

Topic 1.1: Moles and Molar Mass

Revision Notes · AP Chemistry

UNIT 1 · ATOMIC STRUCTURE AND PROPERTIES

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The Mole: Counting Particles You Cannot See

Atoms and molecules are far too small to count one by one, yet chemistry constantly needs to know how many of them are reacting. The mole solves this by counting particles in bulk. One mole of any substance contains exactly 6.022×10^{23} representative particles. This fixed number is Avogadro's number, written $N_a = 6.022 \times 10^{23} \text{ mol}^{-1}$.

The mole works like any other counting unit. A dozen always means 12, whether it counts eggs or pencils; a mole always means 6.022×10^{23} , whether it counts atoms, molecules or ions. What changes is the particle being counted, and the formula or name of the substance tells you which one applies:

- An element such as Fe is counted in **atoms**.
- A molecular substance such as H₂O or CO₂ is counted in **molecules**.
- An ionic compound such as NaCl has no separate molecules, so it is counted in **formula units**, the simplest whole number ratio of its ions.
- Ions themselves can be counted directly, as in 1 mol of Na⁺ ions.

The relationship between amount and particle number is $N = n \times N_a$, where N is the number of particles and n is the amount in moles. Rearranged, $n = N / N_a$ converts a particle count back into moles. Every conversion in this topic runs through the mole, because the mole is the one quantity that connects the particle world to the measurable world of grams and litres.

Molar Mass: The Mass of One Mole

Counting particles is only useful if the count can be weighed. Molar mass is the bridge: it is the mass of one mole of a substance, expressed in grams per mole (g/mol). Its value comes straight from the periodic table. The molar mass of an element is its atomic mass read in grams, so one mole of carbon atoms has a mass of 12.01 g and one mole of iron has a mass of 55.85 g.

For a compound, add the atomic masses of every atom in the formula, each multiplied by its subscript. For water, $M(\text{H}_2\text{O}) = 2(1.008) + 16.00 = 18.02 \text{ g/mol}$. Parentheses multiply everything inside them: for calcium nitrate, $M(\text{Ca}(\text{NO}_3)_2) = 40.08 + 2(14.01) + 6(16.00) = 164.10 \text{ g/mol}$, because the subscript 2 outside the bracket doubles the nitrogen and all three oxygens of the nitrate group.

Molar mass converts between mass and moles in both directions:

$$n = m / M \quad \cdot \quad m = n \times M$$

where m is the mass in grams. Because 1 mol of any substance contains the same number of particles but a different mass, molar mass is also a sanity check. A substance with a molar mass near 100 g/mol gives about 0.01 mol per gram, never 10 mol per gram. If a calculation says otherwise, a conversion has been inverted.

Converting Between Grams, Moles and Particles

Nearly every mole calculation is a short journey with the mole at the centre:

grams ↔ moles ↔ particles

Mass converts to moles using molar mass. Moles convert to particles using Avogadro's number. There is no direct conversion between grams and particles; the route always passes through moles. The reliable method is dimensional analysis: write each conversion factor as a fraction, arrange the fractions so the unwanted units cancel, and check that the unit left standing is the one the question asked for. If the units do not cancel to the target unit, the setup is wrong before any arithmetic begins.

The same framework extends one step further. Moles of a compound convert to moles of its individual particles using the subscripts in the formula, as the next section shows. Later topics hang more destinations on the same centre: solution volume through molarity ($n = M \times V$) and gas volume through the ideal gas law ($PV = nRT$). The strategy never changes: convert what you are

given into moles, use the chemical relationship the question is really about, then convert out of moles into whatever is requested.

Counting Particles Inside a Formula

A chemical formula is a set of ratios. The subscripts state how many atoms of each element, or how many ions of each type, sit inside one formula unit or one molecule, and those ratios scale up unchanged to the mole level. One mole of H_2O contains 2 mol of H atoms and 1 mol of O atoms, 3 mol of atoms in total.

Polyatomic formulas reward careful reading. In 1.00 mol of $\text{Al}_2(\text{SO}_4)_3$ there are:

- 2.00 mol of Al^{3+} ions
- 3.00 mol of SO_4^{2-} ions
- 5.00 mol of ions in total
- 3.00 mol of S atoms
- 12.00 mol of O atoms
- 17.0 mol of atoms in total (2 Al + 3 S + 12 O per formula unit)

Questions exploit the difference between these counts. "How many oxygen atoms" and "how many sulfate ions" have different answers for the same sample, and the formula, read literally, decides both. When a number of particles is needed rather than an amount in moles, finish with Avogadro's number: moles of the specified particle $\times 6.022 \times 10^{23}$.

Worked Examples

Example 1: Particles to moles

A sample contains 3.011×10^{23} molecules of H_2O . How many moles is this?

$$n = N / N_{\text{a}} = (3.011 \times 10^{23}) / (6.022 \times 10^{23} \text{ mol}^{-1}) = 0.500 \text{ mol H}_2\text{O}$$

Half of Avogadro's number of particles is half a mole, whatever the substance. The particle type sets the label on the answer, not the arithmetic.

Example 2: Grams to formula units

How many formula units are in 25.0 g of NaCl?

First the molar mass: $M(\text{NaCl}) = 22.99 + 35.45 = 58.44 \text{ g/mol}$. Then run both steps of the pathway in one chain:

$$25.0 \text{ g NaCl} \times (1 \text{ mol NaCl} / 58.44 \text{ g NaCl}) \times (6.022 \times 10^{23} \text{ formula units} / 1 \text{ mol NaCl}) = 2.58 \times 10^{23} \text{ formula units}$$

Grams cancel in the first factor and moles cancel in the second, leaving formula units, the particle name an ionic compound requires.

Example 3: Moles to mass

What is the mass of 0.750 mol of CO_2 ?

$$M(\text{CO}_2) = 12.01 + 2(16.00) = 44.01 \text{ g/mol}$$

$$m = n \times M = 0.750 \text{ mol} \times 44.01 \text{ g/mol} = 33.0 \text{ g CO}_2$$

Example 4: Counting atoms inside a formula

How many oxygen atoms are present in 0.250 mol of $\text{Al}_2(\text{SO}_4)_3$?

Each formula unit holds $4 \times 3 = 12$ oxygen atoms, so the amount of oxygen atoms is $0.250 \text{ mol} \times 12 = 3.00 \text{ mol}$ of O atoms. Converting to a count:

$$3.00 \text{ mol} \times 6.022 \times 10^{23} \text{ atoms/mol} = 1.81 \times 10^{24} \text{ oxygen atoms}$$

Two separate ideas are chained here, the subscript ratio first, then Avogadro's number, and swapping their order is a common way to lose the question.

Traps Specific to This Topic

Moles of a compound are not moles of its atoms or ions. One mole of H_2O is 3 mol of atoms, and one mole of $\text{Al}_2(\text{SO}_4)_3$ is 5 mol of ions but 17 mol of atoms. Always name the particle before reaching for a number.

Avogadro's number converts moles to particles, never grams to particles. Any setup that multiplies a mass by 6.022×10^{23} has skipped the molar mass step, and its answer is wrong by orders of magnitude.

Parentheses multiply every atom inside them. $\text{Ca}(\text{NO}_3)_2$ contains 6 oxygen atoms, not 3, and 2 nitrogen atoms, not 1. Molar mass errors in AP questions are planted exactly here.

Read the formula exactly as written, hydrates included. $\text{CuSO}_4 \cdot 5\text{H}_2\text{O}$ carries the mass of all five water molecules, giving a molar mass of about 249.7 g/mol, not the 159.6 g/mol of the anhydrous salt alone.

Units decide the method. If a question gives grams and asks for molecules, the pathway is fixed: grams to moles to molecules. Writing the units into every step and cancelling them is not decoration; it is how you catch an inverted conversion factor before it costs the mark.

Key Relationships to Remember

- $N_a = 6.022 \times 10^{23}$ particles per mole
- $N = n \times N_a$ and $n = N / N_a$
- $M = \text{sum of (atomic mass} \times \text{subscript) over the whole formula, in g/mol}$
- $n = m / M$ and $m = n \times M$
- Subscripts give particle ratios that hold at the mole level exactly as written

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