Topic 3.4: Ideal Gas Law

Revision Notes · AP Chemistry

UNIT 3 · PROPERTIES OF SUBSTANCES AND MIXTURES

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Four Variables, One Equation

A sample of gas is completely described by four macroscopic variables: pressure (P), the force per unit area from particles colliding with the walls; volume (V), the space the gas fills, which is the container's volume; temperature (T); and amount (n), in moles.

The ideal gas law ties them together: PV = nRT. R is the ideal gas constant, and its numerical value follows the units chosen for pressure and volume: 0.08206 $L \cdot atm \cdot mol^{-1} \cdot K^{-1}$ with litres and atmospheres, 8.314 $L \cdot kPa \cdot mol^{-1} \cdot K^{-1}$ with litres and kilopascals. Choose the R that matches the data before substituting anything.

Temperature carries one hard rule: it must be in kelvin, T(K) = T(°C) + 273.15. Kelvin is an absolute scale whose zero is a true zero of average kinetic energy, so ratios of kelvin temperatures mean something; ratios of Celsius temperatures do not.

Pressure arrives in whichever unit the apparatus reports, and the conversions are exact: 1 atm = 760 torr = 760 mmHg = 101.325 kPa. Convert the pressure into the unit your chosen R expects before substituting, not after.

Reading the Equation: Ratios Before Numbers

Most questions in this topic are ratio questions in disguise. Hold the right variables constant and the equation reads as a proportionality.

At fixed n and T, P is inversely proportional to V: halving the volume doubles the pressure. A plot of P against V is a hyperbola; a plot of P against 1/V is a straight line through the origin with slope nRT.

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At fixed n and P, V is directly proportional to T. At fixed n and V, P is directly proportional to T. Both proportionalities run on kelvin temperatures only.

When n stays fixed while conditions change, the two states obey the combined gas law, $P_1V_1/T_1 = P_2V_2/T_2$. Cancel whichever variable does not change and the simpler laws fall out.

Avogadro's principle is the same equation read once more: at the same T and P, equal volumes of any gases hold equal numbers of moles, whatever the gas. At STP (0 °C and 1 atm) one mole of an ideal gas occupies 22.4 L, so half a mole occupies 11.2 L.

Mixtures: Each Gas Pays Its Own Share

In a mixture of nonreacting gases, each gas contributes the pressure it would exert alone in the container. Dalton's law adds the shares: P_total = $P_1 + P_2 + P_3 + \ldots$

A gas's share is set by its mole fraction, X_A = n_A/n_total, through P_A = X_A × P_total. Mole fractions sum to 1, and partial pressures sum to the total.

Because a partial pressure depends only on that gas's own n, V and T, pumping more of a second gas into a rigid container leaves the first gas's partial pressure exactly where it was. Only the total climbs.

Gas collected over water is never dry: the collected volume holds the gas plus water vapour. The pressure to use in PV = nRT is the dry share, P_gas = P_total − P_{H_2O}, with the vapour pressure read at the collection temperature.

Density and Molar Mass From the Same Equation

Replace n in the ideal gas law with m/M and density appears: d = PM/RT. Density rises with pressure and with molar mass, and falls as temperature rises. Volume is never needed: it cancels in the rearrangement.

Turned around, the same relation measures molar mass: M = dRT/P. Weigh a gas, measure its density at a known pressure and temperature, and the molar mass, often enough to identify the gas, follows.

At one T and P, density ranks gases exactly as molar mass ranks them: at STP, CO_2 (44 g/mol) is denser than N_2 (28 g/mol), which is denser than He (4 g/mol).

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Worked Examples

Example 1: A straight substitution, units first. Find the volume of 0.500 mol of gas at 300 K and 1.20 atm. Pressure is in atmospheres and the answer is wanted in litres, so R = 0.08206 L·atm·mol^{-1}·K^{-1}. V = nRT/P = (0.500 × 0.08206 × 300) ÷ 1.20 = 12.309 ÷ 1.20 ≈ 10.3 L. Moles, kelvin and atmospheres cancel, leaving litres.

Example 2: The combined gas law by factors. A gas at 2.00 atm fills 6.00 L at 300 K. It is heated to 450 K and compressed to 3.00 L. Temperature scales pressure by 450 ÷ 300 = 1.5; halving the volume scales it by 6.00 ÷ 3.00 = 2. So P_2 = 2.00 × 1.5 × 2 = 6.00 atm, with no R and no moles anywhere in sight.

Example 3: A gas collected over water. Oxygen is collected over water at 25 °C under a total pressure of 760 torr, where water's vapour pressure is 24 torr, and the collected volume is 500 mL. Dry pressure: 760 − 24 = 736 torr, which is 736 ÷ 760 = 0.968 atm. Then n = PV/RT = (0.968 × 0.500) ÷ (0.08206 × 298) = 0.484 ÷ 24.45 ≈ 0.0198 mol. Skipping the vapour correction would have overstated the amount by about 3%.

Example 4: Molar mass from density. A gas has density 1.80 g/L at 1.00 atm and 300 K. M = dRT/P = (1.80 × 0.08206 × 300) ÷ 1.00 = 44.3 g/mol. That sits close to CO_2 at 44.01 g/mol, so carbon dioxide is the plausible identity.

Traps That Cost Marks in Topic 3.4

Running ratios in Celsius. A gas heated from 27 °C to 127 °C at constant pressure does not expand by 127 ÷ 27. In kelvin the temperatures are 300 K and 400 K, so the volume grows by a factor of 4/3. Convert before any ratio, every time.

Using the total pressure for a gas collected over water. The gauge reads gas plus vapour. Subtract the water vapour pressure first, or every mole count and volume built on it runs high.

Pairing the wrong R with the data. Pressure in torr with R = 0.08206 mixes two unit systems and wrecks the answer silently. Convert the pressure to atmospheres, or pick the R built for the units in front of you.

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Believing added gas squeezes the gases already present. In a rigid container each gas's partial pressure answers to its own moles alone. Adding a second gas raises the total; the first gas's share is untouched.

Giving equal volumes equal masses. Equal volumes at the same T and P hold equal moles, never equal masses. A litre of CO_2 outweighs a litre of He elevenfold, because each mole does.

Hunting for a volume in density problems. The density relation d = PM/RT has no V in it, by design. If a density question seems to be missing its volume, it is not missing anything.

Key Relationships to Remember

- PV = nRT, with T always in kelvin (T(K) = T(°C) + 273.15) and R chosen to match the pressure and volume units (0.08206 L·atm·mol^{-1}·K^{-1}; 8.314 L·kPa·mol^{-1}·K^{-1})
- At fixed n and T: P ∝ 1/V (P against V is a hyperbola; P against 1/V is a straight line through the origin, slope nRT)
- At fixed n: $P_1V_1/T_1 = P_2V_2/T_2$ (the combined gas law); at fixed P, V ∝ T; at fixed V, P ∝ T
- At the same T and P, equal volumes of gases hold equal moles; at STP one mole of an ideal gas occupies 22.4 L
- Dalton's law: P_total = $P_1 + P_2 + \ldots$; mole fraction X_A = n_A/n_total; P_A = X_A × P_total; a gas's partial pressure depends only on its own n, V and T
- Gas collected over water: P_gas = P_total − P_{H_2O}, using the vapour pressure at the collection temperature
- Density: d = PM/RT (rises with P and M, falls with T, needs no volume); molar mass: M = dRT/P

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