PHYSICAL CHEMISTRY

Avogadro's Number

The count behind the mole, molar masses, particle arithmetic, and a 10-question practice set with full solutions.

SUBJECT	LEVEL	INCLUDES
Chemistry	High school and early college	Practice set, answer key, revision sheet

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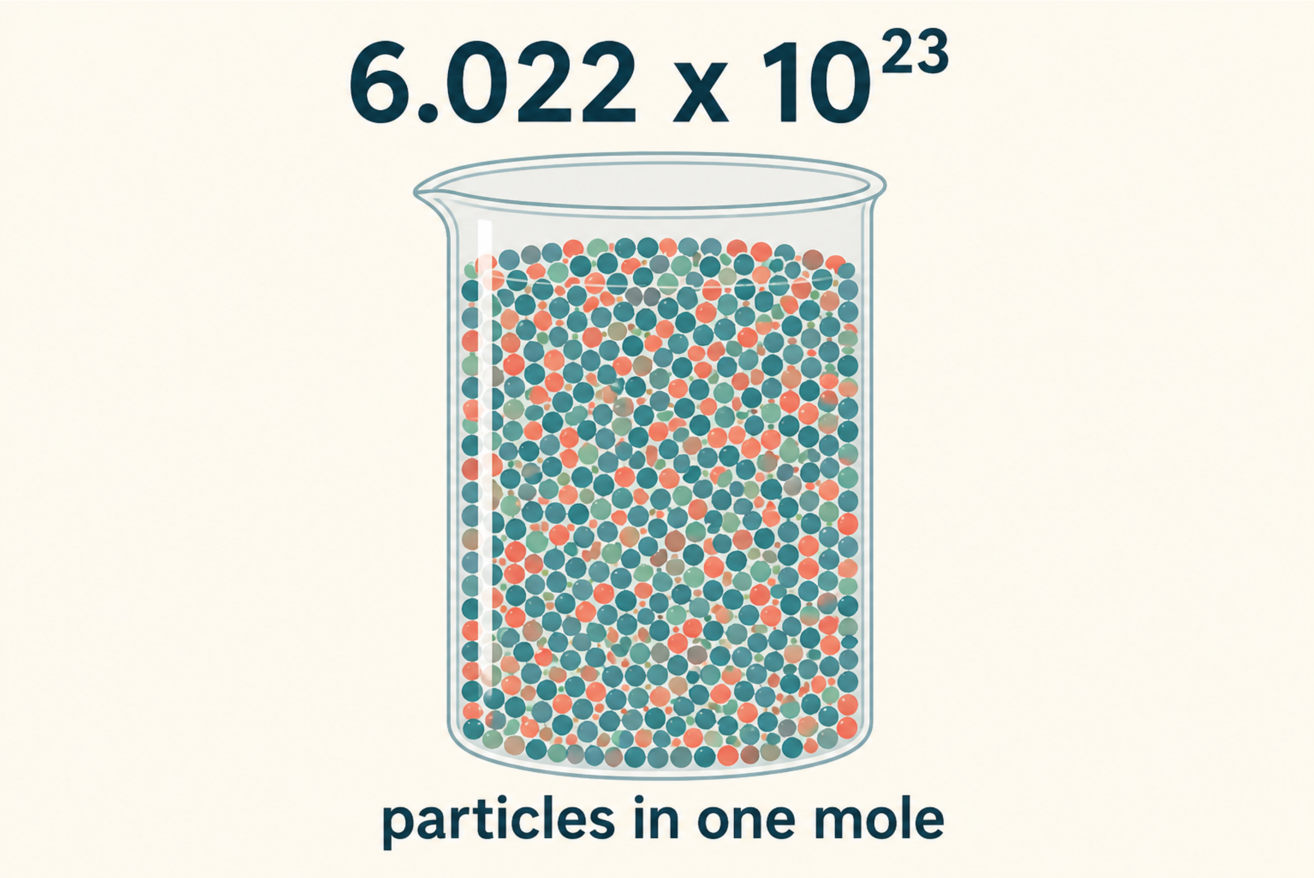
INSIDE

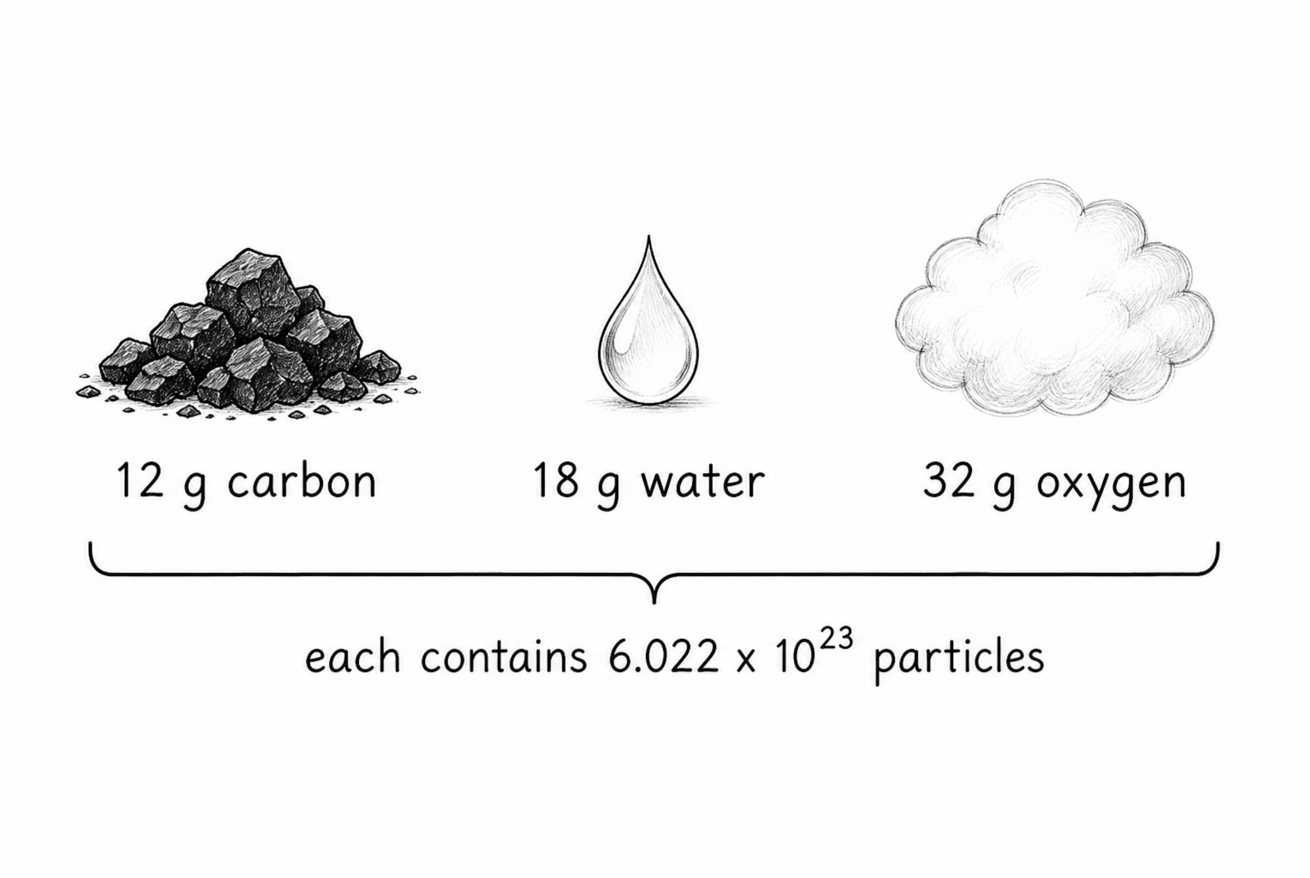
Avogadro's number is the bridge between the atomic world and the everyday one. It is the number of atoms or molecules in exactly one mole of any substance: $N_A = 6.022 \times 10^{23}$. That's about 600 sextillion. A teaspoon of water contains roughly 7×10^{22} water molecules, close to an eighth of a mole. The number is so large it almost defies intuition, but it is what makes chemistry calculations work at the scale we actually measure things, grams, liters, moles, not individual atoms.



One mole equals 6.022×10^{23} particles, Avogadro's number, the bridge between the atomic and the everyday.

The Number Itself

Avogadro's number, also called Avogadro's constant, has the precise modern value:



Different substances, wildly different masses, the same particle count: this is the definition of a mole in one picture.

$$N_A = 6.02214076 \times 10^{23}\ \text{mol}^{-1}$$

Since the 2019 redefinition of the SI base units, the value of N_A is fixed by definition, it's not measured anymore. The mole is now defined as the amount of substance containing exactly $6.02214076 \times 10^{23}$ elementary entities. Whatever you're counting (atoms, molecules, ions, electrons), N_A of them is one mole.

What a Mole Means

A mole is just a counting unit, like a dozen (12) or a gross (144), except much larger. A mole of anything is 6.022×10^{23} of that thing:

- 1 mole of carbon atoms = 6.022×10^{23} carbon atoms
- 1 mole of water molecules = 6.022×10^{23} H_2O molecules
- 1 mole of electrons = 6.022×10^{23} electrons
- 1 mole of golf balls = 6.022×10^{23} golf balls (a quantity larger than every grain of sand on Earth)

Chemists use moles instead of counting individual atoms because atoms are far too small to weigh one at a time. A mole of anything weighs a measurable, useful amount, typically a few grams to a few hundred grams. The mole connects the microscopic to the macroscopic.

Molar Mass, The Practical Connection

The molar mass of an element is the mass of one mole of its atoms, expressed in grams per mole. The number on the periodic table under each element symbol is the molar mass:

- Hydrogen (H): molar mass 1.008 g/mol → 6.022×10^{23} H atoms weigh 1.008 g
- Carbon (C): molar mass 12.011 g/mol → 6.022×10^{23} C atoms weigh 12.011 g
- Oxygen (O): molar mass 16.00 g/mol → 6.022×10^{23} O atoms weigh 16.00 g
- Iron (Fe): molar mass 55.85 g/mol → 6.022×10^{23} Fe atoms weigh 55.85 g

For compounds, add up the molar masses of all atoms in the formula. Water (H_2O) has molar mass $2 \times 1.008 + 16.00 = 18.02$ g/mol. A mole of water is 18.02 g, about a tablespoon.

Using Avogadro's Number in Calculations

The three quantities, mass, moles, and number of particles, connect through two ratios:

$$\text{moles} = \frac{\text{mass (g)}}{\text{molar mass (g/mol)}}$$

$$\text{number of particles} = \text{moles} \times N_A$$

Worked example

How many water molecules are in 36 grams of water?

1. Convert mass to moles: $36\,\text{g} \div 18.02\,\text{g/mol} \approx 2.00\,\text{mol}$.

2. Convert moles to molecules: $2.00\,\text{mol} \times 6.022 \times 10^{23}\,\text{mol}^{-1} \approx 1.20 \times 10^{24}$ molecules.

That's a lot of molecules in two tablespoons of water. The same logic works for every chemical calculation, convert to moles first, then convert moles to whatever you need.

How Avogadro's Number Was Measured

The number is named after Amedeo Avogadro, the Italian physicist who proposed in 1811 that equal volumes of gases at the same temperature and pressure contain the same number of particles. Avogadro himself never calculated the number, he died in 1856, before anyone knew how.

The first credible estimate came from Johann Josef Loschmidt in 1865, who used kinetic theory of gases and the size of molecules to estimate the number per unit volume. In German chemistry, the constant is sometimes still called the Loschmidt number.

Jean Perrin's 1908 experiments on Brownian motion of small particles in water gave a value accurate to a few percent. His work also confirmed the reality of atoms, a question that was still genuinely debated in serious scientific circles at that time. Perrin won the 1926 Nobel Prize in Physics partly for this.

Modern measurements use silicon crystal lattices: a near-perfect sphere of pure silicon-28 is measured precisely for mass, volume, and lattice spacing, and N_A is calculated from the result. The 2019 SI redefinition then fixed N_A at exactly $6.02214076 \times 10^{23}$, turning a measured quantity into a defined one.

Why the Number Is So Large

Avogadro's number is large because atoms are very small. The diameter of a typical atom is about 10^{-10} m. To stack enough atoms to make a substance you can hold in your hand and weigh on a kitchen scale, you need on the order of 10^{23} of them. The specific value comes from the conjunction of the atomic mass unit (1/12 the mass of a carbon-12 atom) and the gram (an arbitrary human-scale unit). If we had picked a different mass unit, Avogadro's number would be different, there's nothing fundamental about the specific value, just about its order of magnitude.

Related study notes: Mole Concept, Periodic Table, Stoichiometry, Chemical Bonding.

Practice Questions

Work each question before reading its solution. The set runs from direct recall and substitution to the applied questions that exams actually use to separate grades.

Q1. State Avogadro's number and what it counts.

SOLUTION

$N_A = 6.02214076 \times 10^{23}$ per mole, exact by the 2019 SI definition. It is the number of elementary entities, atoms, molecules, ions, electrons, whatever is specified, in 1 mole. The mole is to chemists what the dozen is to bakers, scaled to atomic reality.

Q2. Why does 1 mole of carbon-12 weigh about 12 g? What links the atomic mass unit to the gram?

SOLUTION

N_A was historically chosen so that 1 mole of atoms of mass 12 u weighs 12 g: the mole is the bridge $1\ \mathrm{u} \approx 1.6605 \times 10^{-24}\ \mathrm{g} = 1/N_A\ \mathrm{g}$. Atomic masses read from the periodic table in u become gram masses per mole unchanged. That numerical coincidence is the entire convenience of the unit.

Q3. How many molecules are in 36 g of water?

SOLUTION

Molar mass of $H_2O = 18$ g/mol, so 36 g is 2 mol: $2 \times 6.022 \times 10^{23} = 1.204 \times 10^{24}$ molecules. And 3 times as many atoms, 3.6×10^{24}, since each molecule carries 3.

Q4. What is the mass of a single water molecule?

SOLUTION

$m = \dfrac{18\ \mathrm{g/mol}}{6.022 \times 10^{23}\ /\mathrm{mol}} = 2.99 \times 10^{-23}$ g. Division by N_A converts any molar mass to a per-particle mass; multiplication converts back. That two-way bridge is most of stoichiometry.

Q5. How many moles, and how many atoms, are in 5.4 g of aluminium ($M = 27$ g/mol)?

SOLUTION

$n = 5.4/27 = 0.2$ mol; atoms $= 0.2 \times 6.022 \times 10^{23} = 1.2 \times 10^{23}$. The chain mass → moles → particles always runs through the molar mass first.

Q6. At STP, 1 mole of any ideal gas fills 22.4 L. How many CO_2 molecules occupy 5.6 L at STP?

SOLUTION

$n = 5.6/22.4 = 0.25$ mol, so $0.25 \times 6.022 \times 10^{23} = 1.5 \times 10^{23}$ molecules. Avogadro's own insight: equal gas volumes at equal conditions hold equal counts, regardless of the gas. The molar volume turns gas measurement into particle counting.

Q7. Give 2 comparison anchors that make the size of N_A graspable.

SOLUTION

A mole of water is 18 mL, a sip; a mole of marbles would bury the Earth kilometers deep. 6×10^{23} millimeters is about 63 million light-years, far past the Andromeda galaxy. The number is ordinary in the lab only because atoms are that small.

Q8. A drop of water (0.05 mL) evaporates. Roughly how many molecules leave, and what does this say about "trace amounts"?

SOLUTION

0.05 g is $0.05/18 \approx 2.8 \times 10^{-3}$ mol $\approx 1.7 \times 10^{21}$ molecules. Even a vanishing residue contains astronomical particle counts, which is why contamination matters in chemistry and why "chemical-free" is a marketing phrase, not a physical possibility.

Q9. Balance meets count: how many grams of oxygen react completely with 4 g of hydrogen in $2H_2 + O_2 \rightarrow 2H_2O$?

SOLUTION

4 g H_2 is 2 mol; the equation demands half as many moles of O_2, 1 mol, which is 32 g. Equations balance in moles, never in grams: the mole is what lets a mass ratio (4 g : 32 g) come from a simple count ratio (2 : 1).

Q10. How was N_A measured well enough to become an exact definition? Name 1 classical and 1 modern route.

SOLUTION

Perrin's 1908 analysis of Brownian motion in colloids gave the first solid values and made atoms respectable, earning the 1926 Nobel. The modern route counted atoms in a nearly perfect 1 kg silicon-28 sphere by X-ray crystallography: lattice spacing plus sphere volume yields the atom count to 9 digits. In 2019 the measured value was frozen into the definition.

Answer Key

QUICK ANSWERS

Q1 6.022×10^{23} entities per mole **Q2** the mole converts u to g: $M(\text{g/mol}) = m(\text{u})$ **Q3** 1.2×10^{24} **Q4** $\approx 3 \times 10^{-23}$ g **Q5** 0.2 mol; 1.2×10^{23} atoms

Q6 1.5×10^{23} **Q7** a sip of water vs an Earth-burying marble pile **Q8** $\sim 10^{21}$: traces are still astronomical counts **Q9** 32 g **Q10** Perrin's Brownian motion; the silicon sphere

Revision Sheet

The number

$N_A = 6.02214076 \times 10^{23}\ \text{mol}^{-1}$, exact since 2019.

Equations count moles

Coefficients are mole ratios. Convert masses to moles before comparing.

The bridge

moles $= \dfrac{\text{mass}}{M}$; particles $=$ moles $\times N_A$; per-particle mass $= M/N_A$.

Scale anchor

A mole of water: one sip. A mole of marbles: buries the planet.

Gas shortcut

At STP, 1 mol of ideal gas = 22.4 L: volume counts particles.

u to g

$1\ \text{u} = 1/N_A$ g: table masses in u are molar masses in g/mol.

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