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Topic 3.5: Kinetic Molecular Theory

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 postulates and the reading of Maxwell-Boltzmann curves, Section B builds fluency with average kinetic energy, root-mean-square speeds and Graham's law, and Section C mixes everything in exam style.

Two habits carry the whole topic: temperature goes into every formula in kelvin, and molar mass goes into the speed formula in kilograms per mole.

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) State the five postulates of the kinetic molecular theory of an ideal gas. (b) Which postulate explains why a gas exerts pressure on the walls of its container? (c) Which postulate explains why an ideal gas can never condense into a liquid, however cold it gets? (d) Two particles collide and one leaves the collision moving faster than before. State what must have happened to the other particle, and name the postulate your answer rests on.

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A2. (a) State what the average kinetic energy of gas particles depends on, and two things it does not depend on. (b) Samples of He and Xe are both held at 400 K. Compare their average kinetic energies per mole and per particle. (c) One flask holds 1 mol of N_2 at 300 K and another holds 3 mol of N_2 at 300 K. Compare the average kinetic energy per particle in the two flasks, and compare their total kinetic energies. (d) Explain why these comparisons would collapse if the temperatures were quoted and ratioed in degrees Celsius.

A3. A Maxwell-Boltzmann curve plots molecular speed horizontally. (a) State what the vertical axis shows. (b) State what the total area under the curve represents. (c) State what the peak of the curve marks. (d) The same sample is heated and a second curve is drawn on the same axes. State two ways the second curve differs from the first and one way it does not.

A4. (a) Define effusion and diffusion, making the difference between them clear. (b) A helium filled balloon and a nitrogen filled balloon are left in the same room. The helium balloon shrinks noticeably faster. Name the law that explains this and the particle property that decides it. (c) Perfume molecules move at hundreds of metres per second, yet the scent takes minutes to cross a room. Explain the delay at particle level. (d) State whether diffusion rates for two gases follow the same square root of molar mass comparison as effusion rates, and why.

A5. (a) State what 0 K means for particle motion and average kinetic energy in the classical kinetic model, and why it is called absolute zero. (b) A gas is heated from 27 °C to 327 °C. Convert both temperatures to kelvin. (c) Using your kelvin values, state the factor by which the average kinetic energy changes. (d) A student answers part (c) with 327 ÷ 27 ≈ 12. Explain the error in one sentence.

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Section B: Core Practice

B1. Use average kinetic energy per mole = (3/2)RT, with R = 8.314 $J \cdot mol^{-1} \cdot K^{-1}$. (a) Calculate the average kinetic energy per mole of a gas at 300 K, in joules and in kilojoules. (b) Calculate the average kinetic energy per particle at 300 K (Avogadro constant 6.022×10^{23} mol^{-1}). (c) Calculate the average kinetic energy per mole at 600 K, and state the factor by which it has grown from part (a).

B2. Use $u_{rms} = \sqrt{(3RT/M)}$ for O_2 (M = 0.03200 kg/mol). (a) Calculate u_{rms} at 300 K. (b) Calculate u_{rms} at 150 K, showing the square root factor you use. (c) State what happens to the average kinetic energy per mole between 300 K and 150 K, and reconcile that with your answers to (a) and (b).

B3. H_2 (2.016 g/mol) and CO_2 (44.01 g/mol) are held at 298 K in separate vessels. (a) Calculate the ratio $u_{rms}(H_2) \div u_{rms}(CO_2)$ from the molar masses alone. (b) Given $u_{rms}(H_2)$ = 1,920 m/s at 298 K, calculate $u_{rms}(CO_2)$. (c) Explain, using $KE = \frac{1}{2}mv^2$, why the ratio in (a) is a square root and not the plain mass ratio of about 22.

B4. A fixed amount of He effuses through a pinhole in 30 s under fixed conditions. (a) State the ratio of the effusion rates of He and N_2 (28.02 g/mol; He 4.003 g/mol). (b) Calculate the time the same amount of N_2 takes to effuse through the same pinhole under the same conditions. (c) Explain why the time ratio in (b) is the rate ratio of part (a) turned upside down.

B5. An unknown gas takes 60 s for a fixed amount to effuse through a pinhole. The same amount of CH_4 (16.04 g/mol) takes 30 s under identical conditions. (a) State the ratio of the effusion rate of the unknown gas to that of CH_4. (b) Calculate the molar mass of the unknown gas. (c) Suggest an identity for the unknown gas from SO_2 (64.07 g/mol), CO_2 (44.01 g/mol) and Cl_2 (70.90 g/mol).

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B6. Maxwell-Boltzmann curves are drawn for the same sealed sample of O_2 at 300 K and at 1,200 K. (a) State which curve's peak sits further to the right. (b) Compare the areas under the two curves. (c) Compare the heights of the two peaks, and connect your answer to part (b). (d) State the factor by which the average kinetic energy changes from 300 K to 1,200 K, and the factor by which u_{rms} changes.

Section C: Exam Style

C1. A glass tube 1.00 m long is fitted at one end with cotton wool soaked in concentrated ammonia solution (NH_3, 17.03 g/mol) and at the other end with cotton wool soaked in concentrated hydrochloric acid (HCl, 36.46 g/mol). The gases diffuse along the tube and react where they meet, forming a white ring of ammonium chloride. (a) Calculate the ratio of the diffusion rates, rate(NH_3) ÷ rate(HCl). (b) Calculate the distance of the white ring from the HCl end of the tube. (c) Explain why the ring does not form at the midpoint. (d) The whole tube is warmed to a higher uniform temperature. State what happens to each diffusion rate, to the position of the ring, and to the time the ring takes to appear.

C2. Two rigid containers have the same volume. Container X holds 1.0 mol of He at 300 K. Container Y holds 1.0 mol of Ar (39.95 g/mol) at 600 K. (a) Compare the average kinetic energy per mole in the two containers, giving both values. (b) Calculate u_{rms} for the gas in each container (He 4.003 g/mol) and state which gas's particles move faster on average. (c) Compare the pressures in the two containers, using the particle model and the ideal gas equation. (d) Explain how the container with the higher average kinetic energy can hold the slower particles, using $KE = \frac{1}{2}mv^2$.

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C3. An unknown noble gas has u_{rms} = 300 m/s at 300 K. (a) Calculate the molar mass of the gas, in kg/mol and in g/mol. (b) Suggest its identity from Kr (83.80 g/mol), Xe (131.29 g/mol) and Rn (222 g/mol). (c) Calculate u_{rms} of the same gas at 1,200 K. (d) State the average kinetic energy per mole of this gas at 300 K, and explain why the answer would be identical for either of the other two candidates at the same temperature.

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Answer Key

A1. (a) The particles' own volume is negligible beside the container's volume; the particles move constantly, randomly, in straight lines between collisions; the particles exert no attractions or repulsions on each other except during collisions; all collisions are perfectly elastic, so total kinetic energy is conserved; the average kinetic energy of the particles is directly proportional to the kelvin temperature. (b) The constant random motion postulate: moving particles collide with the walls, each collision transfers momentum, and the accumulated impacts are the pressure. (c) The no intermolecular forces postulate: condensation needs attractions to pull particles into a liquid, and the ideal model sets those attractions at exactly zero. (d) The other particle must leave more slowly, carrying less kinetic energy, so that the pair's total is unchanged; that is the perfectly elastic collision postulate. One particle can gain only what the other loses.

A2. (a) It depends on the absolute (kelvin) temperature alone. It does not depend on the identity of the gas, its molar mass, its pressure, or how much gas is present. (b) They are equal, per mole and per particle: at the same temperature every ideal gas has the same average kinetic energy, and 400 K fixes that value for helium and xenon alike, even though a xenon atom weighs about 33 times a helium atom. (c) The average per particle is identical, since both flasks sit at 300 K. The totals differ: the 3 mol flask holds three times as many particles each carrying the same average, so its total kinetic energy is three times the 1 mol flask's. (d) Celsius zero is not a true zero of kinetic energy, so Celsius ratios are meaningless for energy: 27 °C to 327 °C is a ratio of about 12 in Celsius but only 600 K to 300 K, a factor of 2, on the scale the proportionality actually uses.

A3. (a) The number of particles (the fraction of particles) travelling at each speed. (b) The total number of particles in the sample. (c) The most probable speed, the speed at which the largest number of particles travel. (d) On heating, the peak moves to the right, to higher speeds, and the curve becomes lower and broader. What does not change is the area under the curve: the sample is sealed, so the particle count, and therefore the area, is the same.

A4. (a) Effusion is the escape of gas particles through a tiny hole into a vacuum, one particle at a time through an opening too small for bulk flow. Diffusion is the gradual spreading and mixing of one gas through another (or through space) driven by random motion, with no hole involved. (b) Graham's law. The deciding property is molar mass: helium atoms (4.003 g/mol) are far lighter than nitrogen molecules (28.02 g/mol), so at the same temperature they move faster and slip through the balloon's microscopic pores at $\sqrt{(28.02 \div 4.003)} \approx 2.65$ times the rate. (c) The molecules are continually colliding with air molecules, and each collision knocks them onto a new straight run in a new direction. The path is a zigzag of very short runs, so the net displacement grows far more slowly than the instantaneous speed suggests. (d) Yes. Diffusion is driven by the same molecular speeds that set effusion rates, so the lighter gas diffuses faster by the same $\sqrt{(M_2/}$

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M_1) comparison; only the setting differs, mixing through a gas rather than escaping through a hole.

A5. (a) In the classical model it is the temperature at which average translational kinetic energy reaches zero and particle motion reaches its minimum: the floor of the temperature scale, the zero from which the kelvin scale is built, rather than an arbitrary reference point like the freezing of water. (b) 27 °C = 300.15 K, about 300 K; 327 °C = 600.15 K, about 600 K. (c) Average kinetic energy is proportional to the kelvin temperature, so it changes by 600.15 ÷ 300.15 ≈ 2.00: it doubles. (d) The student ratioed Celsius readings, whose zero is not the zero of kinetic energy; only kelvin ratios carry the proportionality.

B1. (a) (3/2) × 8.314 × 300 = 3,741 J per mole, about 3.74 kJ/mol. (b) 3,741 ÷ (6.022×10^{23}) ≈ 6.21×10^{-21} J per particle. (c) (3/2) × 8.314 × 600 = 7,483 J per mole, about 7.48 kJ/mol. That is exactly twice part (a), as it must be: the temperature doubled and average kinetic energy is directly proportional to it.

B2. (a) $u_{rms} = \sqrt{(3 \times 8.314 \times 300 \div 0.03200)} = \sqrt{233{,}831} \approx 484$ m/s. (b) Speed scales with the square root of temperature: $\sqrt{(150 \div 300)} = \sqrt{0.5} \approx 0.707$, so $u_{rms} = 484 \times 0.707 \approx 342$ m/s (direct substitution gives the same 342 m/s). (c) The average kinetic energy per mole halves, from 3.74 kJ to 1.87 kJ, because it is directly proportional to T. There is no contradiction: energy follows T itself, speed follows √T, and $\sqrt{(1/2)} \approx 0.707$ is exactly the factor that halves $\frac{1}{2}mv^2$.

B3. (a) rate and speed ratios run as $\sqrt{(M_2/M_1)}$: $\sqrt{(44.01 \div 2.016)} = \sqrt{21.83} \approx 4.67$. Hydrogen's root-mean-square speed is about 4.67 times carbon dioxide's. (b) $u_{rms}(CO_2) = 1{,}920 \div 4.67 \approx 411$ m/s. (c) At one temperature the average $\frac{1}{2}mv^2$ is the same for both gases, so v^2 must stand in the inverse mass ratio, and v therefore stands in the inverse square root of the mass ratio. The mass ratio is about 22, but speeds answer to $\sqrt{22} \approx 4.7$, not to 22 itself.

B4. (a) rate(He) ÷ rate(N_2) = $\sqrt{(28.02 \div 4.003)} = \sqrt{7.00} \approx 2.65$. Helium effuses about 2.65 times as fast. (b) For a fixed amount, time is inversely proportional to rate, so time(N_2) = 30 × 2.65 ≈ 79 s. (c) Rate measures amount per unit time, so a gas that effuses at 1/2.65 of the rate needs 2.65 times as long to pass the same amount: the time ratio is the rate ratio inverted, $\sqrt{(M_1/M_2)}$ instead of $\sqrt{(M_2/M_1)}$.

B5. (a) The unknown passes the same amount in twice the time, so its rate is half of methane's: rate(unknown) ÷ rate(CH_4) = 30 ÷ 60 = 0.500. (b) Graham's law: $0.500 = \sqrt{(16.04 \div M)}$, so $0.250 = 16.04 \div M$ and $M = 16.04 \div 0.250 = 64.16$, about 64.2 g/mol. Equivalently, halving the rate means quadrupling the mass: 16.04 × 4 = 64.16 g/mol. (c) SO_2 at 64.07 g/mol matches the calculated 64.2 g/mol within rounding; CO_2 (44.01) and Cl_2 (70.90) are both well away from it.

B6. (a) The 1,200 K curve's peak sits further right: its particles hold four times the average kinetic energy and twice the root-mean-square speed. (b) The areas are equal. The sample is sealed, the particle count is unchanged, and area under the curve is that count. (c) The 1,200 K peak is lower. The same area spread over a wider range of speeds cannot stand as tall; the 300 K curve,

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compressed into a narrower band, peaks higher. (d) Average kinetic energy scales with T: 1,200 ÷ 300 = 4, so it grows by a factor of 4. Root-mean-square speed scales with √T: √4 = 2, so it doubles.

C1. (a) rate(NH_3) ÷ rate(HCl) = √(36.46 ÷ 17.03) = √2.141 ≈ 1.46. Ammonia diffuses about 1.46 times as fast as hydrogen chloride. (b) In any given time the gases travel distances in the rate ratio, 1.46 to 1, sharing the 1.00 m between them. HCl's share is 1.00 × (1 ÷ 2.46) ≈ 0.41 m, so the ring forms about 0.41 m from the HCl end (about 0.59 m from the NH_3 end). (c) The midpoint would require equal distances in equal times, which means equal rates. Hydrogen chloride molecules are heavier (36.46 against 17.03 g/mol), move more slowly at the same temperature, and so cover less of the tube before the two fronts meet; the meeting point is dragged towards the HCl end. (d) Both gases' speeds rise by the same √T factor, so both rates increase in the same proportion, the ring forms sooner, but the ratio of rates is unchanged and the ring forms in exactly the same position, about 0.41 m from the HCl end.

C2. (a) Average kinetic energy per mole = (3/2)RT: in X, (3/2) × 8.314 × 300 = 3,741 J, about 3.74 kJ/mol; in Y, (3/2) × 8.314 × 600 = 7,483 J, about 7.48 kJ/mol. The argon holds twice the average kinetic energy per mole, because its temperature is twice as high, and the gas's identity plays no part. (b) u_{rms}(He at 300 K) = √(3 × 8.314 × 300 ÷ 0.004003) = √1,869,248 ≈ 1,367 m/s. u_{rms}(Ar at 600 K) = √(3 × 8.314 × 600 ÷ 0.03995) = √374,598 ≈ 612 m/s. The helium particles move faster on average, by a factor of about 2.2. (c) The containers have the same volume and the same number of moles, so by PV = nRT the pressure follows the temperature: Y's pressure is twice X's. In particle terms, argon's collisions strike the walls with twice the average energy, and that outweighs helium's more frequent, weaker impacts. (d) Average kinetic energy is ½mv^2, a product of mass and speed squared. An argon atom carries about ten times the mass of a helium atom, so it can hold twice the kinetic energy while moving at well under half the speed: the doubled temperature raises the energy each particle carries, but the heavy particle banks that energy in its mass rather than in speed.

C3. (a) Rearranged, M = 3RT ÷ u_{rms}^2 = (3 × 8.314 × 300) ÷ 300^2 = 7,483 ÷ 90,000 = 0.0831 kg/mol, which is 83.1 g/mol. (b) Kr at 83.80 g/mol is the close match; Xe (131.29) and Rn (222) are far too heavy to give so high a speed at 300 K. (c) Speed scales with √T: √(1,200 ÷ 300) = 2, so u_{rms} = 300 × 2 = 600 m/s. (d) (3/2) × 8.314 × 300 = 3,741 J, about 3.74 kJ per mole. Average kinetic energy is set by the kelvin temperature alone: at 300 K every ideal gas, whatever its particles weigh, has the same average per mole. The three candidates differ in speed, not in average energy.

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