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Topic 3.4: Ideal Gas Law

Worksheet · AP Chemistry

UNIT 3 · PROPERTIES OF SUBSTANCES AND MIXTURES

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How to Use This Worksheet

Work through the sections in order. Section A checks the ideas, Section B builds fluency with the ideal gas equation, the combined gas law, partial pressures and density, and Section C mixes everything in exam style.

Two habits carry the whole topic: convert temperature to kelvin before anything else, and choose the value of R that matches the pressure and volume units in front of you.

Do not look at the answer key until you have attempted every question. The key gives full working, so compare your reasoning, not just your final answer.

Section A: Recall and Concepts

A1. (a) Name the four variables in PV = nRT, and state what each one measures. (b) State the rule for temperature in every gas law calculation, and give the conversion from Celsius. (c) State the value of R to use when pressure is in atmospheres and volume in litres, and the value to use when pressure is in kilopascals and volume in litres. (d) A pressure is reported as 1520 torr. Convert it to atmospheres and to kilopascals.

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A2. A fixed amount of an ideal gas is held at constant temperature. (a) Describe the shape of a plot of P against V. (b) Describe the shape of a plot of P against 1/V, and state what its slope equals. (c) State, as proportionalities, the relationships tested by Boyle's law, Charles's law and the pressure-temperature law, naming what is held constant in each. (d) At constant n and P, describe a plot of V against the kelvin temperature, state what its slope equals, and state where the line points at its lower end.

A3. Two flasks, each 5.0 L, sit at the same temperature and pressure. One holds N_2 (28.02 g/mol), the other CO_2 (44.01 g/mol). (a) Compare the numbers of moles in the two flasks, and the numbers of molecules. (b) Calculate the factor by which the mass of gas in one flask exceeds the mass in the other, and state which flask is heavier. (c) State how the densities of the two gases compare, and connect the comparison to the density equation.

A4. A rigid tank holds O_2 at a partial pressure of 2.0 atm. N_2 is pumped in at constant temperature until the total pressure reads 5.0 atm. (a) State the partial pressure of O_2 after the N_2 is added. (b) State the partial pressure of N_2 in the final mixture. (c) Explain, using PV = nRT applied to the oxygen alone, why adding the nitrogen did not change the oxygen's partial pressure.

A5. A gas is collected by bubbling it through water into an inverted bottle. (a) Explain why the collected gas is described as wet. (b) Write the equation that gives the pressure of the dry gas from the measured total pressure. (c) A student puts the total pressure straight into PV = nRT to find the moles of gas collected. State the direction of the error in the mole count, and why it arises. (d) State what the water vapour pressure in the bottle depends on, and what it does not depend on.

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Section B: Core Practice

B1. A 5.00 L vessel holds 0.250 mol of an ideal gas at 350 K. (a) Calculate the pressure in the vessel, in atmospheres. (b) The vessel is rigid and sealed. Calculate the new pressure if the gas is cooled to 175 K.

B2. A sample of gas occupies 2.50 L at 1.50 atm and 298 K. (a) Calculate the amount of gas, in moles. (b) The gas is O_2 (32.00 g/mol). Calculate the mass of the sample.

B3. A gas occupies 750 mL at 1.20 atm and 300 K. It is taken to a pressure of 0.900 atm and a temperature of 350 K, with no gas lost or added. (a) Calculate the new volume, in litres. (b) State which variable the combined gas law required to stay fixed, and which simpler law would have applied if the temperature had not changed.

B4. Use d = PM/RT for O_2 (32.00 g/mol). (a) Calculate the density of O_2 at 2.00 atm and 300 K. (b) Calculate the density at 1.00 atm and 300 K, and state the pattern your two answers show. (c) Calculate the density at 2.00 atm and 600 K, and explain the result in one sentence.

B5. A rigid container holds 3.0 mol of H_2 and 1.0 mol of N_2 at a total pressure of 6.0 atm. (a) Calculate the mole fraction of each gas. (b) Calculate the partial pressure of each gas. (c) A further 1.0 mol of N_2 is pumped in at constant volume and temperature. Calculate the new total pressure and the new partial pressure of each gas. (d) Explain why the partial pressure of H_2 behaves as it does in part (c).

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B6. Hydrogen is collected over water at 25 °C. The total pressure is 745 torr, the vapour pressure of water at 25 °C is 24 torr, and the collected volume is 500 mL. (a) Calculate the partial pressure of the dry H_2, in torr. (b) Convert that pressure to atmospheres. (c) Calculate the amount of H_2 collected, in moles. (d) Calculate the mass of H_2 collected (2.016 g/mol).

Section C: Exam Style

C1. Flask A (2.0 L) contains He at 3.0 atm. Flask B (4.0 L) contains Ar at 1.5 atm. The flasks are connected and the gases mix at constant temperature. (a) Calculate the partial pressure of He in the combined volume. (b) Calculate the partial pressure of Ar in the combined volume. (c) State the total pressure after mixing. (d) Calculate the mole fraction of He in the mixture, explaining how the given pressures and volumes tell you the mole ratio.

C2. An unknown gas is investigated. A 1.65 g sample occupies 0.900 L at 1.05 atm and 305 K. (a) Calculate the amount of gas, in moles. (b) Calculate the molar mass of the gas, and suggest a likely identity from CO_2 (44.01 g/mol), SO_2 (64.07 g/mol) and Ar (39.95 g/mol). (c) Calculate the density of the sample from the data given. Then check it against d = PM/RT using the molar mass of your suggested identity, and comment on the agreement.

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C3. A rigid 10.0 L tank holds air at 1.00 atm and 300 K. (a) Calculate the amount of air in the tank, in moles. (b) The tank is sealed and heated to 450 K. Calculate the new pressure. (c) The experiment is repeated with the tank open to the atmosphere, so the pressure stays at 1.00 atm while the tank is heated from 300 K to 450 K. Calculate the amount of air left in the tank, and the fraction of the original air that escaped. (d) The same tank, at the same pressure and temperature, is filled once with He (4.00 g/mol) and once with CO_2 (44.01 g/mol). Compare the amounts and the masses of gas in the two fillings.

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Answer Key

A1. (a) P is pressure, the force per unit area exerted by the gas on the container walls through particle collisions. V is volume, the space the gas occupies, equal to the container's volume. T is absolute temperature, the measure of the particles' average kinetic energy. n is the amount of gas, in moles. (b) Temperature must be in kelvin, because gas law proportionalities only hold on an absolute scale: T(K) = T(°C) + 273.15. (c) With atmospheres and litres, R = 0.08206 $L \cdot atm \cdot mol^{-1} \cdot K^{-1}$. With kilopascals and litres, R = 8.314 $L \cdot kPa \cdot mol^{-1} \cdot K^{-1}$. (d) 1520 torr ÷ 760 torr = 2.00 atm; 2.00 × 101.325 = 202.65 kPa, about 203 kPa.

A2. (a) A hyperbola: pressure falls as volume rises, steeply at first and then ever more gently, never reaching either axis. (b) A straight line through the origin, because P = (nRT) × (1/V) with nRT constant; the slope equals nRT. (c) Boyle's law: $P \propto 1/V$ at constant n and T. Charles's law: $V \propto T$ at constant n and P. The pressure-temperature law: $P \propto T$ at constant n and V. In the last two, T is the kelvin temperature. (d) A straight line through the origin, because V = (nR/P) × T with nR/P constant; the slope equals nR/P. Extended downward, the line reaches zero volume at 0 K, which is exactly why the kelvin scale, with its true zero, is the scale gas laws are written in: a Celsius plot would cut the axis at -273.15 °C instead and hide the proportionality.

A3. (a) The moles are equal, and so are the numbers of molecules: at the same T and P, equal volumes of gases hold equal moles, whatever the gas (Avogadro's principle). At STP each 5.0 L flask would hold the same count; the identity of the gas is irrelevant. (b) Mass = moles × molar mass, and the moles are equal, so the masses stand in the molar mass ratio: 44.01 ÷ 28.02 ≈ 1.57. The CO_2 flask holds about 1.57 times the mass of the N_2 flask. (c) Density is mass per volume and the volumes are equal, so the densities stand in the same ratio, CO_2 about 1.57 times denser. The density equation says the same thing: at fixed P and T, d = PM/RT is proportional to M alone.

A4. (a) 2.0 atm, unchanged. (b) Partial pressures add to the total, so $P_{N_2} = 5.0 - 2.0 = 3.0$ atm. (c) For the oxygen alone, $P_{O_2} = n_{O_2}RT/V$. Pumping in nitrogen changes neither the moles of O_2, nor the tank's volume, nor the temperature, so every quantity on the right hand side of that equation is what it was, and P_{O_2} cannot change. The nitrogen contributes its own separate share on top.

A5. (a) The gas bubbles through water and collects in contact with it, so water evaporates into the collected volume until its vapour is saturated; the bottle holds the gas mixed with water vapour, not the gas alone. (b) $P_{gas} = P_{total} - P_{H_2O}$. (c) The mole count comes out too large. The total pressure counts both the gas and the vapour pushing together, so crediting all of it to the collected gas overstates its pressure, and n = PV/RT converts that overstated pressure into overstated moles. (d) The vapour pressure depends on the temperature alone (and on the liquid being water). It does not depend on the volume collected, the amount of the other gas, or the total pressure.

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B1. (a) P = nRT/V = (0.250 × 0.08206 × 350) ÷ 5.00 = 7.18 ÷ 5.00 ≈ 1.44 atm. (b) At fixed n and V, P ∝ T. Cooling from 350 K to 175 K halves the kelvin temperature, so the pressure halves: 1.44 ÷ 2 = 0.718 atm.

B2. (a) n = PV/RT = (1.50 × 2.50) ÷ (0.08206 × 298) = 3.75 ÷ 24.45 ≈ 0.153 mol. (b) Mass = 0.153 × 32.00 ≈ 4.91 g.

B3. (a) Combined gas law: $V_2 = P_1V_1T_2 \div (T_1P_2)$ = (1.20 × 750 × 350) ÷ (300 × 0.900) = 315,000 ÷ 270 ≈ 1,167 mL, which is about 1.17 L. Both changes push the volume up: the pressure fall multiplies it by 1.20 ÷ 0.900, and the kelvin temperature rise multiplies it by 350 ÷ 300, so 750 × (1.20 ÷ 0.900) × (350 ÷ 300) ≈ 1,167 mL by the factor route too. (b) The amount of gas, n, had to stay fixed; the equation is just PV = nRT written for two states with the same nR. Had the temperature not changed, the pressure and volume terms alone would remain, which is Boyle's law.

B4. (a) d = PM/RT = (2.00 × 32.00) ÷ (0.08206 × 300) = 64.0 ÷ 24.62 ≈ 2.60 g/L. (b) d = (1.00 × 32.00) ÷ 24.62 ≈ 1.30 g/L. Halving the pressure halved the density: at fixed T, density is directly proportional to pressure. (c) d = (2.00 × 32.00) ÷ (0.08206 × 600) = 64.0 ÷ 49.24 ≈ 1.30 g/L. Doubling the kelvin temperature exactly undid the doubling of pressure, because density follows the ratio P/T.

B5. (a) Total moles = 4.0. X_{H_2} = 3.0 ÷ 4.0 = 0.75; X_{N_2} = 1.0 ÷ 4.0 = 0.25. (b) P_{H_2} = 0.75 × 6.0 = 4.5 atm; P_{N_2} = 0.25 × 6.0 = 1.5 atm. Check: 4.5 + 1.5 = 6.0 atm. (c) Total moles are now 5.0, and at fixed V and T the total pressure scales with total moles: 6.0 × (5.0 ÷ 4.0) = 7.5 atm. P_{H_2} stays 4.5 atm, so P_{N_2} = 7.5 − 4.5 = 3.0 atm. (d) The hydrogen's partial pressure is fixed by its own moles, the volume and the temperature, none of which changed; the added nitrogen raised only the total and its own share.

B6. (a) P_{H_2} = 745 − 24 = 721 torr. (b) 721 ÷ 760 = 0.949 atm. (c) n = PV/RT = (0.949 × 0.500) ÷ (0.08206 × 298) = 0.474 ÷ 24.45 ≈ 0.0194 mol. (d) Mass = 0.0194 × 2.016 ≈ 0.0391 g.

C1. (a) Once connected, the helium spreads through 2.0 + 4.0 = 6.0 L. By Boyle's law its pressure falls by the volume ratio: P_{He} = 3.0 × (2.0 ÷ 6.0) = 1.0 atm. (b) The argon spreads the other way: P_{Ar} = 1.5 × (4.0 ÷ 6.0) = 1.0 atm. (c) Total pressure = 1.0 + 1.0 = 2.0 atm. (d) At the common temperature, each gas's mole count is proportional to its original PV product: He contributes 3.0 × 2.0 = 6.0, Ar contributes 1.5 × 4.0 = 6.0. Equal products mean equal moles, so X_{He} = 6.0 ÷ 12.0 = 0.50, which matches the equal partial pressures in the combined volume.

C2. (a) n = PV/RT = (1.05 × 0.900) ÷ (0.08206 × 305) = 0.945 ÷ 25.03 ≈ 0.0378 mol. (b) M = mass ÷ moles = 1.65 ÷ 0.0378 ≈ 43.7 g/mol. Of the three candidates, CO_2 at 44.01 g/mol is by far the closest; Ar (39.95) and SO_2 (64.07) are both well outside experimental error of 43.7. (c) Density from the data: d = 1.65 ÷ 0.900 ≈ 1.83 g/L. Check with M = 44.01: d = PM/RT = (1.05 × 44.01) ÷ 25.03 ≈ 1.85 g/L. The two values agree to within about 1%, the small gap being the same experimental error that put the measured molar mass at 43.7 rather than 44.01, so the identification is consistent.

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C3. (a) n = PV/RT = (1.00 × 10.0) ÷ (0.08206 × 300) = 10.0 ÷ 24.62 ≈ 0.406 mol. (b) Sealed means n and V fixed, so P ∝ T: P_2 = 1.00 × (450 ÷ 300) = 1.50 atm. (c) Open means P and V fixed, so n ∝ 1/T: n_2 = 0.406 × (300 ÷ 450) = 0.271 mol. The escaped amount is 0.406 − 0.271 = 0.135 mol, which is 0.135 ÷ 0.406 = one third of the original air; heating drove it out as the remaining air expanded. (d) The amounts are identical: at the same P, V and T the mole count is fixed by PV = nRT alone, 0.406 mol either way. The masses differ by the molar mass ratio: CO_2's mass is 44.01 ÷ 4.00 ≈ 11.0 times the helium mass, about 17.9 g against about 1.62 g if figures are wanted.

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