

The Mole: Counting Atoms by Weighing

- **Goal:** connect invisible atoms and molecules to masses a balance can measure
- **Central question:** how do we count particles far too small to see?
- **Roadmap:** carbon-12 standard, the mole, Avogadro's number, molar mass, calculations
- **Why it matters:** predicting reaction amounts, reading labels, preparing solutions
- For students, returning adult learners, and educators



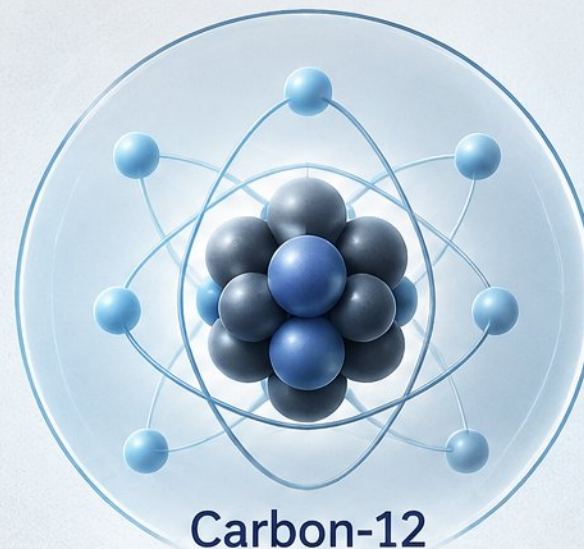
Why Chemists Count By Weighing

- Atoms are far too small to count one by one
- Reactions need fixed particle ratios, not convenient masses
- Weighing is fast, precise, and repeatable
- Mass is the practical bridge to particle count
- Like buying rice, nails, or coins by the kilogram



Relative Mass and the Carbon-12 Standard

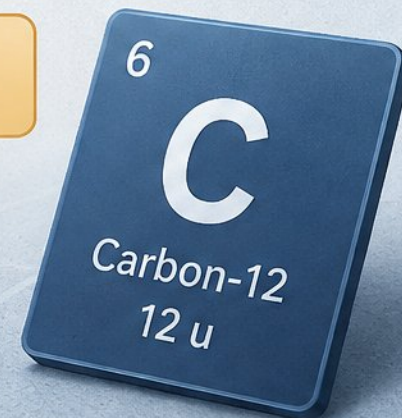
- Atomic masses are relative, not absolute
- Carbon-12 is the reference: exactly 12 u
- Unified atomic mass unit (u) = dalton (Da)
- $1 \text{ Da} \approx 1.660\,539\,066\,60 \times 10^{-27} \text{ kg}$
- Periodic table masses are isotope-weighted averages



Unified atomic
mass unit (u)

dalton (Da)

$$1 \text{ u} = 1 \text{ Da}$$



Reading Atomic Masses from the Periodic Table

SAME ELEMENT • DIFFERENT ISOTOPES • AN AVERAGE MASS

- Periodic table masses are averages, not whole numbers, due to natural isotope abundance.
- Atomic number = protons; mass number = protons + neutrons in one isotope.
- Average atomic mass is the weighted average for the element.
- Carbon's standard atomic weight is about 12.011 u, since natural carbon includes carbon-13 and carbon-14.
- $H \approx 1.008 \text{ u}$, $O \approx 15.999 \text{ u}$, so $H_2O \approx 18.015 \text{ u}$.
- Classroom rounding: use two decimal places and state your precision.

Diagram illustrating the components of an element tile for Carbon (C):

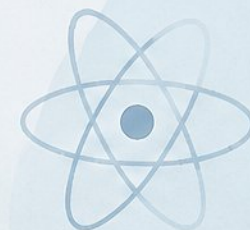
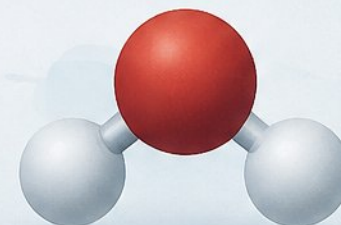
- Atomic number = protons**: 6
- Symbol and name**: C, Carbon
- Average atomic mass (in u)**: 12.011

Diagram illustrating the components of an element tile for Hydrogen (H):

- Atomic number**: 1
- Symbol and name**: H, Hydrogen
- Average atomic mass (in u)**: 1.008

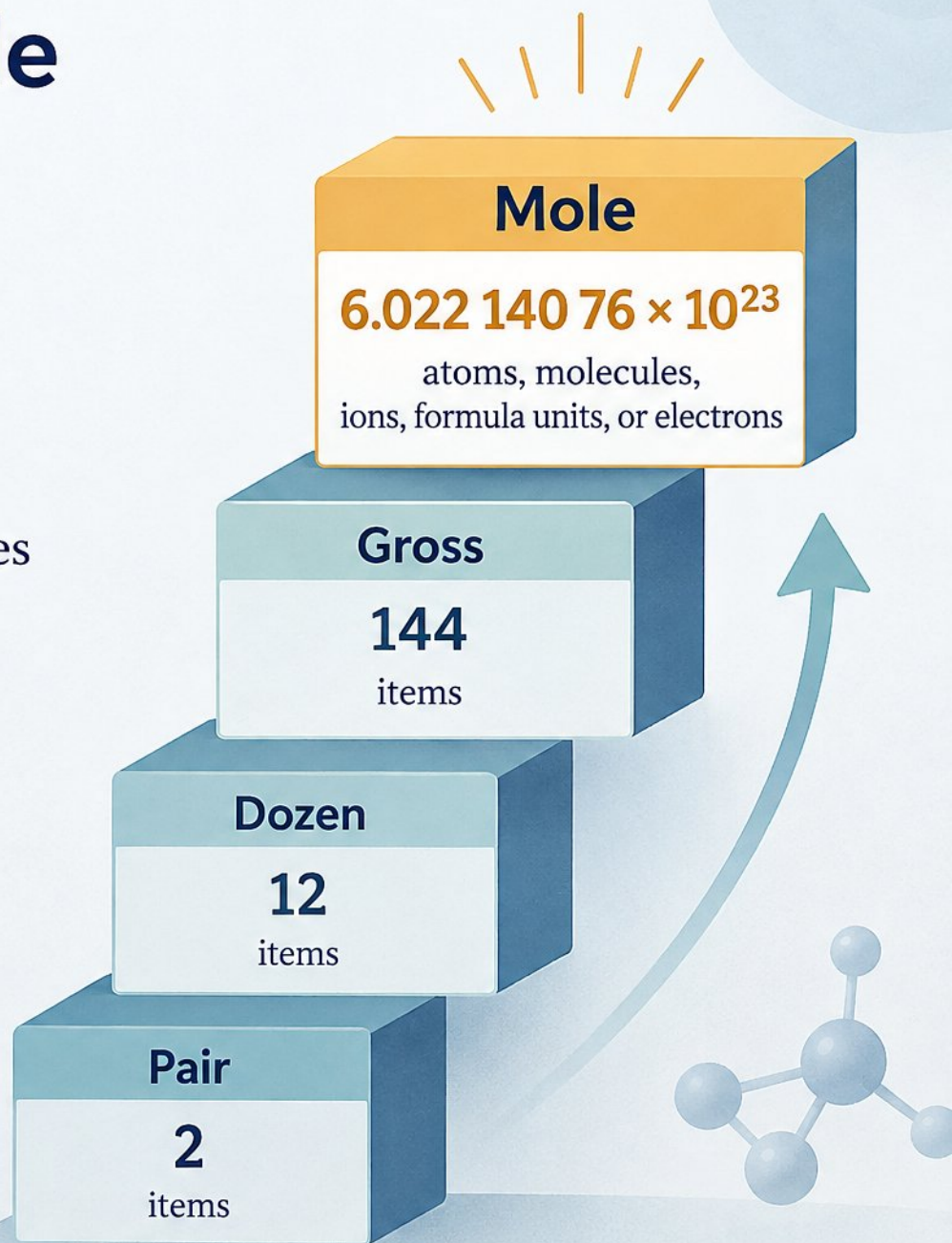
Diagram illustrating the components of an element tile for Oxygen (O):

- Atomic number**: 8
- Symbol and name**: O, Oxygen
- Average atomic mass (in u)**: 15.999



Introducing the Mole as a Counting Unit

- **Mole (mol):** SI base unit for amount of substance, symbol n
- One mole contains exactly $6.022\,140\,76 \times 10^{23}$ elementary entities
- That fixed number is the Avogadro constant's value in mol^{-1} — the Avogadro number
- Like a pair (2) or dozen (12): a mole is just a much bigger count
- Always name the entity: atoms, molecules, ions, formula units, or electrons



The SI Redefinition of the Mole (2019)

A MORE PRECISE AND UNIVERSAL DEFINITION

1971

CARBON-12 DEFINITION

Defined the mole as the amount of substance containing as many entities as there are atoms in 0.012 kg of carbon-12.



0.012 kg
of
carbon-12



2019

FIXED AVOGADRO CONSTANT

Defines the mole by fixing the Avogadro constant to an exact value.



$6.022\,140\,76 \times 10^{23}$
 mol^{-1}



- Agreed by the 26th CGPM (Nov 2018); effective 20 May 2019
- Replaced the 1971 definition tied to 0.012 kg of carbon-12
- Now fixes the Avogadro constant: exactly $6.022\,140\,76 \times 10^{23} \text{ mol}^{-1}$
- The Avogadro number now has no uncertainty
- Carbon-12's molar mass is now experimental, not exactly 12 g/mol
- Molar mass constant M_u stays $\sim 1 \text{ g/mol}$; classroom math unaffected

SAME AMOUNT OF MATTER.
A MORE PRECISE FOUNDATION.



MEASUREMENT
CONNECTS
OUR WORLD

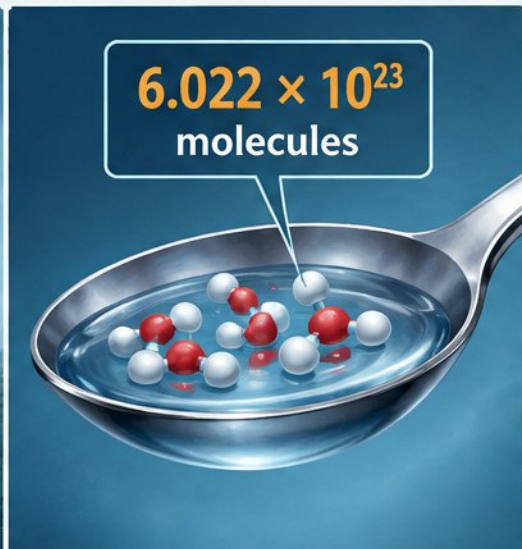
Avogadro's Number: Scale and Meaning



- 6.022×10^{23} is enormous, not a small correction factor
- A mole of rice grains would cover Earth's land kilometers deep
- Counting one per second would take about **20 trillion years**
- One mole of water: only ~18 mL, yet 6.022×10^{23} molecules
- It is a count, not a mass, and never varies by substance



A mole of rice grains
would cover Earth's
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One mole of water:
only ~18 mL, yet
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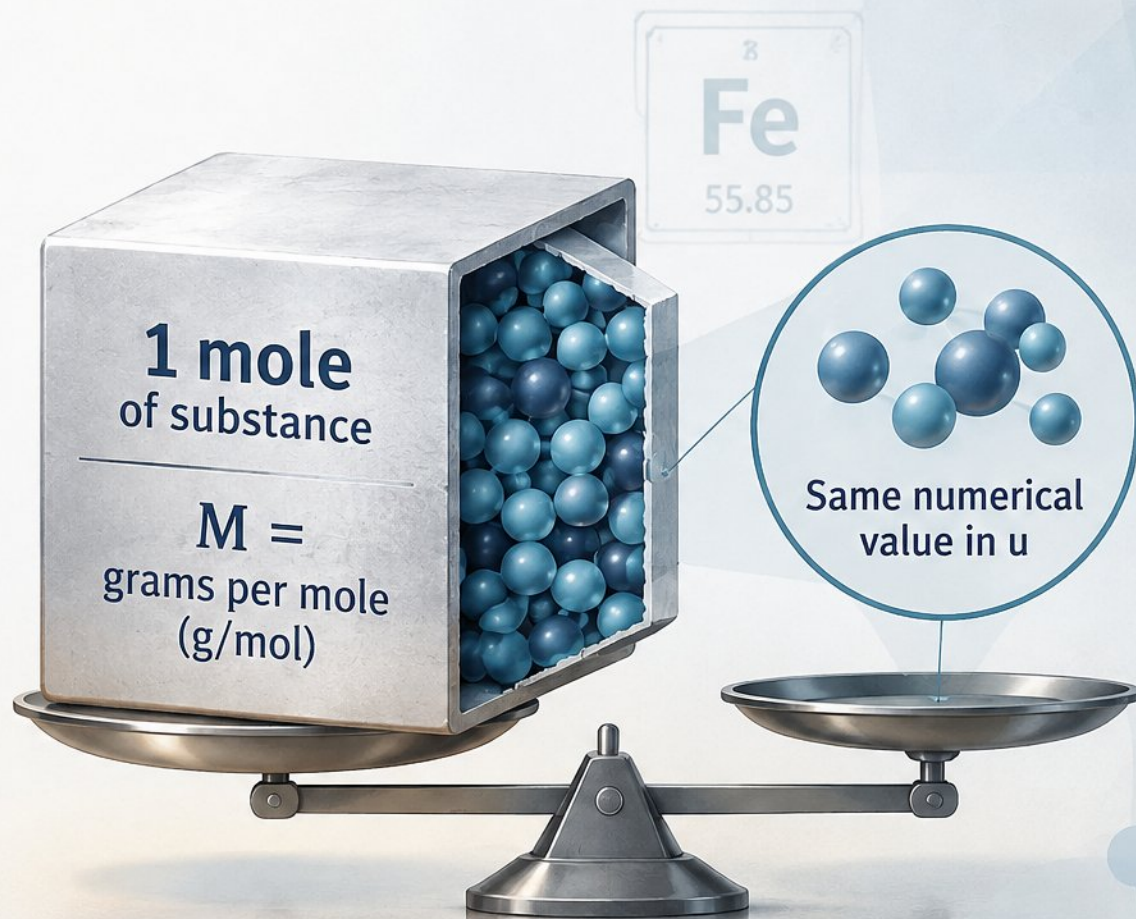


It is a count, not a mass,
and never varies
by substance

Molar Mass: The Mass of One Mole



- Molar mass M = mass in grams of one mole (g/mol)
- Numerical value in g/mol equals relative atomic, molecular, or formula mass in u
- Intensive property: does not depend on sample size
- $M = m / n$
- Elements: read from periodic table (Fe $\approx$ 55.85 g/mol)
- H₂ or O₂: multiply atomic mass by atoms per molecule
- Compounds: sum atomic masses $\times$ subscripts



Worked Example: Molar Mass of Compounds



$$2(1.008) + 1(16.00) \approx \mathbf{18.016 \text{ g/mol}}$$



$$12.01 + 2(16.00) \approx \mathbf{44.01 \text{ g/mol}}$$



$$22.99 + 35.45 \approx \mathbf{58.44 \text{ g/mol}}$$



$$40.08 + 28.02 + 96.00 \approx \mathbf{164.10 \text{ g/mol}}$$



$$53.96 + 96.21 + 192.00 \approx \mathbf{342.17 \text{ g/mol}}$$

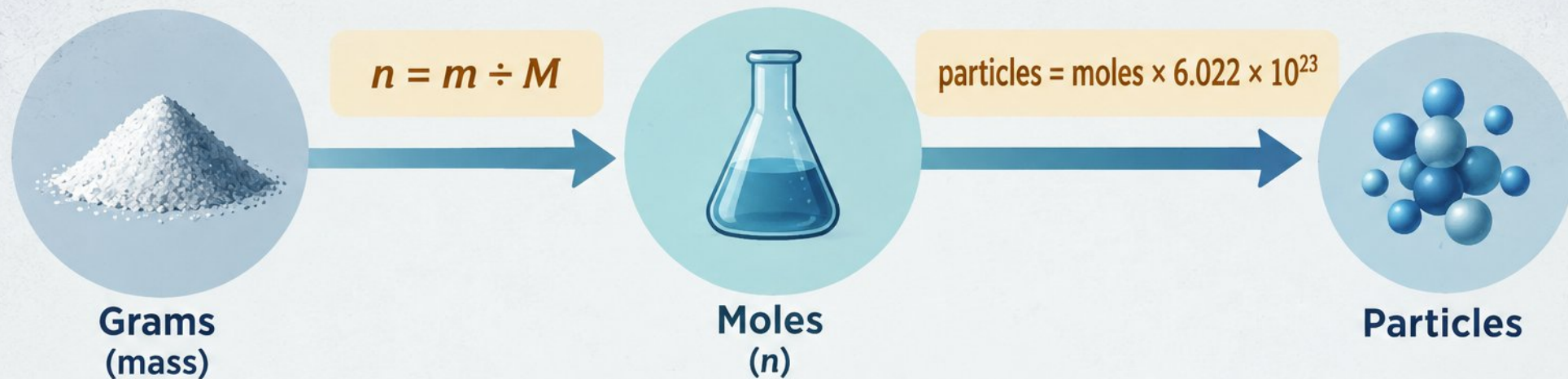


$$159.62 + 90.08 \approx \mathbf{249.70 \text{ g/mol}}$$



Core Calculations: Mass, Moles, and Particles

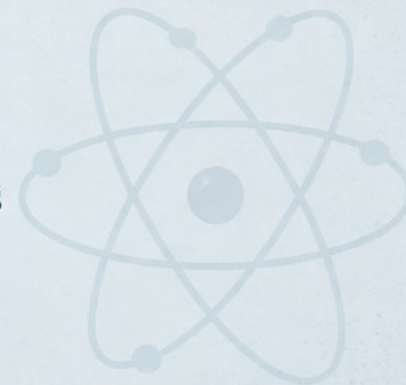
SAME SUBSTANCE. DIFFERENT UNITS. A CLEAR PATH.



Dimensional
Analysis Example

$$\frac{\text{g}}{1} \times \frac{1 \text{ mol}}{\cancel{\text{M g}}} \times \frac{6.022 \times 10^{23} \text{ particles}}{\cancel{1 \text{ mol}}} = \text{particles}$$

- Central relationship: $n = m \div M$ (moles = mass $\div$ molar mass)
- Moles to particles: $\text{particles} = \text{moles} \times 6.022 \times 10^{23}$
- Two-step route: $\text{grams} \div M \rightarrow \text{moles} \times \text{Avogadro's number} \rightarrow \text{particles}$
- Use dimensional analysis so unwanted units cancel cleanly
- Watch for wrong molar mass, inverted factors, lost sig figs



Guided Practice: Conversion Set



Use dimensional analysis to solve each conversion. Show your setup and check your units.

1 Mass to moles: $72.0 \text{ g Ar} \div 39.95 \approx 1.80 \text{ mol}$

2 Mass to moles: $108 \text{ g H}_2\text{O} \div 18.02 \approx 5.99 \text{ mol}$

3 Moles to mass: $3.00 \text{ mol CaCl}_2 \times 110.98 \approx 333 \text{ g}$

4 Moles to particles: $2.00 \text{ mol H}_2\text{O} \times 6.022 \times 10^{23} \approx 1.20 \times 10^{24} \text{ molecules}$

5 Mass to particles: $20.0 \text{ g Cl}_2 \div 70.90 \times 6.022 \times 10^{23} \approx 1.70 \times 10^{23} \text{ molecules}$

6 Reverse: $3.01 \times 10^{23} \text{ H}_2\text{O} \div 6.022 \times 10^{23} \approx 0.500 \text{ mol} \rightarrow \times 18.02 \approx 9.01 \text{ g}$

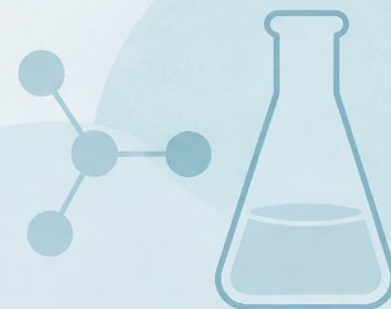


Classroom Demonstration: One Molar Display

SAME PARTICLE COUNT • DIFFERENT MASSES • DIFFERENT VOLUMES



- Goal: one mole of different substances — same particle count, different masses and volumes
- Show in identical containers: carbon 12.0 g, water 18.0 g (18.0 mL)
- Add NaCl 58.5 g, copper 63.6 g, sulfur 32.1 g, sucrose 342 g
- Extension: set out 0.100, 0.500, 1.00, 2.00, and 5.00 mol NaCl piles
- Ask learners to estimate each pile, then verify by weighing
- Bring Me a Mole: measure 1 mol aluminum foil or 0.250 mol NaCl
- Cover containers, keep water sealed against evaporation, label every sample

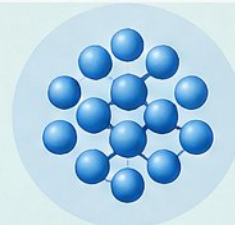


SAME AMOUNT
OF PARTICLES.
DIFFERENT PROPERTIES.

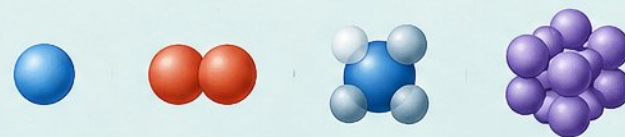
Common Misconceptions and Teaching Strategies



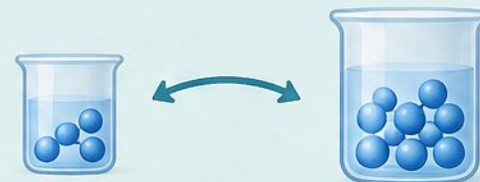
Myth: the mole is a mass — it is an amount



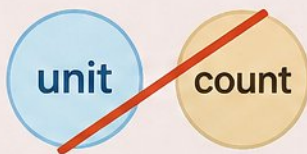
Myth: Avogadro's number applies only to atoms



Molar mass is intensive; it ignores sample size



Mole is a unit; Avogadro number is the count



Teaching Strategies

- Name the entity, label units, use particle diagrams
- Ask learners to explain reasoning, not just answer
- Connect number and mass aspects; avoid pure rote
- For adults: use concrete contexts and quick confidence checks

Assessment, Answer Keys, and Next Steps

PRACTICE | CHECK UNDERSTANDING | PLAN AHEAD



Mixed Practice

- Mixed practice: mass $\rightarrow$ moles, moles $\rightarrow$ particles, particles $\rightarrow$ mass



Mass



Moles



Particles



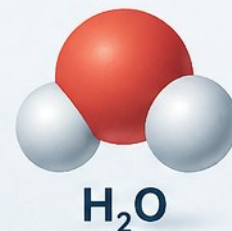
Exit Ticket

- Exit ticket: one conceptual question plus one calculation item



Sample Items

- Sample items: 75.0 g Fe $\approx$ 1.34 mol; 2.50 mol H₂O $\approx$ 1.51×10^{24} molecules



Model Answer Keys

- Model answer keys showing dimensional analysis, units, and sig figs

$$75.0 \text{ g Fe} \times \frac{1 \text{ mol Fe}}{55.85 \text{ g Fe}} = 1.34 \text{ mol Fe}$$

Units
cancel

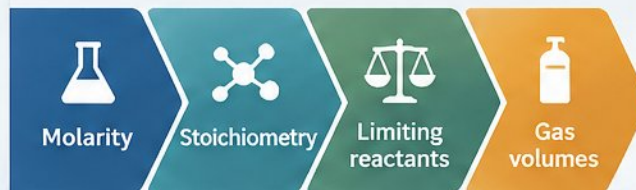
Use correct
conversion factor

Report with
correct sig figs



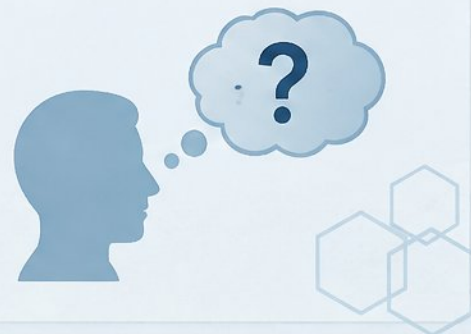
Next Steps

- Next steps: molarity, stoichiometry, limiting reactants, gas volumes



Educator Reflection

- Educator reflection: which step caused the most hesitation?



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