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Ideal Gas Laws

Study Guide — States of Matter & Gases

Pre-med/IB-style questions building mastery of gas behavior: Boyle/Charles/Gay-Lussac/Avogadro laws, PV=nRT, unit traps, STP and molar volume, Dalton's law and partial pressures (including gas over water), gas stoichiometry with volumes, density and molar mass from gas data, diffusion/effusion, and when real gases deviate from ideal behavior.

70 items — Study Guide with Answers

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Generated February 20, 2026

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1 **Which pair of assumptions is made in the ideal gas model?**

- A Gas particles are stationary; gas particles have very large volumes.
- B **Gas particles have negligible volume; there are no intermolecular forces between particles.** ✓
- C Gas particles attract strongly; collisions are inelastic.
- D Gas particles have fixed positions; pressure is due to attraction to container walls.
- E Gas particles have variable molar mass; temperature depends on pressure only.

▶ **Explanation:** An ideal gas is modeled as point particles (negligible volume) with no intermolecular attractions/repulsions; collisions are elastic and pressure comes from collisions with container walls.

2 **In gas law calculations (e.g., PV = nRT), temperature must be expressed in:**

- A °C (Celsius)
- B °F (Fahrenheit)
- C **K (Kelvin)** ✓
- D atm (atmospheres)
- E mol (moles)

▶ **Explanation:** Gas laws depend on absolute temperature. Kelvin is an absolute scale (0 K = absolute zero), so it must be used in proportional relationships like PV = nRT.

3 **What is 27°C in Kelvin (K)?**

- A 246 K
- B 273 K

C **300 K** ✓

D 327 K

E 400 K

▶ **Explanation:** K = °C + 273 (approximately). So 27 + 273 = 300 K.

4 **Boyle's law describes the relationship between pressure and volume for a fixed amount of gas at constant temperature:**

A P is directly proportional to V

B **P is inversely proportional to V** ✓

C P is independent of V

D P is proportional to V^2

E P is proportional to $1/V^2$

▶ **Explanation:** Boyle's law: PV = constant (at constant T and n). So if V decreases, P increases so that PV stays constant.

5 **A gas is compressed at constant temperature from 6.0 L to 2.0 L. What happens to its pressure?**

A Pressure halves

B Pressure stays the same

C **Pressure triples** ✓

D Pressure becomes 6 times larger

E Pressure becomes one-third

▶ **Explanation:** At constant T (Boyle's law), P1V1 = P2V2. If volume is divided by 3 (6.0→2.0), pressure must multiply by 3.

6 **Charles's law describes the relationship between volume and temperature for a fixed amount of gas at constant pressure:**

- A V is inversely proportional to T(K)
- B **V is directly proportional to T(K)** ✓
- C V is directly proportional to 1/T(K)
- D V is independent of temperature
- E V is proportional to T(°C)

▶ **Explanation:** Charles's law: V/T(K) = constant (at constant P and n). Temperature must be in Kelvin because the relationship uses absolute temperature.

7 **At constant pressure, a gas is heated from 300 K to 600 K. What happens to its volume?**

- A It halves
- B It stays the same
- C **It doubles** ✓
- D It triples
- E It becomes four times larger

▶ **Explanation:** At constant pressure (Charles's law), V T(K). Doubling T from 300→600 K doubles V.

8 **A gas has volume 2.0 L at 300 K. At constant pressure, what is its volume at 450 K?**

- A 1.0 L
- B 2.5 L
- C **3.0 L** ✓
- D 4.0 L
- E 6.0 L

▶ **Explanation:** V2 = V1 × (T2/T1) = 2.0 × (450/300) = 2.0 × 1.5 = 3.0 L.

9 **Gay-Lussac's law relates pressure and temperature when volume and amount are constant:**

- A P is inversely proportional to T(K)
- B **P is directly proportional to T(K)** ✓
- C P is directly proportional to V
- D P is proportional to 1/V
- E P is proportional to T(°C)

▶ **Explanation:** At constant V and n, P/T(K) = constant. Heating increases particle kinetic energy and collision frequency/force, increasing pressure.

10 **A gas has pressure 1.0 atm at 300 K in a rigid container. What is its pressure at 450 K (same container)?**

- A 0.67 atm
- B 1.0 atm

C **1.5 atm** ✓

D 2.0 atm

E 3.0 atm

▶ **Explanation:** P2 = P1 × (T2/T1) = 1.0 × (450/300) = 1.5 atm.

11 Avogadro's law states that at constant temperature and pressure, the volume of a gas is proportional to:

A Its molar mass

B **The number of moles (amount of gas)** ✓

C The number of neutrons

D The density of the gas

E The number of protons

▶ **Explanation:** Avogadro's law: V n (at constant T and P). More moles means more particles, so the gas occupies more volume.

12 At constant temperature and pressure, a gas sample increases from 2.0 mol to 3.0 mol. What happens to its volume?

A It decreases by one-third

B It stays the same

C **It increases by 50%** ✓

D It doubles

E It triples

▶ **Explanation:** V n, so V2/V1 = n2/n1 = 3.0/2.0 = 1.5. That is a 50% increase.

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13 **Which equation is the combined gas law for a fixed amount of gas?**

- A P1 + V1 = P2 + V2
- B P1V1T1 = P2V2T2
- C **P1V1/T1 = P2V2/T2** ✓
- D P1/T1 = V2/P2
- E PV = nT

▶ **Explanation:** For constant n, P1V1/T1 = P2V2/T2 (T in Kelvin). It combines Boyle, Charles, and Gay-Lussac into one relationship.

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14 **A gas has P1 = 1.0 atm, V1 = 2.0 L, T1 = 300 K. If P2 = 0.50 atm and T2 = 450 K, what is V2? (Assume amount of gas is constant.)**

- A 1.5 L
- B 3.0 L
- C 4.0 L
- D **6.0 L** ✓
- E 9.0 L

▶ **Explanation:** V2 = V1 × (P1/P2) × (T2/T1) = 2.0 × (1.0/0.50) × (450/300) = 2.0 × 2 × 1.5 = 6.0 L.

15 **Which set of conditions corresponds to STP in many high-school gas law problems?**

- A 25°C and 1 atm
- B **0°C and 1 atm** ✓
- C 0°C and 2 atm
- D 100°C and 1 atm
- E 273°C and 1 atm

▶ **Explanation:** STP is commonly taken as 0°C (273 K) and 1 atm. (Some conventions use 1 bar; questions should state the convention.)

16 **Assuming an ideal gas occupies 22.4 L per mole at STP, what volume does 0.25 mol occupy at STP?**

- A 2.24 L
- B **5.60 L** ✓
- C 11.2 L
- D 22.4 L
- E 44.8 L

▶ **Explanation:** $V = n \times 22.4 = 0.25 \times 22.4 = 5.60$ L.

17 **Assuming 22.4 L/mol at STP, 11.2 L of an ideal gas at STP corresponds to how many moles?**

- A 0.25 mol

B **0.50 mol** ✓

C 1.0 mol

D 2.0 mol

E 4.0 mol

▶ **Explanation:** n = V/22.4 = 11.2/22.4 = 0.50 mol.

18 **Using PV = nRT with R = 0.0821 L · atm · mol^{-1} · K^{-1}, which unit set must be used to avoid unit mismatch?**

A P in kPa, V in mL, T in °C

B **P in atm, V in L, T in K** ✓

C P in mmHg, V in L, T in °C

D P in atm, V in m^3, T in K

E P in mol, V in L, T in K

▶ **Explanation:** The value R = 0.0821 is specifically for P in atm, V in liters, and T in Kelvin. Other units require a different R value or conversions.

19 **A gas sample has n = 0.50 mol, V = 10.0 L, T = 300 K. What is its pressure? (R = 0.0821 L · atm · mol^{-1} · K^{-1})**

A 0.41 atm

B 0.82 atm

C **1.23 atm** ✓

D 2.46 atm

E 12.3 atm

▶ **Explanation:** $P = nRT/V = (0.50)(0.0821)(300)/10.0 = (0.50)(24.63)/10.0 = 12.315/10.0$ 1.23 atm.

20 **A gas sample has $P = 2.0$ atm, $V = 4.0$ L, $n = 0.50$ mol. What is its temperature? ($R = 0.0821$ L · atm · mol^{-1} · K^{-1})**

- A 49 K
- B 98 K
- C **195 K** ✓
- D 390 K
- E 780 K

▶ **Explanation:** $T = PV/(nR) = (2.0\times4.0)/(0.50\times0.0821) = 8.0/0.04105$ 195 K.

21 **A student uses $PV = nRT$ but accidentally substitutes $T = 25$ (°C) instead of 298 (K). How will the calculated number of moles compare to the correct value?**

- A It will be too small because 25 is smaller than 298
- B **It will be too large because the denominator RT is too small** ✓
- C It will be unchanged because Celsius and Kelvin differ by a constant
- D It will be negative because Celsius can be negative
- E It cannot be predicted without knowing the gas identity

▶ **Explanation:** $n = PV/(RT)$. Using 25 instead of 298 makes RT much smaller, so $PV/(RT)$ becomes much larger. This is a classic trap: gas laws require absolute temperature.

22 **Dalton's law of partial pressures states that, for a mixture of non-reacting gases:**

- A The gas with the greatest molar mass has the greatest partial pressure
- B **Total pressure equals the sum of the partial pressures of each gas** ✓
- C All gases have equal partial pressures regardless of amounts
- D Partial pressure depends only on volume, not on moles
- E Total pressure equals the product of the partial pressures

▶ **Explanation:** Dalton's law: $P_{total} = P1 + P2 + ...$ for non-reacting gases in the same container at the same T and V.

23 **A gas mixture has total pressure 3.0 atm. Two components have partial pressures 1.2 atm and 0.8 atm. What is the partial pressure of the third gas?**

- A 0.6 atm
- B **1.0 atm** ✓
- C 1.2 atm
- D 1.8 atm
- E 2.2 atm

▶ **Explanation:** $P3 = P_{total} - (P1 + P2) = 3.0 - (1.2 + 0.8) = 1.0$ atm.

24 **A mixture contains 2 mol He and 3 mol Ne in the same container. If total pressure is 1.0 atm, what is the partial pressure of He?**

- A 0.20 atm

B **0.40 atm** ✓

C 0.50 atm

D 0.60 atm

E 0.80 atm

▶ **Explanation:** Mole fraction $x_{He} = 2/(2+3) = 2/5 = 0.40$. Partial pressure $P_{He} = x_{He} \times P_{total} = 0.40 \times 1.0 = 0.40$ atm.

25 **A gas mixture has $P_{total} = 4.0$ atm and $P_{N2} = 3.0$ atm. What is the mole fraction of N2?**

A 0.25

B 0.50

C **0.75** ✓

D 1.00

E 3.0

▶ **Explanation:** For ideal mixtures, $P_i = x_i\ P_{total}$. So $x_{N2} = P_{N2}/P_{total} = 3.0/4.0 = 0.75$.

26 **A gas is collected over water at 25°C. Atmospheric pressure is 760 mmHg and water vapor pressure at 25°C is 24 mmHg. What is the pressure of the dry gas?**

A **736 mmHg** ✓

B 760 mmHg

C 784 mmHg

D 24 mmHg

E 0 mmHg

▶ **Explanation:** Collected gas pressure includes water vapor: $P_{atm} = P_{gas} + P_{H2}O$. So $P_{gas} = 760 - 24 = 736$ mmHg.

27 **A gas is collected over water. Atmospheric pressure is 101.3 kPa and water vapor pressure is 3.2 kPa. What is the pressure of the dry gas?**

- A **98.1 kPa ✓**
- B 101.3 kPa
- C 104.5 kPa
- D 3.2 kPa
- E 32.0 kPa

▶ **Explanation:** $P_{gas} = P_{atm} - P_{H2}O = 101.3 - 3.2 = 98.1$ kPa.

28 **For 2H2(g) + O2(g) → 2H2O(g), at the same temperature and pressure, 10 L of H2 would require what volume of O2 for complete reaction?**

- A 2.5 L
- B **5.0 L ✓**
- C 10 L
- D 20 L
- E 40 L

▶ **Explanation:** At the same T and P, gas volumes follow mole ratios. The ratio H2:O2 is 2:1, so 10 L H2 needs 5 L O2.

29 **For $2H_2(g) + O_2(g) \rightarrow 2H_2O(g)$, at the same temperature and pressure, reacting 10 L of H_2 with excess O_2 would produce what volume of $H_2O(g)$?**

- A 5 L
- B **10 L** ✓
- C 15 L
- D 20 L
- E 40 L

▶ **Explanation:** The ratio $H_2:H_2O$ is 2:2 = 1:1, so equal gas volumes of H_2 consumed produce equal gas volumes of H_2O (if treated as gas at the same conditions).

30 **For $N_2(g) + 3H_2(g) \rightarrow 2NH_3(g)$, at the same temperature and pressure, 15 L of H_2 reacting with excess N_2 produces what volume of NH_3?**

- A 5.0 L
- B **10.0 L** ✓
- C 15.0 L
- D 20.0 L
- E 30.0 L

▶ **Explanation:** Volume ratios match mole ratios: 3 volumes H_2 produce 2 volumes NH_3. So NH_3 volume = $15 \times (2/3) = 10$ L.

31 **For $N_2(g) + 3H_2(g) \rightarrow 2NH_3(g)$, you mix 10 L N_2 with 20 L H_2 at the same temperature and pressure. What is the maximum NH_3 volume formed?**

- A 10.0 L

B **13.3 L** ✓

C 20.0 L

D 26.7 L

E 40.0 L

▶ **Explanation:** Need 3 L H2 per 1 L N2. With 20 L H2, the reaction can use at most $20/3 = 6.67$ L N2, so H2 is limiting. NH3 volume $= 20 \times (2/3) = 13.3$ L.

32 **In the same scenario (10 L N2 and 20 L H2), what volume of N2 remains unreacted (same T and P)?**

A 0 L

B **3.33 L** ✓

C 6.67 L

D 10.0 L

E 13.3 L

▶ **Explanation:** H2 limits. 20 L H2 consumes N2 volume $= 20/3 = 6.67$ L. Starting with 10 L N2, leftover $= 10 - 6.67 = 3.33$ L.

33 **Graham's law: The rate of diffusion/effusion is inversely proportional to the square root of molar mass. Compared to O2 (M = 32), H2 (M = 2) effuses at what relative rate?**

A H2 is 2 times faster

B **H2 is 4 times faster** ✓

C H2 is 8 times faster

D H2 is 16 times faster

E H2 is 4 times slower

▶ **Explanation:** $rate_{H2}/rate_{O2} = \sqrt{(M_{O2}/M_{H2})} = \sqrt{(32/2)} = \sqrt{16} = 4$. H2 effuses 4× faster.

34 **Using Graham's law, He (M = 4) effuses how many times faster than CO2 (M = 44)?**

A 1.0×

B 2.2×

C **3.3×** ✓

D 11×

E 44×

▶ **Explanation:** $rate_{He}/rate_{C}O2 = \sqrt{(44/4)} = \sqrt{11}$ 3.32, so about 3.3× faster.

35 **An unknown gas effuses at half the rate of N2 (M = 28). What is the molar mass of the unknown gas (approx.)?**

A 14 g/mol

B 28 g/mol

C 56 g/mol

D **112 g/mol** ✓

E 224 g/mol

▶ **Explanation:** $rate_{unknown}/rate_{N2} = 0.5 = \sqrt{(28/M)}$. Squaring: $0.25 = 28/M \rightarrow M = 28/0.25 = 112$ g/mol.

36 **Ideal gas behavior is most closely approached when gases are at:**

- A High pressure and low temperature
- B **Low pressure and high temperature** ✓
- C High pressure and high temperature
- D Low pressure and low temperature
- E Any pressure as long as temperature is 0°C

▶ **Explanation:** At low pressure, particles are far apart (volume and attractions matter less). At high temperature, kinetic energy overwhelms intermolecular attractions. Both favor ideal behavior.

37 **Real gases deviate MOST from ideal behavior under which conditions?**

- A Low pressure and high temperature
- B **High pressure and low temperature** ✓
- C Low pressure and low temperature
- D High temperature only
- E When volume is measured in liters

▶ **Explanation:** High pressure forces particles close together (finite volume and attractions matter). Low temperature reduces kinetic energy so attractions become significant—so deviation increases.

38 **According to kinetic molecular theory, gas pressure is mainly caused by:**

- A Attraction between gas particles and the container walls
- B Repulsion between gas particles

C **Collisions of gas particles with the container walls** ✓

D The mass of neutrons in the gas particles

E The volume of the container being large

▶ **Explanation:** Pressure is the result of many collisions of gas particles with the container walls, transferring momentum and exerting force per unit area.

39 **Average kinetic energy of gas particles is directly proportional to:**

A Pressure

B Volume

C **Temperature in Kelvin** ✓

D Molar mass

E Number of neutrons

▶ **Explanation:** For gases, average kinetic energy depends only on absolute temperature: higher T(K) means higher average kinetic energy.

40 **If the Kelvin temperature of an ideal gas sample is doubled, what happens to the average kinetic energy of its particles?**

A It halves

B It stays the same

C **It doubles** ✓

D It quadruples

E It becomes $\sqrt{2}$ times larger

▶ **Explanation:** Average kinetic energy T(K). Doubling T doubles average kinetic energy.

41 **The root-mean-square (rms) speed of gas particles is proportional to:**

A T

B **√T** ✓

C 1/T

D M

E M×T

▶ **Explanation:** For an ideal gas, u_{rms} √(T/M). So with fixed molar mass, rms speed is proportional to √T.

42 **At the same temperature, which gas has the highest average molecular speed?**

A **H2** ✓

B N2

C O2

D CO2

E Cl2

▶ **Explanation:** At the same T, lighter gases move faster (u_{rms} 1/√M). H2 has the smallest molar mass, so it has the highest speed.

43 **At the same temperature and pressure, which gas would have the greatest density (assume ideal behavior)?**

A H_2

B He

C N_2

D O_2

E **Cl_2** ✓

▶ **Explanation:** For ideal gases at the same P and T, density $d = PM/RT$, so density is proportional to molar mass M. Cl_2 has the highest molar mass listed.

44 **What is the density of CO_2 at 1.00 atm and 300 K? ($M(CO_2)=44$ g/mol, $R=0.0821\ L \cdot atm \cdot mol^{-1} \cdot K^{-1}$)**

A 0.18 g/L

B 0.74 g/L

C **1.79 g/L** ✓

D 4.40 g/L

E 17.9 g/L

▶ **Explanation:** $d = PM/RT = (1.00\times44)/(0.0821\times300) = 44/24.63$ 1.79 g/L.

45 **A gas has density 1.25 g/L at 1.00 atm and 300 K. What is its molar mass? ($R = 0.0821\ L \cdot atm \cdot mol^{-1} \cdot K^{-1}$)**

A 5.0 g/mol

B 18 g/mol

C **31 g/mol** ✓

D 44 g/mol

E 62 g/mol

▶ **Explanation:** From d = PM/RT, M = dRT/P = 1.25×0.0821×300/1.00 1.25×24.63 30.8 g/mol (~31 g/mol).

46 **O2 gas (M = 32 g/mol) has density 1.50 g/L at 300 K. What is the pressure? (R = 0.0821 L · atm · mol^{-1} · K^{-1})**

A 0.58 atm

B 0.94 atm

C **1.15 atm** ✓

D 2.30 atm

E 4.60 atm

▶ **Explanation:** From d = PM/RT → P = dRT/M = (1.50×0.0821×300)/32 = (1.50×24.63)/32 36.945/32 1.15 atm.

47 **How many moles of gas are in a 5.0 L container at 2.0 atm and 300 K? (R = 0.0821 L · atm · mol^{-1} · K^{-1})**

A 0.081 mol

B **0.41 mol** ✓

C 1.0 mol

D 2.0 mol

E 4.1 mol

▶ **Explanation:** n = PV/RT = (2.0×5.0)/(0.0821×300) = 10/24.63 0.406 mol (~0.41 mol).

48 Convert 380 mmHg to atm (1 atm = 760 mmHg).

A 0.25 atm

B **0.50 atm** ✓

C 1.0 atm

D 2.0 atm

E 760 atm

▶ **Explanation:** 380/760 = 0.50 atm.

49 Convert 2.5 atm to kPa (1 atm = 101.3 kPa).

A 25.3 kPa

B 101 kPa

C 203 kPa

D **253 kPa** ✓

E 506 kPa

▶ **Explanation:** 2.5 × 101.3 253 kPa.

50 **A gas has volume 1.00 L at 27°C and constant pressure. What is its volume at 127°C? (Assume ideal behavior.)**

- A 1.10 L
- B 1.27 L
- C **1.33 L** ✓
- D 2.00 L
- E 4.70 L

▶ **Explanation:** Convert to Kelvin: 27°C = 300 K, 127°C = 400 K. At constant P, V T, so V2 = 1.00×(400/300) = 1.33 L. (4.70 L is the common Celsius-trap.)

51 **A gas has V1 = 2.0 L at P1 = 1.0 atm and T1 = 300 K. What is V2 at P2 = 0.50 atm and T2 = 600 K?**

- A 2.0 L
- B 4.0 L
- C 6.0 L
- D **8.0 L** ✓
- E 12.0 L

▶ **Explanation:** V2 = V1×(P1/P2)×(T2/T1) = 2.0×(1.0/0.50)×(600/300) = 2.0×2×2 = 8.0 L.

52 **A gas has V1 = 3.0 L at P1 = 2.0 atm and T1 = 250 K. What is V2 at P2 = 1.0 atm and T2 = 300 K? (n constant)**

- A 1.8 L
- B 3.6 L

C **7.2 L** ✓

D 12.0 L

E 18.0 L

▶ **Explanation:** V2 = 3.0×(2.0/1.0)×(300/250) = 3.0×2×1.2 = 7.2 L.

53 At constant volume and temperature, the number of moles of a gas is doubled. What happens to the pressure?

A It halves

B It stays the same

C **It doubles** ✓

D It quadruples

E It becomes zero

▶ **Explanation:** From PV = nRT, at constant V and T, P n. Doubling n doubles P.

54 At constant temperature and pressure, the amount of gas increases by 50%. What happens to volume?

A Volume decreases by 50%

B Volume stays the same

C **Volume increases by 50%** ✓

D Volume doubles

E Volume quadruples

▶ **Explanation:** At constant T and P, Avogadro's law gives V n. Increasing n by 50% increases V by 50%.

55 **A fixed amount of gas is compressed isothermally so that its volume becomes one-quarter of its original value. What happens to pressure?**

- A Pressure becomes one-quarter
- B Pressure becomes one-half
- C Pressure stays the same
- D **Pressure becomes 4 times larger** ✓
- E Pressure becomes 16 times larger

▶ **Explanation:** Boyle's law: P 1/V at constant T. If $V \to V/4$, then $P \to 4P$.

56 **Which graph would be a straight line for Boyle's law (constant T and n)?**

- A P vs V
- B **P vs 1/V** ✓
- C V vs T(°C)
- D P vs T(K)
- E V vs P^2

▶ **Explanation:** Boyle's law gives $P = k(1/V)$. So plotting P against 1/V is linear (straight line). P vs V is a curve (hyperbola).

57 **Which graph should be approximately a straight line passing through the origin (for ideal behavior) for Charles's law?**

- A **V vs T(K)** ✓
- B V vs T(°C)
- C P vs V
- D P vs 1/T(K)
- E P vs V^2

▶ **Explanation:** Charles's law: V T(K). Using Kelvin gives a line through the origin (0 K → 0 volume in the extrapolated model). Using Celsius shifts the graph.

58 **Extrapolating the V vs T(K) line for an ideal gas suggests that gas volume would reach zero at approximately:**

- A 0°C
- B **−273°C** ✓
- C 273°C
- D −100°C
- E 100°C

▶ **Explanation:** Absolute zero is 0 K, which corresponds to about −273°C. The extrapolation gives the idea of an absolute temperature scale.

59 **For a fixed amount of gas, if pressure and volume are both held constant, which relationship must be true?**

- A n is proportional to T

- B **nT is constant** ✓
- C n/V is constant
- D P/T is constant
- E V/T is constant

▶ **Explanation:** From $PV = nRT$, if P and V are constant, then nT must be constant (since PV/R is constant). If n increases, T must decrease to keep nT constant.

60 **A sealed rigid container holds O2 at some temperature. If you add helium to the container and keep the temperature constant, what happens to the partial pressure of O2?**

- A It increases because total pressure increases
- B It decreases because helium "dilutes" oxygen
- C **It stays the same because P_{O2} depends only on n_{O2}, T, and V** ✓
- D It becomes zero because oxygen is displaced
- E It depends on molar mass, so it must change

▶ **Explanation:** For ideal gases, $P_i = n_i RT/V$. Adding another gas changes total pressure, but n_{O2}, T, and V are unchanged, so P_{O2} stays the same.

61 **A gas mixture is in a container at constant temperature. The container's volume is doubled without changing the number of moles of each gas. What happens to each partial pressure?**

- A Each partial pressure doubles
- B Each partial pressure stays the same
- C **Each partial pressure halves** ✓

D Only the heaviest gas partial pressure changes

E Partial pressures cannot be predicted without knowing the gases

▶ **Explanation:** $P_i = n_iRT/V$. If V doubles (T and n_i constant), each P_i halves. Therefore total pressure also halves.

62 **A gas mixture is in a rigid container at constant temperature. Half of the gas molecules (of every component) are removed. What happens to total pressure?**

A It doubles

B It stays the same

C **It halves** ✓

D It becomes zero

E It increases by 50%

▶ **Explanation:** At constant V and T, pressure is proportional to total moles n_{total}. Removing half the molecules halves n_{total}, so total pressure halves.

63 **For 2H2(g) + O2(g) → 2H2O(g), 1.00 L of O2 at 1.00 atm and 300 K reacts with excess H2. How many moles of H2O form? ($R = 0.0821\ L \cdot atm \cdot mol^{-1} \cdot K^{-1}$)**

A 0.0203 mol

B 0.0406 mol

C **0.0812 mol** ✓

D 0.162 mol

E 0.406 mol

▶ **Explanation:** First find moles O2: $n = PV/RT = 1.00\times1.00/(0.0821\times300) = 1/24.63$ 0.0406 mol. Stoichiometry: 1 mol O2 → 2 mol H2O, so H2O moles $2\times0.0406 = 0.0812$ mol.

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64 **CaCO3(s) → CaO(s) + CO2(g). If 5.00 g CaCO3 decomposes completely, what volume of CO2 is produced at 1.00 atm and 300 K? (Mr(CaCO3)=100.0, R=0.0821)**

- A 0.246 L
- B 0.616 L
- C **1.23 L** ✓
- D 2.46 L
- E 12.3 L

▶ **Explanation:** Moles CaCO3 = 5.00/100.0 = 0.0500 mol → 0.0500 mol CO2. Volume: $V = nRT/P = 0.0500\times0.0821\times300/1.00 = 0.0500\times24.63$ 1.23 L.

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65 **What mass of N2 is contained in 10.0 L at 1.00 atm and 300 K? (M(N2)=28 g/mol, R=0.0821)**

- A 1.14 g
- B 5.6 g
- C **11.4 g** ✓
- D 28.0 g
- E 114 g

▶ **Explanation:** $n = PV/RT = 10.0/(0.0821\times300) = 10.0/24.63$ 0.406 mol. Mass = nM 0.406×28 11.4 g.

66 **Which statement is correct at the same temperature and pressure?**

A 1.0 L of He contains more molecules than 1.0 L of CO2

B 1.0 L of He contains fewer molecules than 1.0 L of CO2

C **1.0 L of He contains the same number of molecules as 1.0 L of CO2** ✓

D Only gases with the same molar mass have equal molecules in equal volumes

E The number of molecules depends only on molar mass, not volume

▶ **Explanation:** Avogadro's law: equal volumes of gases at the same T and P contain equal numbers of molecules, regardless of gas identity.

67 **At constant temperature and pressure, 0.25 mol He occupies 3.0 L. What volume will 0.50 mol He occupy?**

A 1.5 L

B 3.0 L

C 4.5 L

D **6.0 L** ✓

E 12.0 L

▶ **Explanation:** At constant T and P, V n. Doubling moles from 0.25 to 0.50 doubles volume from 3.0 L to 6.0 L.

68 **One reason the measured pressure of a real gas can be LOWER than the ideal gas prediction (at the same T, V, n) is that:**

A Gas particles have no volume

B **Intermolecular attractions pull particles away from the walls, reducing collision force ✓**

C The container walls absorb gas and increase collisions

D Real gases have fewer moles than ideal gases

E Temperature becomes negative at high pressure

▶ **Explanation:** Attractions between particles reduce the momentum delivered to container walls, lowering measured pressure compared to ideal predictions, especially at moderate pressures/low temperatures.

69 One reason the measured pressure of a real gas can be HIGHER than the ideal gas prediction at very high pressures is that:

A **Gas particles occupy significant volume, reducing free space and increasing collision frequency ✓**

B Intermolecular attractions become infinite and increase pressure

C Molar mass increases with pressure

D Avogadro's constant changes at high pressure

E Temperature automatically becomes higher

▶ **Explanation:** At very high pressure, the finite volume of gas particles matters: the free volume is smaller than V, so collisions happen more frequently than the ideal model assumes, raising pressure.

70 Which statement is true for an ideal gas (a key consequence of the ideal gas model)?

A Internal energy depends on pressure only

B Internal energy depends on volume only

C **Internal energy depends only on temperature ✓**

D Internal energy depends on molar mass only

E Internal energy is always zero for gases

▶ **Explanation:** In the ideal gas model, there are no intermolecular potential energy effects, so internal energy is purely kinetic and depends only on temperature.