ExamFitter

examfitter.com · Free AP Chemistry quizzes, revision notes and worksheets

Topic 3.5: Kinetic Molecular Theory

Revision Notes · AP Chemistry

UNIT 3 · PROPERTIES OF SUBSTANCES AND MIXTURES

On ExamFitter.com: Topic 3.5 Quiz · Topic 3.5 Worksheet · Unit 3: All Topics · AP Chemistry Home

The Model: Five Postulates

Kinetic molecular theory explains the macroscopic properties of gases through the motion of their particles. It rests on five postulates.

Particles are points. Their own volume is negligible beside the container's: nearly all of a gas is empty space.

Particles move constantly and randomly. They travel in straight lines until they collide. Pressure is what those collisions do to the walls: each impact transfers momentum, and countless impacts per second register as a steady force per unit area.

Particles exert no forces on each other. Except at the instant of collision there are no attractions or repulsions between them. With nothing pulling them together, an ideal gas could never condense.

Collisions are perfectly elastic. Total kinetic energy after a collision equals the total before it: individual particles speed up and slow down, but the total is conserved.

Average kinetic energy is set by temperature alone. It is directly proportional to the kelvin temperature, as the next section shows.

The postulates are idealizations: at high pressure the particles' own volume stops being negligible, and at low temperature the attractions the model ignores start to matter. Topic 3.6 measures those departures.

© 2026 ExamFitter.com · Free AP Chemistry revision notes

Temperature Is Average Kinetic Energy

For one mole of an ideal gas, the average kinetic energy is (3/2)RT, with R = 8.314 J·mol^{-1}·K^{-1}. Per particle it reads (3/2)kT, where k is the Boltzmann constant.

Nothing about the gas appears in the formula except through T. At 298 K the average kinetic energy is about 3.72 kJ per mole for helium, for oxygen, for xenon: identical, because the temperature is identical. Two gases at one temperature share the same average kinetic energy per particle and per mole.

Average is not total. Three moles at 300 K hold three times the total kinetic energy of one mole, while the average per particle is exactly the same in both.

Kelvin is compulsory here. The proportionality runs to absolute zero: at 0 K the classical model gives zero average kinetic energy, the limit at which translational motion reaches its minimum. Celsius has its zero in the wrong place for these ratios.

Same Energy, Different Speeds

Kinetic energy is ½mv^2, so a fixed average energy shared between a light particle and a heavy one cannot mean a shared speed: the light particle must move faster to carry the same energy on less mass.

The characteristic speed is the root-mean-square speed, $u_{rms} = \sqrt{(3RT/M)}$. M must be in kilograms per mole for the speed to come out in metres per second.

Two proportionalities follow. For one gas, u_{rms} grows with the square root of the kelvin temperature. Between gases at one temperature, u_{rms} falls with the square root of the molar mass.

At 298 K the speeds spread widely: hydrogen about 1,920 m/s, nitrogen about 515 m/s, xenon about 238 m/s. Same average kinetic energy in every case; very different speeds.

The Maxwell-Boltzmann Distribution

Plot speed horizontally and the number of particles travelling at that speed vertically and the Maxwell-Boltzmann curve appears: it rises from zero to a peak, then falls away in a long tail of fast particles.

The peak marks the most probable speed. Individual speeds scatter widely either side of it; only the distribution holds still.

© 2026 ExamFitter.com · Free AP Chemistry revision notes

The area under the whole curve counts the particles in the sample, and for a sealed sample that count cannot change: the area is fixed whatever the temperature does to the shape.

Heat the gas and the peak slides right, drops lower and broadens over a wider range of speeds. Cool it and the peak moves left, narrows and grows taller.

Swap the gas instead, at one temperature, and mass does the shifting: a heavier gas peaks further left, narrower and taller, a lighter gas further right and flatter. Xenon sits far to the left of helium, squeezed into a narrow band of slow speeds.

Effusion and Diffusion: Graham's Law

Effusion is escape through a tiny hole into a vacuum: one particle at a time slipping through a pinhole too small for bulk flow. Diffusion is the spreading of one gas through another until the mixture is uniform.

Both are driven by molecular speed, so both obey the same comparison. Graham's law: the rate of effusion is inversely proportional to the square root of the molar mass, $\text{rate}_1/\text{rate}_2 = \sqrt{(M_2/M_1)}$.

A helium balloon dies before a nitrogen one because helium atoms, effusing at $\sqrt{(28.02 \div 4.003)} \approx 2.65$ times the rate, slip through the rubber's microscopic pores far faster.

Diffusion looks slow despite those speeds: a perfume molecule takes minutes to cross a room because collisions keep knocking it off course, and its zigzag path of tiny straight runs leaves only a small net drift.

Worked Examples

Example 1: One gas at 298 K, its energy and its speed. Average kinetic energy per mole: $(3/2) \times 8.314 \times 298 = 3{,}716$ J, about 3.72 kJ; per particle, $3{,}716 \div (6.022 \times 10^{23}) \approx 6.17 \times 10^{-21}$ J. For N_2 at this temperature, $u_{rms} = \sqrt{(3 \times 8.314 \times 298 \div 0.02802)} = \sqrt{265{,}265} \approx 515$ m/s, with M converted to 0.02802 kg/mol first. Heat the nitrogen to 596 K and its speed scales by $\sqrt{(596 \div 298)} = \sqrt{2}$: $515 \times 1.41 \approx 728$ m/s. Its average kinetic energy doubled; its speed rose by only 41%.

Example 2: Helium against oxygen at 298 K. Speed ratio $= \sqrt{(32.00 \div 4.003)} \approx 2.83$. Helium's u_{rms} is about 1,363 m/s against oxygen's 482 m/s. Both carry the same average kinetic energy per mole, 3.72 kJ; helium pays for it in speed because its particles carry so little mass.

© 2026 ExamFitter.com · Free AP Chemistry revision notes

> **Example 3: Where two diffusing gases meet. CH_4 (16.04 g/mol) and SO_2 (64.07 g/mol) enter opposite ends of a tube together. Their rates stand in the ratio √(64.07 ÷ 16.04) ≈ 2.00, so methane covers twice the distance. They meet one third of the way along from the SO_2 end.**

Traps That Cost Marks in Topic 3.5

Same temperature, same speed. The equality at one temperature is kinetic energy, not speed. Offered two gases at 300 K, the lighter one is faster, every time.

Doubling the temperature doubles the speed. It doubles the average kinetic energy. Speed follows the square root: double T and speeds grow by √2 ≈ 1.41; quadruple T and they only double.

Running the ratio in Celsius. Heating from 27 °C to 327 °C looks like a twelvefold jump. In kelvin it is 300 K to 600 K: kinetic energy doubles, speeds rise by 1.41. Convert first, then ratio.

Letting the area under the curve change. Heating reshapes the curve but cannot change its area: the sample holds the same particles. A hotter curve is lower and broader precisely because the area is conserved.

Feeding grams per mole into u_{rms}. With M left as 28.02 instead of 0.02802 kg/mol, the N_2 speed at 298 K comes out near 16 m/s, about 32 times too small. The formula wants kilograms per mole when R is in joules.

Inverting Graham's law. Rate falls as mass rises, so the other gas's mass sits under the square root: $\text{rate}_1/\text{rate}_2 = \sqrt{(M_2/M_1)}$. A gas four times as heavy effuses half as fast, not a quarter as fast.

© 2026 ExamFitter.com · Free AP Chemistry revision notes

Key Relationships to Remember

- The five postulates: negligible particle volume; constant random straight line motion; no intermolecular forces except in collisions; perfectly elastic collisions; average kinetic energy proportional to kelvin temperature; and pressure as momentum transfer from collisions with the walls
- Average kinetic energy per mole = (3/2)RT and per particle = (3/2)kT: set by the kelvin temperature only, never by the gas's identity or molar mass; total kinetic energy also scales with the number of moles
- $u_{rms} = \sqrt{(3RT/M)}$, with M in kg/mol and R = 8.314 $J \cdot mol^{-1} \cdot K^{-1}$; $u_{rms} \propto \sqrt{T}$ for one gas and $\propto 1/\sqrt{M}$ between gases at one temperature
- Maxwell-Boltzmann curve: peak is the most probable speed; area under it is the total number of particles, constant for a sealed sample; heating shifts the peak right, lower and broader; a heavier gas peaks left, narrower and taller
- Graham's law: $rate_1/rate_2 = \sqrt{(M_2/M_1)}$ for effusion, with diffusion rates following the same comparison; for a fixed amount, effusion times stand in the inverse ratio, $time_1/time_2 = \sqrt{(M_1/M_2)}$

Continue Revising

Topic 3.5 Quiz: Kinetic Molecular Theory

Topic 3.5 Worksheet

Unit 3: All Topics

© 2026 ExamFitter.com · Free AP Chemistry revision notes