pressure has nothing to do with what the molecules are and everything to do with how many there are and how fast they are moving.

A worked case in units you would actually read off a gauge. A 20 L cylinder sits at 150 kPa absolute on a 20 $^{\circ}\text{C}$ morning. Rearranged for amount, $n = PV/RT = (150\,000 \times 0.020)/(8.314 \times 293.15) = 3000/2437 \approx 1.23 \text{ mol}$. Note what had to happen before the arithmetic: pascals not kilopascals, cubic metres not

litres, and kelvin not Celsius. This page converts your entries for you, but the discipline is worth keeping in your head, because a scrap of paper will not.

The law arrived in pieces. Boyle established PV constant at fixed temperature in 1662; Charles and Gay-Lussac tied volume and pressure to temperature around 1800; Avogadro proposed in 1811 that equal volumes of gases hold equal numbers of particles. Émile Clapeyron folded them into a single expression in 1834. Kinetic theory later derived the whole thing from mechanics: treat molecules as point masses that bounce elastically and never attract one another, average over their collisions with the walls, and $PV = nRT$ falls out, with RT revealed as a measure of the average kinetic energy per mole.

Those two assumptions are also the fine print, and here a common textbook line deserves correcting. **The ideal gas law is a limit, not a fact about gases.** Real molecules do occupy volume and do attract each other, so the equation is exact only as pressure approaches zero and the gas gets out of its own way. Near condensation it fails plainly: at 100 atm, or anywhere close to the boiling point, the error runs to tens of percent and you want van der Waals or a compressibility factor. Under ordinary room conditions the error is well under 1%, which is why the approximation earns its keep.

Two errors account for most wrong answers on this page, and both are unit errors rather than physics errors. The first is feeding in Celsius. Doubling a gas from 20 °C to 40 °C does not double anything: in kelvin that is 293 to 313, a rise of 7%, and a calculation that used 20 and 40 would be wrong by a factor of nearly two. The second is feeding in a gauge pressure. A tire gauge reading 220 kPa means 321 kPa absolute; P here is absolute pressure, measured from vacuum, because the equation counts molecular impacts and vacuum is where there are none. A third, quieter trap: the familiar 22.4 L per mole belongs to 0 °C and 1 atm. IUPAC redefined standard pressure to 100 kPa in 1982, and at that pressure the molar volume is 22.71 L. Both numbers circulate, and quoting one against the other's conditions is a 1.3% error hiding inside a memorised constant.

Worked example 4.1. Given $P = 1$ atm, $n = 1$ mol, $T = 273.15$ K; find V (volume). *1 mol at 0 °C and 1 atm → 22.414 L (molar volume at STP).*

4.2 Boyle's Law

(Chemistry 11-10) $P_1 V_1 = P_2 V_2$

Where,

P_1 = Initial pressure (kPa)

V_1 = Initial volume (L)

P_2 = Final pressure (kPa)

V_2 = Final volume (L)

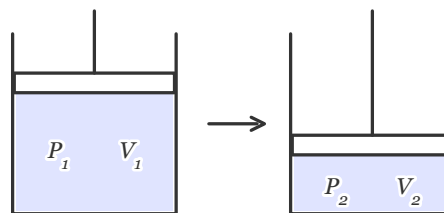


Figure 4.2: boyle's law.

Robert Boyle published his gas law in 1662, making it one of the oldest quantitative laws in physics: at constant temperature, the pressure and volume of a trapped gas are inversely proportional, so their product never changes. Cap a syringe and squeeze: compress 60 mL of air at 100 kPa down to 20 mL and the pressure climbs to 300 kPa. Molecularly, shrinking the space raises how often molecules hammer the walls, and pressure rises in exact proportion.

The law matters wherever gas gets squeezed. Scuba divers learn it first: air breathed at depth expands as they ascend, which is why the cardinal rule is never to hold your breath on the way up. Boyle's law assumes the temperature and the amount of gas stay fixed; change either and you need Charles's law or the combined gas law instead.

Worked example 4.2. Given $P_1 = 1$ atm, $V_1 = 2$ L, $P_2 = 4$ atm; find V_2 (final volume). *2 L at 1 atm → 4 atm gives 0.5 L.*

4.3 Charles's Law

(Chemistry 11-11)
$$\frac{V_1}{T_1} = \frac{V_2}{T_2}$$

Where,

V_1 = Initial volume (L)

T_1 = Initial absolute temperature (°C)

V_2 = Final volume (L)

T_2 = Final absolute temperature (°C)

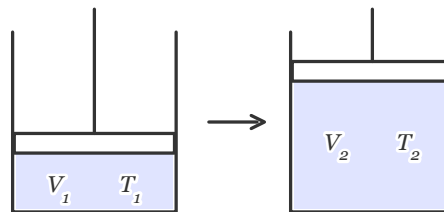


Figure 4.3: charles's law.

Hold the pressure on a gas constant and its volume follows its absolute temperature in strict proportion: warm it by 10% and it swells by 10%. The reason is worth having rather than memorising. Temperature is a measure of how fast the molecules are moving; pressure is how hard their impacts push on each square metre of wall. If the gas is to keep pushing with the same force while its molecules move faster, the only thing it can do is spread out, so that each patch of wall is struck less often. Volume rises exactly as fast as temperature to keep that balance, and $V_1/T_1 = V_2/T_2$ is that sentence in symbols.

A 2.5 L balloon leaves a 22 °C room and goes into a –18 °C freezer. In kelvin those are 295.15 and 255.15, so $V_2 = 2.5 \times 255.15/295.15 = 2.16$ L: it loses about 14% of its volume and visibly puckers. Bring it back out and it recovers. Nothing left the balloon; the same molecules simply stopped needing as much room.

Jacques Charles found this with hydrogen balloons around 1787 and never published it. Joseph Louis Gay-Lussac did the careful work and published in 1802, and generously named the result after Charles. The interesting part is what came out of extending the straight line. Plot volume against Celsius temperature and you get a line that, extrapolated backwards, hits zero volume at about –273 °C. No one in 1802 could reach anywhere near that temperature, yet the graph

pointed straight at it. That extrapolation is how absolute zero was first located, and it is why William Thomson could propose an absolute scale in 1848 by simply moving the origin to where the gases were pointing. The modern value, –273.15 °C, is the zero of the kelvin scale.

Which is exactly why a ratio in Celsius is meaningless here. Going from 20 °C to 40 °C does not double the volume; in kelvin that is 293.15 to 313.15, a rise of under 7%. Worse, a Celsius ratio breaks outright when the temperature crosses zero: 0 °C in the denominator gives infinity, and a negative Celsius temperature gives a negative volume. This page converts your °C or °F entries to kelvin before it does anything, but the trap is worth recognising when you meet the equation off-screen.

Two smaller conditions do real work. The pressure must actually be constant: a gas sealed in a rigid tank obeys Gay-Lussac's law instead, where pressure rises and volume does not move at all. And the gas must stay a gas. Cool steam through 100 °C and the relation does not merely become inaccurate, it stops applying: the vapour condenses and the volume collapses by a factor of about 1600. A balloon is also only approximately constant-pressure, since the stretched rubber adds a little tension of its own, which is why the real shrinkage runs slightly under what the arithmetic predicts.

Worked example 4.3. Given $V_1 = 2$ L, $T_1 = 300$ K, $T_2 = 600$ K; find V_2 (final volume). *2 L at 300 K heated to 600 K → 4 L.*

4.4 Gay-Lussac's Law

(Chemistry 11-12)
$$\frac{P_1}{T_1} = \frac{P_2}{T_2}$$

Where,

P_1 = Initial pressure (kPa)

T_1 = Initial absolute temperature (°C)

P_2 = Final pressure (kPa)

T_2 = Final absolute temperature (°C)

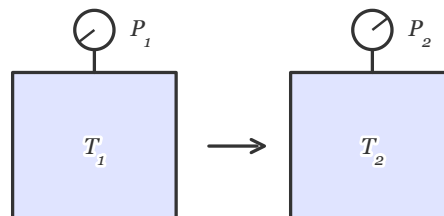


Figure 4.4: gay-lussac's law.

Seal a gas in a container that cannot change size and there is only one thing left for it to do when you heat it: push harder. Pressure then tracks absolute temperature in direct proportion, $P_1/T_1 = P_2/T_2$. The mechanism is the same one behind Charles's law with the outcome swapped. Faster molecules strike the walls both more often and with more momentum each time, and since the walls will not move aside, all of that arrives as pressure.

Here is the case everyone actually meets, worked carefully, because the careless version is the standard error. A tire is set to 220 kPa gauge on a 5 °C morning and warms to 45 °C after an hour on the highway. Convert to absolute pressure first: $220 + 101 = 321 \text{ kPa absolute}$. Then $P_2 = 321 \times 318.15/278.15 = 367 \text{ kPa absolute}$, which is 266 kPa on the gauge: a rise of 46 kPa, about 6.6 psi. Run the same calculation on the gauge reading alone and you get 252 kPa, understating the rise by a third. This is why tire pressures are specified cold, and why topping up a hot tire leaves it soft in the morning.

The relation is usually credited to Gay-Lussac's 1802 paper, though Guillaume Amontons had it a century earlier: around 1702 he built an air thermometer that worked on precisely this principle, and noticed that the pressure line extrapolated toward a temperature below which it could not go. Some texts call it Amontons's law for

that reason. Combine it with Boyle's law and Charles's law and you have the combined gas law; add Avogadro and you have $PV = nRT$, of which this is the constant-volume slice.

Absolute pressure is the trap here, more than absolute

temperature. The ratio P_1/P_2 is only meaningful when both pressures are measured from vacuum, because the equation is counting molecular impacts and a gauge has quietly subtracted an atmosphere from the count. Temperature has the same requirement for the same reason (kelvin, not Celsius) and this page converts your entries on both fronts. But when you meet the equation on paper, ask twice whether the pressure in your hand is gauge or absolute. It usually is gauge; almost every instrument in a mechanical room reads that way.

The other honest limit is that *constant volume* is an idealisation. A tire is not rigid: it grows a little as it warms, which relieves some of the pressure rise, so the measured increase runs slightly below the calculation. A steel cylinder is much closer to the ideal, which is what makes this law genuinely dangerous rather than merely academic. An aerosol can left on a dashboard, a propane cylinder in a closed vehicle, or a sealed pressure vessel in a fire all follow this line with nothing to relieve them, and the pressure keeps climbing until something gives. Relief valves exist because the equation has no upper bound.

Worked example 4.4. Given $P_1 = 3 \text{ atm}$, $T_1 = 300 \text{ K}$, $T_2 = 400 \text{ K}$; find P_2 (final pressure). *3 atm at 300 K heated to 400 K $\rightarrow$ 4 atm (405.3 kPa).*

4.5 Combined Gas Law

(Chemistry 11-13)
$$\frac{P_1 V_1}{T_1} = \frac{P_2 V_2}{T_2}$$

Where,

P_1 = Initial pressure (kPa)

V_1 = Initial volume (L)

T_1 = Initial absolute temperature (°C)

P_2 = Final pressure (kPa)

V_2 = Final volume (L)

T_2 = Final absolute temperature (°C)

The combined gas law merges Boyle's, Charles's, and Gay-Lussac's laws into a single statement: for a fixed amount of gas, PV/T is constant. A weather balloon shows all three variables moving at once. Launched with 2.0 m³ of helium at 101 kPa and 288 K, it rises to where the pressure is 30 kPa and the temperature 228 K; its new volume is $V_2 = P_1 V_1 T_2 \div (P_2 T_1) = 2.0 \times 101 \times 228 \div (30 \times 288) \approx 5.3 \text{ m}^3$: more than double.

Worked example 4.5. Given $P_1 = 1 \text{ atm}$, $V_1 = 1 \text{ L}$, $T_1 = 273.15 \text{ K}$, $P_2 = 0.5 \text{ atm}$, $T_2 = 546.3 \text{ K}$; find V_2 (final volume). *1 L at 1 atm, 273.15 K $\rightarrow$ 0.5 atm, 546.3 K gives 4 L.*

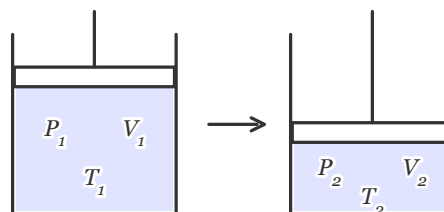


Figure 4.5: combined gas law.

Hold any one variable constant and the named laws drop out: fix T for Boyle's law, fix P for Charles's, fix V for Gay-Lussac's. Temperatures must be absolute (the ratio of 20 °C to 40 °C is not 1:2 but 293:313), and Celsius or Fahrenheit inputs convert to kelvin automatically. Add Avogadro's insight about the amount of gas n and this law becomes the full ideal gas law, $PV = nRT$.

4.6 Gas Volume at STP

(Chemistry 11-14) $V = n V_m$

Where,

V = Gas volume at STP (L)

n = Amount of gas (mol)

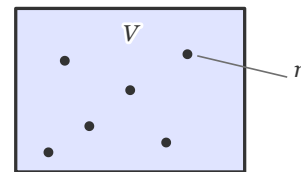


Figure 4.6: gas volume at stp.

Fix the temperature and pressure and a mole of *any* ideal gas occupies the same volume, so converting between moles and litres needs one constant and no information about the substance. This is Avogadro's principle, and it deserves a moment of surprise before it becomes routine. A mole of hydrogen weighs 2 g and a mole of sulfur hexafluoride weighs 146 g, seventy-three times more, yet at the same conditions they fill the same flask. The reason is that pressure comes from the number of impacts and their momentum, and at a common temperature the heavier molecules move proportionally slower. The mass cancels out of everything the container can feel.

This page uses classic STP, 0 °C and 1 atm, where $V_m = 22.414$ L/mol. Burning one mole of methane, $\text{CH}_4 + 2\text{O}_2 \rightarrow \text{CO}_2 + 2\text{H}_2\text{O}$, consumes 2 mol of oxygen and produces 1 mol of carbon dioxide, which is 44.83 L of O_2 in and **22.41 L of CO_2 out** at STP. Note that the two moles of water do not appear in that volume tally, because at 0 °C the water is a liquid and the equation only counts gases.

The constant is not independent; it is $PV = nRT$ evaluated once. $V_m = RT/P = (8.314 \times 273.15)/101\,325 = 0.022414$ m³/mol, and every other molar volume in circulation is the same calculation at different conditions. Historically the logic ran the other way. Gay-Lussac reported in 1808 that gases combine in simple whole-number volume ratios, two volumes of hydrogen to one of

oxygen, and Avogadro's 1811 hypothesis explained why: equal volumes hold equal counts, so the volume ratios *are* the mole ratios of the balanced equation. That insight was ignored for half a century before Cannizzaro revived it at Karlsruhe in 1860.

The error that dominates this page is using 22.4 L/mol for conditions that are not STP. Room temperature is not 0 °C. At 25 °C and 1 atm the molar volume is 24.47 L/mol, so applying 22.4 to a bench-top measurement understates the volume by 9%. And "standard conditions" is not one definition but several: classic STP gives 22.414 L/mol, IUPAC's post-1982 STP (0 °C, 100 kPa) gives 22.711, and SATP (25 °C, 100 kPa) gives 24.79. The spread between the smallest and largest is over 10%, which is far more than the precision most people think they are carrying. Check which convention a source assumes before you borrow its number, and when the conditions are anything other than a listed standard, abandon the shortcut and use the ideal gas law directly.

Two smaller cautions. This is an ideal-gas result, so it degrades for gases near their condensation point: ammonia and sulfur dioxide at STP already deviate by a percent or two, and water vapour at 0 °C is not a gas at all. And when you use volume ratios as mole ratios in a reaction, they apply only to the species that are actually gaseous at the stated conditions. A dissolved or condensed product contributes no volume, and counting it is the quiet way to get a stoichiometry problem wrong.

Worked example 4.6. Given $n = 1$ mol; find V (gas volume at stp). *1 mol ideal gas at STP $\rightarrow$ 22.414 L.*

4.7 Gas Density from Molar Mass

(Chemistry 11-15) $\rho = \frac{PM}{RT}$

Where,

ρ = Gas density (kg/m³)

P = Pressure (kPa)

M = Molar mass (g/mol)

T = Absolute temperature (°C)

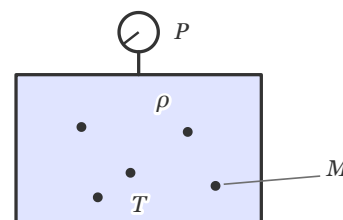


Figure 4.7: gas density from molar mass.

Start from $PV = nRT$, substitute $n = m/M$, and rearrange for mass over volume: $\rho = PM/RT$. What the result says is that a gas has no density of its own. Unlike a solid or a liquid, whose density is close enough to a fixed property to tabulate, a gas takes whatever density its pressure and temperature impose, and the only thing the substance itself contributes is M . Squeeze it and it gets denser in exact proportion; warm it and it thins in inverse proportion.

Air at 101.325 kPa and 20 °C, using $M = 0.028964$ kg/mol: $\rho = (101\,325 \times 0.028964)/(8.314 \times 293.15) = 1.204$ kg/m³, the figure every ventilation calculation starts from. Helium at the same conditions, with $M = 0.0040026$, comes to 0.166 kg/m³. Subtract, and a cubic metre of helium lifts about **1.04 kg**. That is the entire physics of a party balloon, and it explains why a balloon large enough to lift a person has to be the size of a house.

Rearranged for molar mass the equation becomes a measurement rather than a prediction, and a historically important one. Weigh a bulb of known volume empty, fill it with a vapour at measured temperature and pressure, weigh it again, and $M = \rho RT/P$ hands you the molar mass of an unknown. This is the Dumas method, and through the middle of the nineteenth century it was one of the few

routes to a molecular formula. It also settled arguments: measured vapour densities are what showed that many elemental gases travel as diatomic molecules rather than lone atoms.

The dominant error here is the molar mass unit, and it is a clean factor of a thousand. With $R = 8.314$ J/(mol·K) the equation demands M in *kilograms* per mole. Air is 0.029 kg/mol. Enter 29 and the answer comes back as 1204 kg/m³, air denser than water, which at least announces itself. Enter 0.029 when the calculation wanted grams and you get the mirror error. The usual companions apply too: P must be absolute, not a gauge reading, and T must be in kelvin.

Two conceptual notes. There is no such thing as "the molar mass of air" in the strict sense: air is a mixture, and 28.96 g/mol is a mole-weighted average of nitrogen, oxygen and argon. It works precisely because an ideal gas is indifferent to what its neighbours are; only the total count matters. That same indifference produces a result most people find backwards: *humid air is lighter than dry air*. Water is 18 g/mol against air's 29, so at a given pressure and temperature every water molecule that joins the mixture has displaced a heavier one. Muggy days are low-density days, which is why aircraft performance charts include humidity and why a hot, humid runway is a long takeoff.

Worked example 4.7. Given $P = 1$ atm, $M = 32.00$ g/mol, $T = 273.15$ K; find ρ (gas density). *O₂ at STP (1 atm, 273.15 K) → 1.42768 kg/m³.*

Module 4 practice

39. Boyle's law: A gas bubble of 8 mL forms at 150 kPa. Rising water pressure squeezes it down to 3 mL at the same temperature.

Calculate the pressure inside the squeezed bubble.

40. Boyle's law: A gas bubble of 9 mL forms at 160 kPa. Rising water pressure squeezes it down to 3 mL at the same temperature.

Calculate the pressure inside the squeezed bubble.

41. Hot gases: A balloon holds 8 L of air at 127 °C. It is warmed to 227 °C at constant pressure. *Determine the balloon's new volume.*

42. Hot gases: An aerosol can reads 100 kPa at 27 °C. Its rated burst pressure is 220 kPa. *Determine the temperature, in kelvin, at which the can would reach its burst pressure.*

43. The combined gas law: A gas sample of 900 mL at 100 kPa and 300 K is transferred into a smaller vessel of 600 mL and brought to 400 K. *Calculate the pressure in the new vessel.*

44. The combined gas law: A gas sample sits at 200 kPa in 5 L at 250 K. After a process, the same sample is found at 100 kPa in 8 L. *Determine the sample's final temperature, in kelvin.*

45. The ideal gas law: A rigid 10 L tank holds 3 mol of nitrogen at 275 K. *Calculate the pressure in the tank.*

46. The ideal gas law: A flexible gas bag holds 2 mol of helium at 150 kPa and 280 K. *Determine the volume of the bag.*

47. $PV = nRT$, every way: A steel cylinder in the stockroom is stamped 9 L. Its gauge reads 350 kPa on a day the room sits at 127 °C. The order form wants the contents by amount, not by volume. *Determine the amount of gas in the cylinder.*

48. $PV = nRT$, every way: Before a demonstration, 3 mol of carbon dioxide is charged into a rigid 15 L vessel in a prep room at 27 °C. The safety sheet wants the resulting pressure on record. *Calculate the pressure in the vessel.*

49. Molar volume at STP: A reaction in the fume hood produces 2 mol of hydrogen gas, collected at STP. *Determine the volume the hydrogen occupies.*

50. Molar volume at STP: A gas bag holds 22.4 L of carbon dioxide, measured at STP. *Determine the amount of carbon dioxide in the bag.*

51. Weighing a gas: A balloon is filled with methane ($M = 16.0$ g/mol) at 175 kPa and 27 °C. Air under the same conditions runs about 1.2 g/L. *Calculate the density of the methane, in grams per litre.*

52. Weighing a gas: A storage cylinder delivers compressed air ($M = 29.0 \text{ g/mol}$) into a vessel held at 200 kPa and 27 °C. *Determine the density of the compressed air in the vessel, in grams per litre.*

53. The Pressure Test: Launch morning. A sounding balloon is part-filled with 11.2 L of helium, measured at STP, before the neck is tied off. No calculator today — $V_m = 22.4 \text{ L/mol}$, and every number is chosen to fit in your head. *Determine the amount of helium sealed into the balloon.*

54. The Pressure Test: Lift-off. The balloon rises reading 10 L at 90 kPa and 300 K. At reporting altitude the pressure gauge shows 30 kPa and the thermometer 200 K. Work the change one ratio at a time — line one feeds line two. *Determine the balloon's volume at reporting altitude.*

MODULE 5

Water by the Numbers

Parts per million · Hard water · Grains per gallon · The CaCO_3 yardstick · TDS from conductivity · Chlorine dose · Feeding the plant · The Treatment Plant

5.1 Nutrient ppm from Fertiliser Weight

(Chemistry 11-16)
$$c = \frac{m f}{V}$$

Where,

c = Nutrient concentration (ppm)

m = Fertiliser weight (g)

f = Element share of the fertiliser (%)

V = Water volume (L)

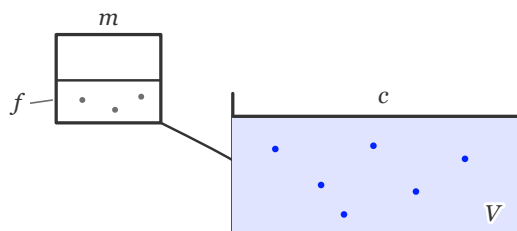


Figure 5.1: nutrient ppm from fertiliser weight.

This is the arithmetic behind every hydroponic feed chart, and doing it yourself is what makes a nutrient program legible rather than a matter of following the bottle. The concentration of a nutrient element is the fertiliser weight times the element's share of that fertiliser, divided by the water volume.

One convenient fact makes the mental version easy: because a litre of water weighs a kilogram, **one gram in one hundred litres is ten ppm**. So 10 g of calcium nitrate at 15.5% nitrogen into 100 L gives 1.55 g of N in 100 L, which is 15.5 ppm.

The trap is the same one that catches field growers, and it is worth stating plainly: fertiliser labels report phosphorus and potassium as oxides, not elements. The P in an N-P-K analysis is P_2O_5 and the K is

K_2O . To get elemental phosphorus multiply by 0.436; for elemental potassium multiply by 0.830. Hydroponic recipes are almost always written in elemental ppm, so feeding to a label's P number without converting overfeeds phosphorus by a factor of 2.3.

Reference targets for a general vegetative feed, in elemental ppm: N 150–200, P 50, K 200, Ca 150–200, Mg 50, S 60, with iron near 3 and the other micronutrients below 1. Those numbers are why calcium and magnesium deserve their own attention: they are needed at concentrations comparable to phosphorus, and a base nutrient designed for soft water will not supply them.

Worked example 5.1. Given $m = 10 \text{ g}$, $f = 15.5 \%$, $V = 100 \text{ L}$; find c (nutrient concentration). *10 g at 15.5% N into 100 L → 15.5 ppm N.*

5.2 TDS Estimated from Conductivity ($TDS = k \times EC$)

(Chemistry 11-17) $TDS = k \times EC$

Where,

TDS = Total dissolved solids (mg/L)

k = TDS/EC factor

EC = Electrical conductivity ($\mu S/cm$)

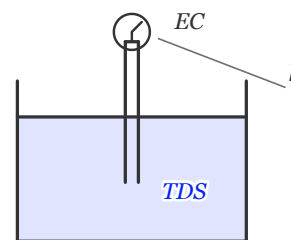


Figure 5.2: tds estimated from conductivity ($tds = k \times ec$).

Evaporating a litre of water to dryness and weighing the residue is the only true TDS measurement, and it takes a drying oven and several hours. A conductivity meter gives an answer in two seconds, and because dissolved salts are what carry the current the two track each other closely enough for daily work: TDS in $mg/L \approx k \times$ conductivity in $\mu S/cm$. The factor k is not a constant of nature: it depends on which ions dominate. Chloride-rich waters run near 0.55, typical fresh surface and well waters around 0.65, and sulphate- or bicarbonate-heavy waters climb toward 0.70. A tower reading 500 $\mu S/cm$ at $k = 0.65$ is carrying roughly 325 mg/L of dissolved solids.

Two things will bite you. First, conductivity is strongly temperature-dependent, about 2% per $^{\circ}C$, so every meaningful reading is temperature-compensated to 25 $^{\circ}C$; a probe without compensation reading warm blowdown will overstate TDS by 15% or more. Second, non-ionic dissolved solids are invisible to the meter. Silica, dissolved organics and sugars contribute real gravimetric TDS but carry almost no current, which is why a high-silica cooling water or a food-plant effluent can gravimetrically test far above what its conductivity predicts. Calibrate k against one oven-dried sample of your own water and it becomes a genuinely reliable daily tool.

Worked example 5.2. Given $k = 0.65$, $EC = 500 \mu S/cm$; find TDS (total dissolved solids). $500 \mu S/cm$ at $k = 0.65 \rightarrow TDS = 325 mg/L$.

5.3 Total Hardness as $CaCO_3$

(Chemistry 11-18) $TH = 2.497 Ca + 4.118 Mg$

Where,

TH = Total hardness as $CaCO_3$ (mg/L)

Ca = Calcium as Ca^{2+} (mg/L)

Mg = Magnesium as Mg^{2+} (mg/L)

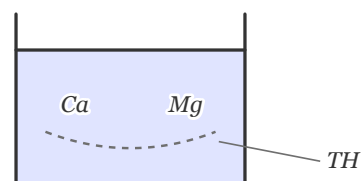


Figure 5.3: total hardness as $CaCO_3$.

Hardness is a property, not a substance: it is the combined effect of every divalent cation in the water, and in practice calcium and magnesium account for essentially all of it. Because they have different atomic weights, the trade converts both to a single reference salt, calcium carbonate, so that numbers from different labs, different waters and different treatment steps can be added and subtracted. The factors come straight from equivalent weights: $50.04/20.04 = 2.497$ for calcium and $50.04/12.15 = 4.118$ for magnesium. A well water reporting 80 mg/L Ca and 25 mg/L Mg therefore has a total hardness of $2.497 \times 80 + 4.118 \times 25 = 199.8 + 103.0 = 303 mg/L$ as $CaCO_3$, or about 17.7 grains per gallon.

The single most common field error is reading a lab sheet that already says "as $CaCO_3$ " and multiplying by 2.497 anyway: instant 2.5 $\times$ overstatement, an oversized softener and a customer paying for salt they do not need. Check the units column before you touch the calculator. The second trap is the magnesium factor: because Mg^{2+} is light, a modest 25 mg/L of it contributes as much hardness as 41 mg/L of calcium, which is why waters from dolomitic aquifers feel far harder than their calcium number suggests and why magnesium-heavy water scales heat exchangers with a stubborn magnesium silicate rather than easily-acid-cleaned calcite.

Worked example 5.3. Given $\text{Ca} = 80 \text{ ppm}$, $\text{Mg} = 25 \text{ ppm}$; find TH (total hardness as CaCO_3). $\text{Ca } 80 \text{ mg/L} + \text{Mg } 25 \text{ mg/L} \rightarrow \text{TH} = 302.7 \text{ mg/L as CaCO}_3$.

5.4 Grains per Gallon $\leftrightarrow$ ppm Hardness

(Chemistry 11-19) $H = 17.118 G$

Where,

H = Hardness as CaCO_3 (mg/L)

G = Hardness in grains per US gallon (gpg)

The grain is a survivor from the old apothecaries' system (1/7000 of a pound, originally the weight of a single barleycorn) and North American water treatment never let it go. One grain of CaCO_3 dissolved in one US gallon is 64.79891 mg in 3.785412 L, which is 17.118 mg/L. That is the whole conversion: 10 gpg is 171 mg/L, and a residential softener rated "30,000 grains" will remove 30,000 grains of hardness before it needs salt. Softener control valves, resin capacity tables and salt-dose settings are all published in grains, while every laboratory report and every regulatory limit is in mg/L, so the technician lives in both worlds at once.

Watch out for the imperial gallon. A British or Canadian "grain per gallon" historically meant grains per *imperial* gallon (4.546 L), which works out to 14.25 mg/L: a 20% difference that will undersize a softener if you take a UK spec sheet at face value. And remember the informal hardness bands are quoted in grains: under 1 gpg is soft, 1–3.5 slightly hard, 3.5–7 moderately hard, 7–10.5 hard, and anything over 10.5 gpg (180 mg/L) is very hard and will scale a water heater within a few years untreated.

Worked example 5.4. Given $G = 10$; find H (hardness as CaCO_3). $10 \text{ grains per gallon} \rightarrow 171.18 \text{ mg/L as CaCO}_3$.

5.5 Ion Concentration as CaCO_3 Equivalent

(Chemistry 11-20) $C_{\text{CaCO}_3} = C_{\text{ion}} \times \frac{50.04}{\text{EW}}$

Where,

C_{CaCO_3} = Concentration as CaCO_3 (mg/L)

C_{ion} = Concentration as the ion (mg/L)

EW = Equivalent weight of the ion (g/mol)

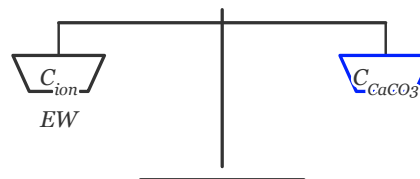


Figure 5.5: ion concentration as CaCO_3 equivalent.

A water analysis lists a dozen ions with a dozen different atomic weights, and you cannot add or subtract them directly: 20 mg/L of calcium and 20 mg/L of sodium are not the same amount of chemistry. Expressing everything "as CaCO_3 " fixes that by converting each ion to the mass of calcium carbonate that carries the same number of charge equivalents. Calcium carbonate has a molar mass of 100.09 and a valence of 2, so its equivalent weight is 50.04 g/eq; divide that by the ion's own equivalent weight and you have the factor. Calcium (EW 20.04) gets 2.497, magnesium (12.15) gets 4.118, sodium (23.0) gets 2.18, bicarbonate (61.0) gets 0.82, and sulphate (48.03) gets 1.04. Forty mg/L of Ca^{2+} becomes $40 \times 2.497 = 99.9 \text{ mg/L as CaCO}_3$.

Once every ion is on the CaCO_3 scale you can do the things that make an analysis useful: check that cations balance anions to within a few percent (the classic sanity test on any lab report), split total hardness into carbonate and non-carbonate fractions by comparing hardness to alkalinity, and size a softener or a dealkalizer directly from the numbers. The trap is the word "equivalent": the divisor is the equivalent weight, molar mass divided by charge, not the molar mass. Using 40.08 for calcium instead of 20.04 halves every hardness figure you produce.

Worked example 5.5. Given $C_{\text{ion}} = 40 \text{ ppm}$, $\text{EW} = 20.04 \text{ g/mol}$; find C_{CaCO_3} (concentration as CaCO_3). $40 \text{ mg/L Ca}^{2+} \text{ at EW } 20.04 \rightarrow 99.88 \text{ mg/L as CaCO}_3$.

5.6 Equivalent Weight from Molar Mass and Valence

(Chemistry 11-21)
$$EW = \frac{M}{z}$$

Where,

EW = Equivalent weight (g/mol)

M = Molar mass (g/mol)

z = Valence (equivalents per mole)

Reactions in water happen equivalent-for-equivalent, not gram-for-gram: one charge neutralizes one charge, one H^+ neutralizes one OH^- . The equivalent weight is simply the mass of a species that carries one mole of that reacting capacity: molar mass divided by valence. Calcium carbonate, $M = 100.087$ with a divalent cation, has $EW = 50.04$ g/eq, which is why 50.04 turns up in every "as $CaCO_3$ " conversion in this trade. Sulphuric acid, $M = 98.08$ with two replaceable protons, has $EW = 49.04$ g/eq, so a pound of pure H_2SO_4

Worked example 5.6. Given $M = 100.087$ g/mol, $z = 2$; find EW (equivalent weight). $CaCO_3$: M 100.087 g/mol, $z = 2 \rightarrow EW = 50.0435$ g/eq.

5.7 Chlorine Dose, Demand and Residual

(Chemistry 11-22)
$$D = C_{\text{demand}} + C_{\text{res}}$$

Where,

D = Chlorine dose applied (mg/L)

C_{demand} = Chlorine demand (mg/L)

C_{res} = Chlorine residual (mg/L)

Chlorine added to water does not stay chlorine. It is consumed by iron, manganese, sulphide, ammonia, and every scrap of organic matter it meets, that consumption is the demand, and only what survives shows up on a DPD test as residual. The bookkeeping is a straight sum: dose equals demand plus residual. Feed 3.3 mg/L into a water with 2.5 mg/L of demand and you keep 0.8 mg/L residual, which is a typical distribution target. There is no way around the demand; it has to be satisfied before a single tenth of a milligram of residual appears, which is why a well with high iron can swallow 4 mg/L and still test zero at the tap.

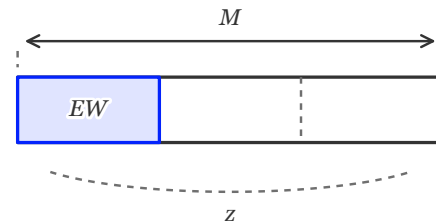


Figure 5.6: equivalent weight from molar mass and valence.

neutralizes almost exactly a pound of alkalinity expressed as $CaCO_3$: the near-1:1 coincidence that lets operators do acid-feed arithmetic in their heads.

The valence you divide by is the one that applies to the reaction in question, and that is where people go wrong. Sodium carbonate reacting as a base uses $z = 2$ (EW 53.0), but if you are counting sodium ions it behaves as $z = 2$ for a different reason. Bicarbonate is $z = 1$ (EW 61.0) in the alkalinity titration but the carbonate it came from is $z = 2$ (EW 30.0). Write down which reaction you mean before you pick z , and the rest of the water chemistry falls into line.

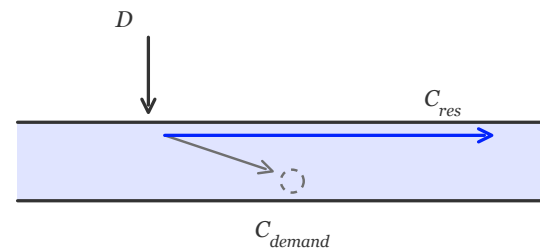


Figure 5.7: chlorine dose, demand and residual.

The practical procedure is a chlorine demand study: dose a set of bottles of the actual water at increasing rates, wait the real contact time (usually 30 minutes, but overnight for a distribution main) and test each. The plot of residual against dose is flat until demand is satisfied, then rises with unit slope. Two field cautions. Demand is not a fixed property: it climbs in warm weather, after a main break, and with every degree of surface-water turbidity, so the study needs repeating seasonally. And test at the far end of the system, not at the plant: the residual that matters legally and microbiologically is the one still standing at the last service connection.

Worked example 5.7. Given $C_{\text{demand}} = 2.5$ ppm, $C_{\text{res}} = 0.8$ ppm; find D (chlorine dose applied). $2.5 \text{ mg/L demand} + 0.8 \text{ mg/L residual} \rightarrow 3.3 \text{ mg/L dose}$.

5.8 Chemical Feed Rate (lb/day = mg/L × MGD × 8.34)

(Chemistry 11-23) $\dot{m} = C Q$

Where,

$\dot{m}$ = Chemical feed rate (kg/h)

C = Dose (mg/L)

Q = Water flow rate (L/min)

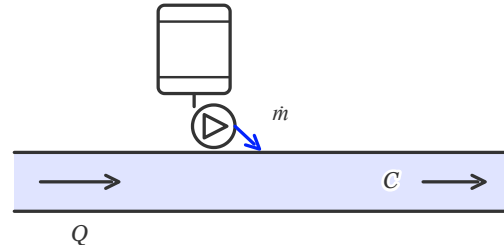


Figure 5.8: chemical feed rate (lb/day = mg/l × mgd × 8.34).

Every North American operator learns this one as a chant: pounds per day equals milligrams per litre times million gallons per day times 8.34. The 8.34 is nothing mysterious: it is the weight of a US gallon of water, 8.345 lb at 60 °F, rounded. Dosing 2 mg/L of chlorine into a 1 MGD plant needs $2 \times 1 \times 8.34 = 16.7$ lb/day. Underneath the trade constant it is simply concentration times flow, mass balance in its plainest form, and this page will do it in kilograms per hour and cubic metres per hour just as happily.

Three warnings from the field. First, this gives pounds of *active chemical*; if you are feeding 12.5% sodium hypochlorite you still have to divide by the strength to get pounds of product, and again by the solution's density to get gallons the pump must deliver. Second, the 8.34 quietly assumes the water is water: feeding into a brine, a slurry or a glycol loop needs the actual density. Third, calibrate the pump against a drawdown cylinder rather than trusting the dial: a diaphragm pump's stroke setting drifts with discharge pressure, and the difference between a calculated feed rate and a measured one is where half of all chemical overspend hides.

Worked example 5.8. Given $\dot{m} = 5$ kg/h, $Q = 100 \text{ m}^3/\text{h}$; find C (dose). $5 \text{ kg/h into } 100 \text{ m}^3/\text{h} \rightarrow 50 \text{ mg/L dose}$.

Module 5 practice

55. Hard water: The lab reports an apartment building's incoming water at 60 mg/L Ca^{2+} and 25 mg/L Mg^{2+} . *Determine the total hardness as CaCO_3 .*

56. Hard water: A rural well sample titrates at 60 mg/L of calcium and 30 mg/L of magnesium. *Calculate the total hardness, expressed as CaCO_3 .*

57. Grains per gallon: A softener installer receives a lab report stating 250 mg/L as CaCO_3 . The softener's hardness dial reads in grains per US gallon. *Determine the dial setting in grains per gallon.*

58. Grains per gallon: A softener's dial has been set at 6 grains per gallon for years. The homeowner's new lab report speaks mg/L as CaCO_3 . *Calculate the hardness the dial represents, in mg/L as CaCO_3 .*

59. The CaCO_3 yardstick: A water-chemistry table lists zinc, Zn^{2+} with a molar mass of 65.38 g/mol and a charge of 2. *Determine the ion's equivalent weight.*

60. The CaCO_3 yardstick: A lab measures 80 mg/L of calcium, Ca^{2+} , whose equivalent weight is 20.04 g/eq. *Determine the concentration expressed as CaCO_3 .*

61. TDS from conductivity: A handheld meter dipped in a rural well reads 800 $\mu\text{S}/\text{cm}$. The local lab's calibration factor for this aquifer is $k = 0.68$. *Estimate the total dissolved solids.*

62. TDS from conductivity: A greenhouse nutrient tank measures 600 $\mu\text{S}/\text{cm}$. The grower's chart uses $k = 0.57$ for this feed blend. *Estimate the total dissolved solids.*

63. Chlorine dose: A well is dosed at 2.2 mg/L of chlorine. Downstream, the free residual measures 0.7 mg/L. *Determine the water's chlorine demand.*

64. Chlorine dose: Jar tests on a surface-water source show it consumes 2 mg/L of chlorine. The operator must keep a free residual of 0.7 mg/L in the distribution system. *Calculate the chlorine dose to apply.*

65. Feeding the plant: A treatment plant doses alum at 20 mg/L. The plant treats 0.5 MGD — million US gallons per day. *Calculate the alum feed rate in pounds per day.*

66. Feeding the plant: A hypochlorinator delivers 208.5 lb/day of chlorine into a plant flow of 2.5 MGD. *Determine the dose the water receives, in mg/L.*

67. The Treatment Plant: Last sample of the shift. A raw-water sample lands on the bench: the meter reads 400 $\mu\text{S}/\text{cm}$ (this source uses $k = 0.6$), and the titration returns 40 mg/L of calcium and 20 mg/L of magnesium — the trade factors round to 2.5 and 4 on this paper. Yesterday's jar test fed 2 mg/L of chlorine and 0.4 mg/L of residual survived. Work each line — every answer feeds the next. *Determine tomorrow's chlorine dose to hold a 0.6 mg/L residual — meter to pump, one line at a time.*

68. The Treatment Plant: Bonus mark, worked backwards: the day book shows a TDS of 280 mg/L logged beside an EC of 400 $\mu\text{S}/\text{cm}$ — but the k column is smudged beyond reading. *Determine the factor k the operator used.*

MODULE 6

Acids, Bases & pH

The pH scale · pH from concentration · Concentration from pH · pH plus pOH · Powers of ten · Titration · The Litmus Final

6.1 pH from Hydrogen Ion Concentration

(Chemistry 11-24) $\text{pH} = -\log_{10} [\text{H}^+]$

Where,

$$\text{pH} = \text{pH}$$

$[\text{H}^+]$ = Hydrogen ion concentration (M)

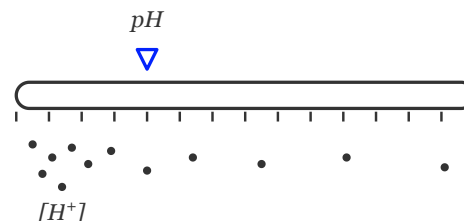


Figure 6.1: pH from hydrogen ion concentration.

Hydrogen ion concentrations in water span more than ten orders of magnitude, so in 1909 the Danish chemist Søren Sørensen, working at the Carlsberg brewery laboratory, compressed them onto a logarithmic scale. Each pH unit is a factor of ten: lemon juice at pH 2 carries a hundred times the hydrogen ion concentration of tomato juice at pH 4. Pure water at 25 °C sits at pH 7, where $[\text{H}^+] = 1.0 \times 10^{-7}$ mol/L. The logarithm takes the concentration in mol/L, and the inverse direction is exact: $[\text{H}^+] = 10^{-\text{pH}}$.

In water treatment pH is the master variable. Boiler water is typically held between 10.5 and 11.5 to suppress corrosion, cooling towers near 7–9 to balance scale against corrosion, and municipal drinking water around 7.0–8.5. A pH swing of a single unit, a tenfold chemistry change, is often the first sign that a chemical feed pump has failed.

Worked example 6.1. Given $[\text{H}^+] = 0.001 \text{ M}$; find pH (ph). $[\text{H}^+] = 1.0\text{e-}3 \text{ mol/L} \rightarrow \text{pH} = 3$.

6.2 pH and pOH Relation

(Chemistry 11-25) $\text{pH} + \text{pOH} = 14$

Where,

$$\text{pH} = \text{pH}$$

$$\text{pOH} = \text{pOH}$$

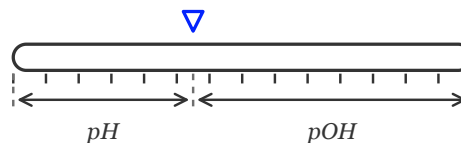


Figure 6.2: pH and pOH relation.

Water is never merely water. A small fraction of it is always pulled apart into H^+ and OH^- , and the equilibrium between them fixes the product of their concentrations at a constant: $K_w = [\text{H}^+][\text{OH}^-] = 1.0 \times 10^{-14}$ at 25 °C. Because it is a *product* that is constant, the two concentrations are locked in a see-saw: push one up and the other must come down by the same factor. Take negative logarithms of both sides and the product becomes a sum: $\text{pH} + \text{pOH} = 14$. So pH and pOH are not two independent measurements. They are one number written two ways, and knowing either gives the other for free.

A worked case. Dissolve enough sodium hydroxide to make a 0.010 M solution: it dissociates completely, so $[\text{OH}^-] = 1.0 \times 10^{-2}$ and $\text{pOH} = 2.00$. Then $\text{pH} = 14 - 2.00 = 12.00$. Check it the long way, $[\text{H}^+] = K_w/[\text{OH}^-] = 10^{-14}/10^{-2} = 10^{-12}$, giving pH 12.00, and the two routes agree, as they must.

The "14" is not a magic number, it is $\text{p}K_w$, and the exact value at 25 °C is 13.995. The whole logarithmic apparatus came from Søren Sørensen at the Carlsberg laboratory in 1909, who needed a compact way to talk about hydrogen ion concentrations spanning more than ten orders of magnitude in brewing. The p-prefix now attaches to anything worth compressing that way: pOH, pKa, pKw, pCa.

Worked example 6.2. Given $\text{pH} = 4.75$; find pOH (poh). $\text{pH } 4.75 \rightarrow \text{pOH} = 9.25$.

The trap is that 14 belongs to 25 °C and nothing else. Self-ionisation is endothermic, so heating water drives it forward and K_w rises: $\text{p}K_w$ is about 14.9 at 0 °C, 13.0 at 60 °C, and 12.0 at 100 °C. Which means *neutral* water at 100 °C has a pH of 6.0, and it is still perfectly neutral, not acidic, because pH still equals pOH. This matters wherever hot water is measured: boiler water sampled at temperature and boiler water sampled after cooling do not read the same, and the difference is the equation moving, not the chemistry changing. Anyone reading pH on a hot sample needs to know whether the meter is compensating and what reference temperature it is compensating to.

Which leads to the deeper misconception. **Neutral does not mean pH 7; neutral means pH equals pOH.** The two coincide at 25 °C by arithmetic accident, and part company at every other temperature. Two smaller notes: the relation is aqueous only (it says nothing about a non-aqueous solvent, which has its own autoionisation constant if it has one at all) and the pH scale is not fenced between 0 and 14. Those bounds are just where 1 M solutions land. Concentrated hydrochloric acid at 12 M has a negative pH, and concentrated sodium hydroxide runs above 14, with the sum still holding.

6.3 Titration: Concentration of an Unknown

(Chemistry 11-26) $C_a = \frac{n C_b V_b}{V_a}$

Where,

C_a = Analyte concentration (M)

V_a = Analyte volume (aliquot) (L)

C_b = Titrant concentration (M)

V_b = Titre volume delivered (L)

n = Mole ratio (analyte per titrant)

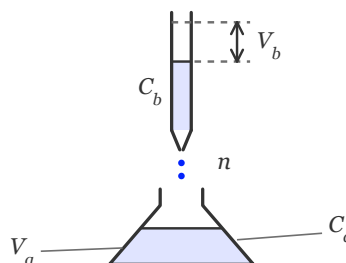


Figure 6.3: titration: concentration of an unknown.

At the equivalence point the moles of titrant delivered, $C_b V_b$, exactly match the moles of analyte present, $C_a V_a$, scaled by the balanced equation's mole ratio n . Everything else in a titration (the burette, the indicator, the swirling) exists only to find that point precisely. Titrate a 25.00 mL aliquot of hydrochloric acid with 0.1000 M sodium hydroxide and take 23.45 mL to reach the endpoint: with $n = 1$, $C_a = (1 \times 0.1000 \times 23.45)/25.00 = 0.09380$ M, good to four figures from nothing but glassware.

The mole ratio is where marks are lost. For a diprotic acid such as H_2SO_4 titrated with NaOH, one mole of acid consumes two of base, so $n = 0.5$: a 25.00 mL aliquot needing 30.00 mL of 0.100 M NaOH is

$0.5 \times 0.100 \times 30.00/25.00 = 0.0600$ M. Karl Friedrich Mohr systematised the whole technique in his 1855 *Lehrbuch der chemisch-analytischen Titrimethode*, introducing the burette clamp and the pinchcock that made reproducible volumetric analysis possible; his methods still underpin water-hardness and chlorine testing today. Note the difference between the endpoint (where the indicator changes) and the equivalence point (where the stoichiometry balances): the gap between them is the indicator error, which is why the indicator is chosen to change colour on the steep part of the titration curve.

Worked example 6.3. Given $V_a = 25.00$ mL, $C_b = 0.1000$ M, $V_b = 23.45$ mL, $n = 1$; find C_a (analyte concentration). *25.00 mL HCl vs 23.45 mL of 0.1000 M NaOH $\rightarrow$ 0.09380 M.*

Module 6 practice

69. pH from concentration: A calibrated meter puts the hydrogen-ion concentration of a tomato-juice sample at 1×10^{-4} mol/L. *Calculate the pH of the sample.*

70. pH from concentration: A calibrated meter puts the hydrogen-ion concentration of a cola sample left overnight to go flat at 1×10^{-3} mol/L. *Calculate the pH of the sample.*

71. Concentration from pH: An acid-rain sample from the roof gauge measures pH 4 on a freshly calibrated meter. *Determine its hydrogen-ion concentration.*

72. Concentration from pH: The lemon juice in a fruit-acidity lab measures pH 2 on a freshly calibrated meter. *Determine its hydrogen-ion concentration.*

73. pH plus pOH: The QC binder quotes a descaling bath at pOH 12. *Determine the bath's pH.*

74. pH plus pOH: The QC binder quotes a descaling bath at pOH 8. *Determine the bath's pH.*

75. Powers of ten: During a stream survey, the outfall sample reads pH 2 while the upstream control reads pH 5. *Determine how many times more acidic the outfall sample is — then its hydrogen-ion concentration.*

76. Powers of ten: Two cleaners share the janitor's shelf: an all-purpose spray at pH 9 and a drain gel at pH 10. *Determine how much more basic the drain gel is — then its pOH.*

77. Titration: A 50.0 mL sample of hydrochloric acid is pipetted into a flask with a few drops of indicator. From the burette, 0.1 mol/L sodium hydroxide brings it to the endpoint at 40.0 mL delivered. Acid and base react one-to-one. *Determine the concentration of the acid.*

78. Titration: A 20.0 mL sample of hydrochloric acid is pipetted into a flask with a few drops of indicator. From the burette, 0.1 mol/L sodium hydroxide brings it to the endpoint at 30.0 mL delivered. Acid and base react one-to-one. *Determine the concentration of the acid.*

79. The Litmus Final: The last station of the practical. A 50.0 mL aliquot of hydrochloric acid — strong, fully dissociated — meets the burette, and the endpoint lands at exactly 25.0 mL of 0.2 mol/L sodium hydroxide, one-to-one. The numbers are chosen to work in your head, and every answer feeds the next line. *Determine the acid's concentration, its pH, and its pOH — one line at a time.*

80. The Litmus Final: Bonus mark, from the basic end: the rinse bath beside the station lists a hydroxide-ion concentration of 1×10^{-4} mol/L. *Determine the bath's pH — two hops, no calculator.*

Answer key

- | | | |
|--|----------------|---------------------------------------|
| 1. 3 mol | 27. 1.5 mol | 54. 30 L |
| 2. 2 mol | 28. 0.2 mol | 55. 252.77 mg/L as CaCO ₃ |
| 3. 80 g | 29. 15 % | 56. 273.36 mg/L as CaCO ₃ |
| 4. 58.5 g/mol | 30. 5 % | 57. 14.6045 gpg |
| 5. 2,408,000,000,000,000,000,000,000 atoms | 31. 0.5 mol/L | 58. 102.708 mg/L as CaCO ₃ |
| 6. 2 mol | 32. 0.25 mol/L | 59. 32.69 g/eq |
| 7. 40 % | 33. 75 mL | 60. 199.76 mg/L as CaCO ₃ |
| 8. 72.73 % | 34. 3.5 mol/L | 61. 544 mg/L |
| 9. 2 mol | 35. 2 mol | 62. 342 mg/L |
| 10. 0.5 mol | 36. 0.75 mol/L | 63. 1.5 mg/L |
| 11. 18 g/mol | 37. 1 mol | 64. 2.7 mg/L |
| 12. 18 g/mol | 38. 2 mol/L | 65. 83.4 lb/day |
| 13. 60 % | 39. 400 kPa | 66. 10 mg/L |
| 14. 60 % | 40. 480 kPa | 67. 240 mg/L |
| 15. 15 g | 41. 10 L | 68. 0.7 (no unit) |
| 16. 40 g | 42. 660 K | 69. 4 (no unit) |
| 17. 1,050 kJ | 43. 200 kPa | 70. 3 (no unit) |
| 18. 625 kJ | 44. 200 K | 71. 0.0001 mol/L |
| 19. 3 mol | 45. 685.9 kPa | 72. 0.01 mol/L |
| 20. 4 mol | 46. 31 L | 73. 2 (no unit) |
| 21. 25 % | 47. 0.95 mol | 74. 6 (no unit) |
| 22. 6,000 kJ | 48. 498.8 kPa | 75. 0.01 mol/L |
| 23. 2.5 mol | 49. 44.8 L | 76. 4 (no unit) |
| 24. 44 g | 50. 1 mol | 77. 0.08 mol/L |
| 25. 1.5 mol/L | 51. 1.12 g/L | 78. 0.15 mol/L |
| 26. 0.4 mol/L | 52. 2.33 g/L | 79. 0.1 mol/L |
| | 53. 0.5 mol | 80. 10 (no unit) |