Lectures: Physics 3306

Provides an introduction to a wide variety of topics in classical (pre-quantum) physics as a bridge to prepare students for subsequent upper-level courses in physics. The topics covered include thermodynamics, fluid mechanics, mechanical waves, optics, radiation, electromagnetic phenomena, atoms, and laboratory techniques. Prerequisites: C- or better in PHYS 1106; and in PHYS 1304 or PHYS 1308.

Saptaparna Bhattacharya

March 6th, 2026

Based on Simon Dalley's lectures taught in Spring 2025

Labs

Lectures

Schedule

No class

Month	Monday	Tuesday	Wednesday	Thursday	Friday	Saturday	Sunday
January	19	20	21 ✅	22	23 ✅	24	25
	26 ❄️🌨️❄❆❄	27	28 ❄️🌨️❄❆❄	29	30 ✅	31	1
February	2 ✅	3	4 ✅	5	6 ✅	7	8
	9 ✅	10	11 HWB due ✅	12	13 ✅	14	15
	16 ✅	17	18 ✅	19	20 HWC due ✅	21	22
	23 Hegi Center ✅	24	25 HWD due ✅	26	27 ✅	28	1
March	2 ✅	3	4 HWE due	5	6 ✅	7	8
	9	10	11	12	13	14	15
	16	17	18	19	20	21	22
	23	24	25	26	27	28	29
April	30	31	1	2	3	4	5

Labs

Lectures

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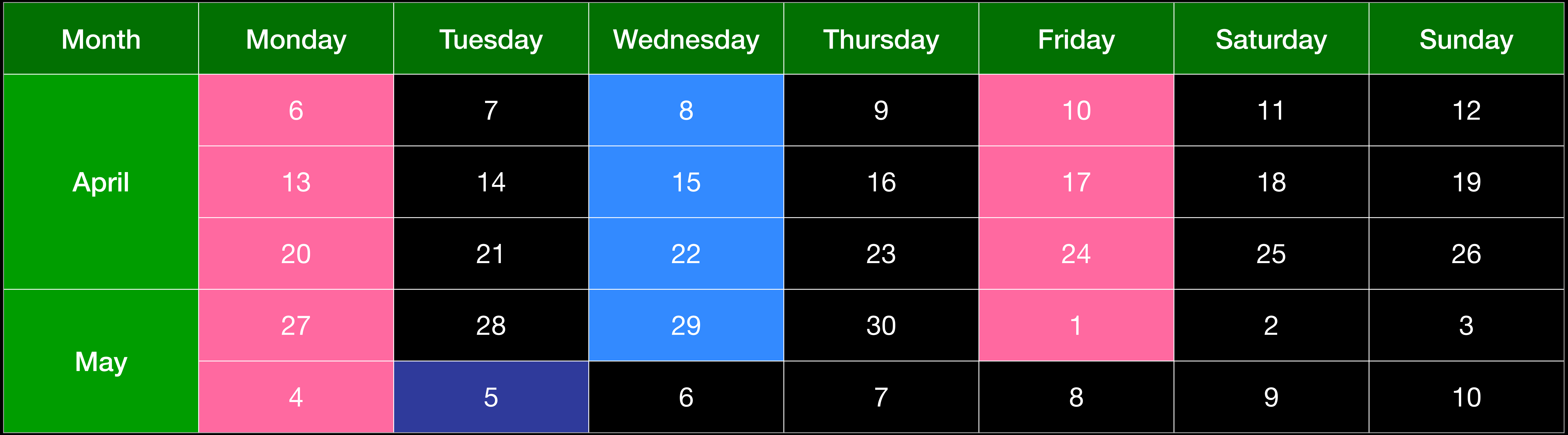

Month	Monday	Tuesday	Wednesday	Thursday	Friday	Saturday	Sunday
April	6	7	8	9	10	11	12
	13	14	15	16	17	18	19
	20	21	22	23	24	25	26
May	27	28	29	30	1	2	3
	4	5	6	7	8	9	10

Kinetic Theory of Gases
Halliday 19.1-19.5

Key concepts: Sample size

- Let’s define a sample size in *moles*
  - Aids in comparison of samples: we can be certain that we are comparing samples that contain the same number of atoms or molecules
  - The *mole* is one of the seven SI base units and is defined as:
    - One *mole* is the number of atoms in a 12 g sample of carbon-12
    - Number of atoms $N_A = 6.02 \times 10^{23}\ \mathrm{mol}^{-1}$ (per mole)
    - Avogadro’s number
    - *Named after Italian scientist Amedeo Avogadro (1776–1856), who suggested that all gases occupying the same volume under the same conditions of temperature and pressure contain the same number of atoms or molecules*

Key concepts: Macroscopic properties of matter

- Can we characterize the macroscopic properties of a gas:
  - pressure and temperature in terms of the behavior of molecules that make it up
  - But, which gas?
  - Through experiments, we have found that if we confine 1 mol samples of various gases in boxes of identical volume and hold the gases at the same temperature, then their measured pressures are almost the same, and at lower densities the differences tend to disappear

Hydrogen
(1 mol)

Oxygen
(1 mol)

Uranium
hexafluoride
(1 mol)

All of these gases are at identical volume and temperature

Key concepts: Ideal gas equation

- All real gases obey the relation (low gas density, low temperature):
  - $pV = nRT$ (ideal gas law)
  - p = absolute pressure, n = number of moles of gas present, T = temperature in kelvins, V = volume occupied
  - $R = 8.31\ \mathrm{J/mol\,.\,K}$
  - One can also write this equation in terms of Boltzmann constant k
    - $k = \dfrac{R}{N_A}$

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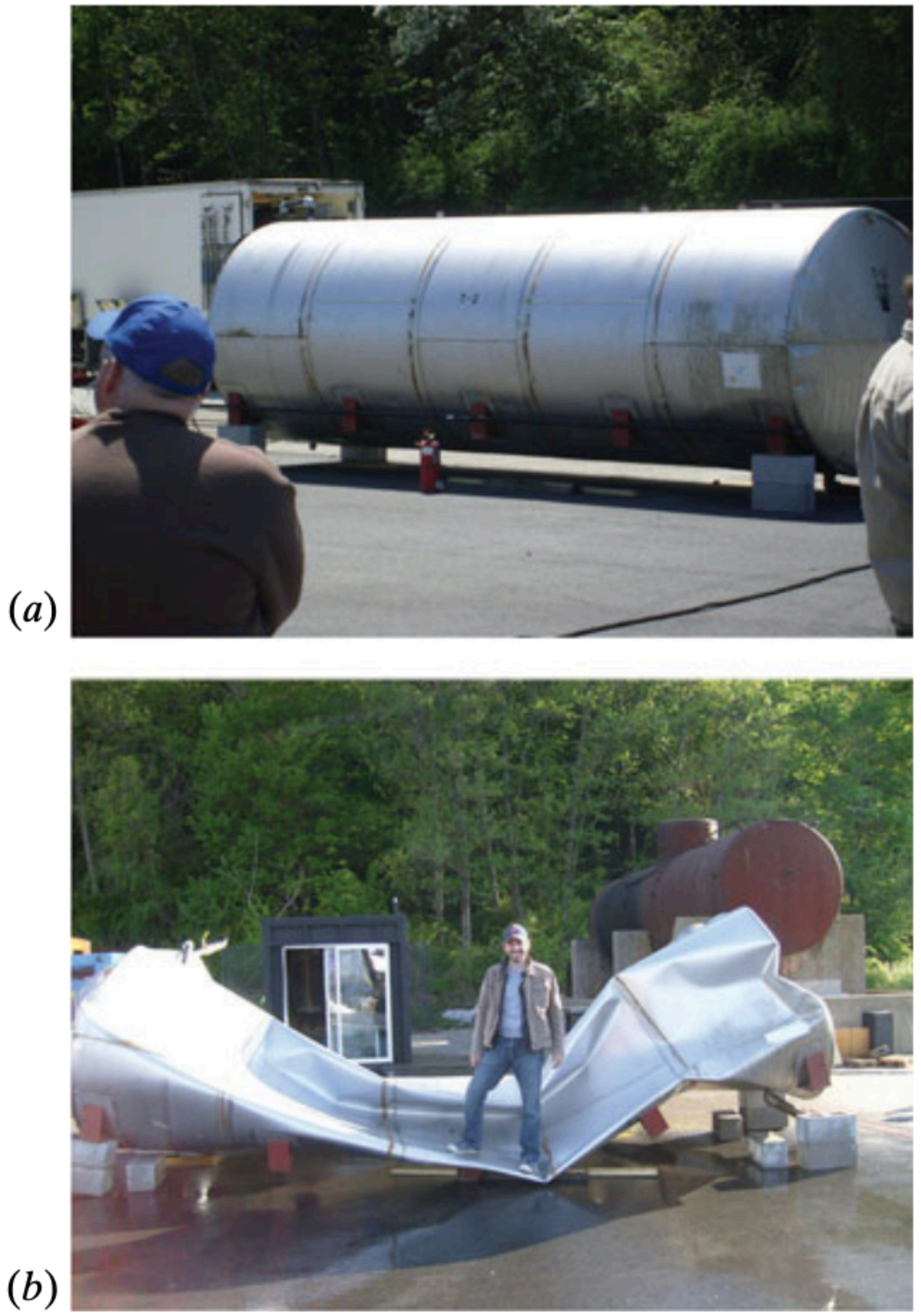

Courtesy www.doctorslime.com

Figure 19-1 (*a*) Before and (*b*) after images of a large steel tank crushed by atmospheric pressure after internal steam cooled and condensed.

Filled with steam at 110˚C

Cold water poured: collapse occurs

What happened?

Key concepts: Ideal gas equation

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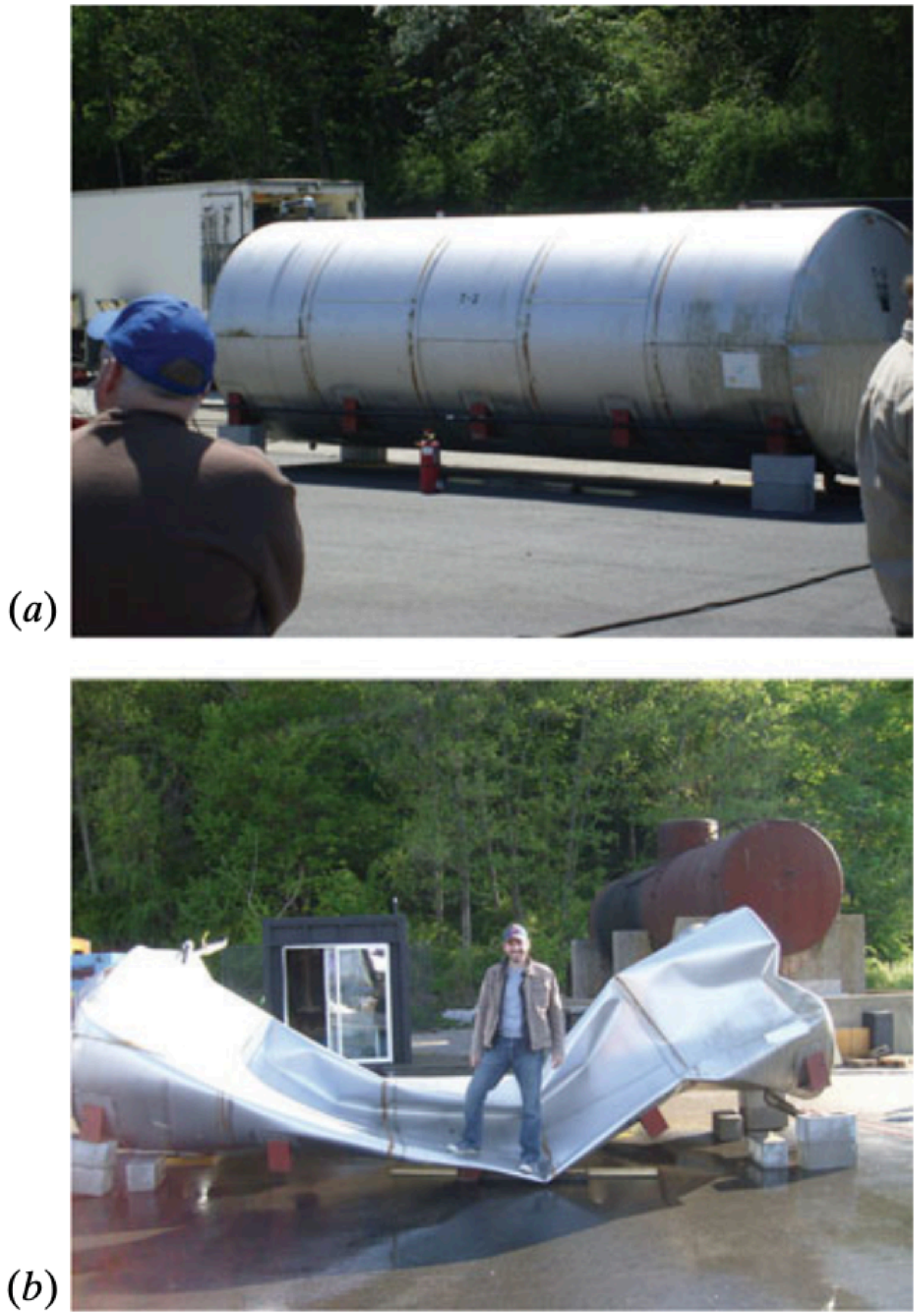

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Cold water poured: collapse occurs

T↓, n↓, V constant, p↓

hydrogen H is collected from space, where it is monatomic, and put into a container where it forms H_2 molecules.

What is the number of moles multiplied by?

a) 2

b) 1/2

c) unchanged

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- **The Concept of a Mole:** A mole is simply a unit of measurement used to count particles, representing a specific number (Avogadro's number).
- **The Process:** You are starting with a certain number of individual, monatomic hydrogen atoms (H). When they are placed in the container, they pair up to form diatomic molecules (H2).
- **The Result:** Because it takes exactly two H atoms to make one H2 molecule, the total number of independent particles in the container is cut in half.

Since the total number of particles is halved, the number of moles is also halved.

Key concepts: Isotherm

- Isothermal expansion:
  - Gas expands from initial volume i (V_i) to final volume f (V_f)
  - Such a process at constant temperature is called an **isothermal expansion**
- On the $p - V$ curve, an isotherm is a curve that connects points that the same temperature
- For n moles, it is a graph of:
  - $p = nRT\dfrac{1}{V} = (\text{a constant})\dfrac{1}{V}$

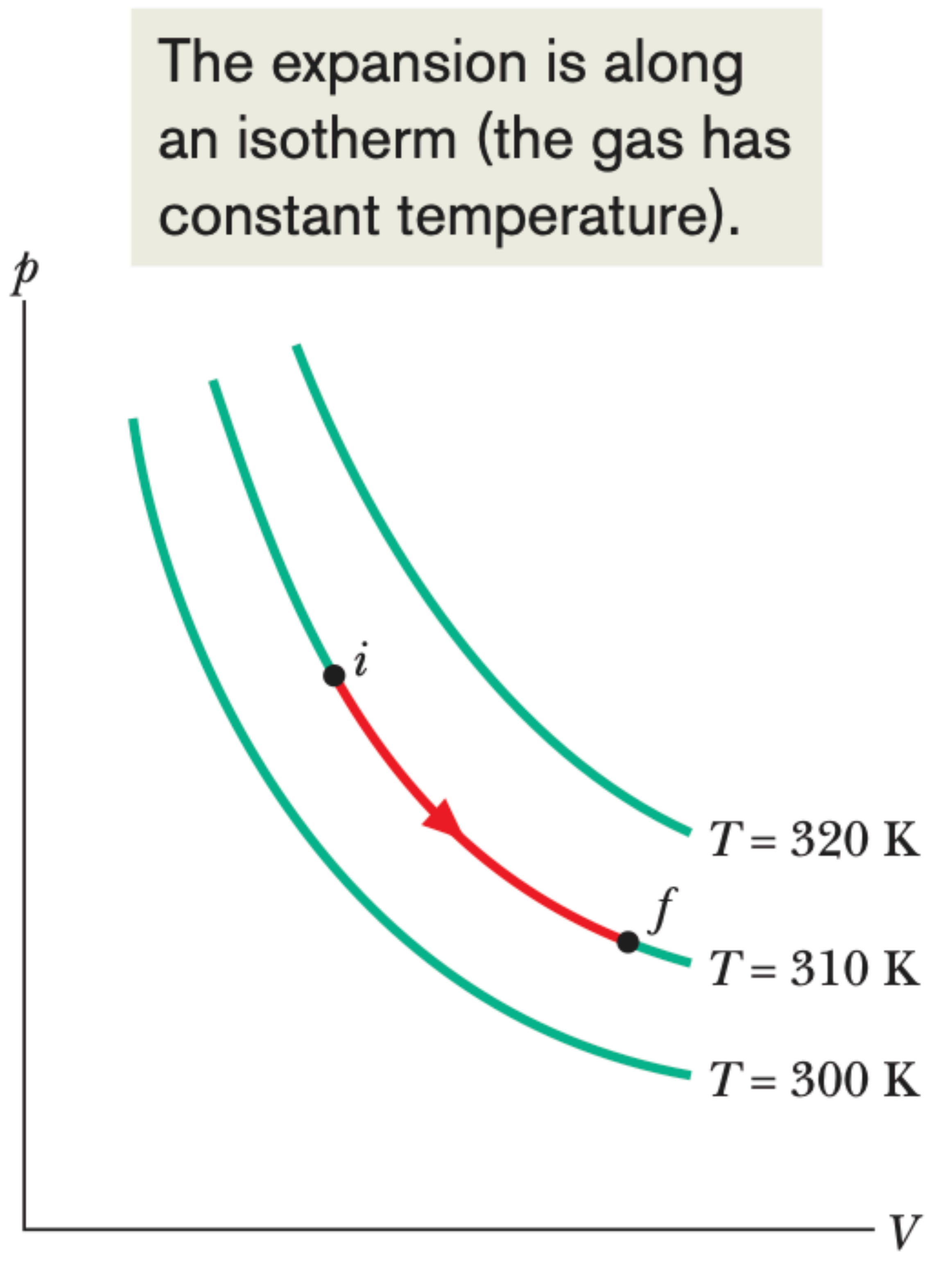
The expansion is along an isotherm (the gas has constant temperature).

Figure 19-2 Three isotherms on a p-V diagram. The path shown along the middle isotherm represents an isothermal expansion of a gas from an initial state i to a final state f. The path from f to i along the isotherm would represent the reverse process—that is, an isothermal compression.

An ideal gas has an initial pressure of 3 Pa and volume 4 m^3.

Which is the final pressure and volume of the process that ends on the same isotherm?

a) 5 Pa and 7 m^3

b) 5 Pa and 2 m^3

c) 6 Pa and 8 m^3

d) 8 Pa and 6 m^3

e) 6 Pa and 2 m^3

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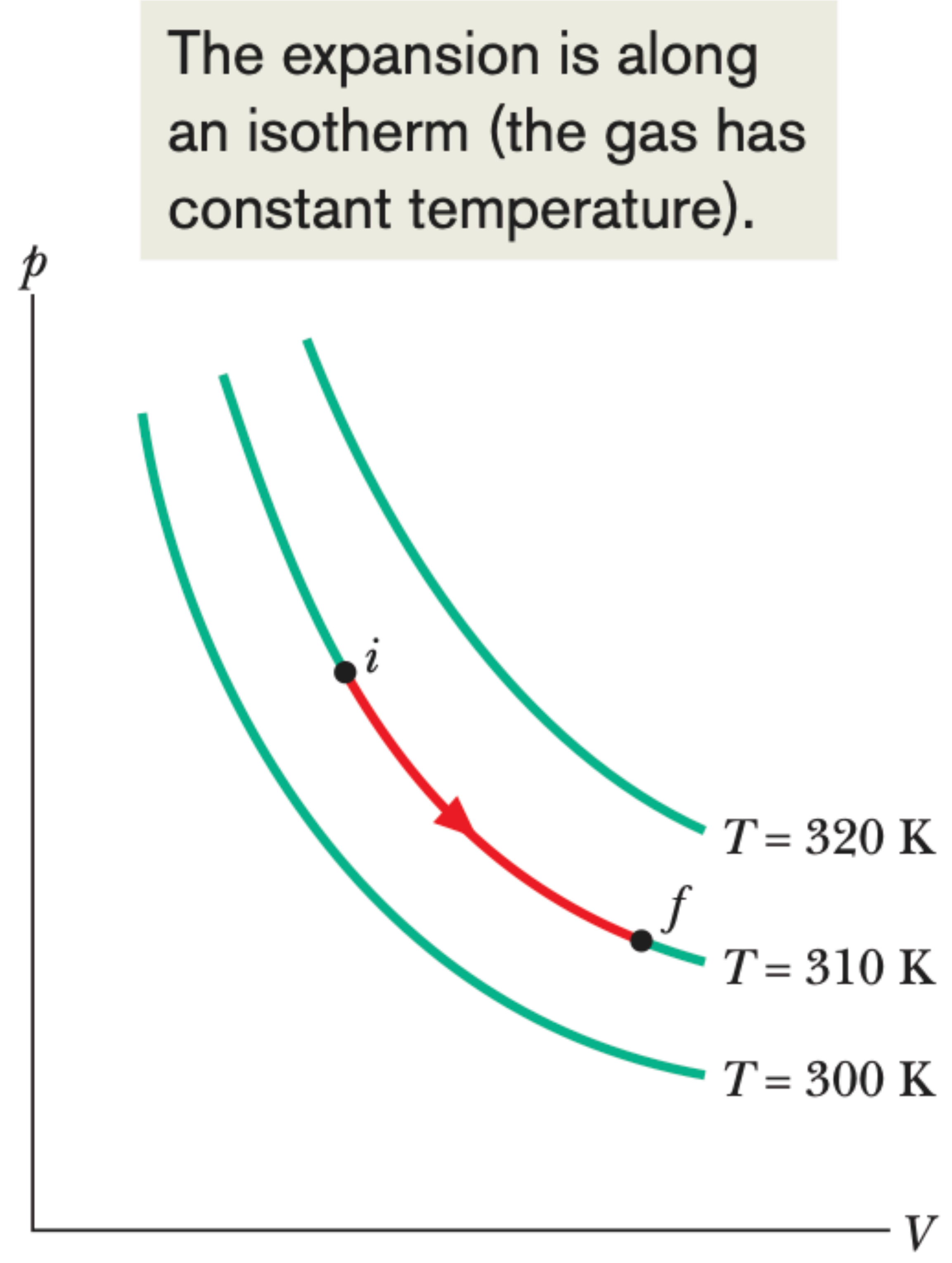
The Concept: The problem states that the process ends on the “same isotherm”. In thermodynamics, an isotherm represents a curve on a graph where the temperature remains constant. According to Boyle's Law, if the temperature and the amount of gas are held constant, the product of the gas's pressure and volume is also constant. * **The Formula:** $P_1V_1 = P_2V_2$

The Calculation: 1. Find the product of your initial state. Your initial pressure is 3 Pa and initial volume is 4 m^3. 2. Multiply these together: 3 * 4 = 12. 3. Because the final state must lie on the same isotherm, its pressure and volume must also multiply to 12. Next, test the given options: * a) 5 * 7 = 35 * b) 5 * 2 = 10 * c) 6 * 8 = 48 * d) 8 * 6 = 48 * **e) 6 * 2 = 12**

Key concepts: Work done

- To find the work done by an ideal gas during an isothermal expansion, we use:
  - $W = \int_{V_i}^{V_f} p dV$
  - Using the ideal gas law:
    - $W = \int_{V_i}^{V_f} \frac{nRT}{V} dV$
    - $W = nRT \int_{V_i}^{V_f} \frac{dV}{V}$
    - $W = nRT \int_{V_i}^{V_f} \frac{dV}{V} = nRT \left[\ln V\right]_{V_i}^{V_f}$
    - $W = nRT \ln \frac{V_f}{V_i}$



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Key concepts: Pressure, temperature and RMS speed

- Change in particle's momentum along the x-axis:
  - $\Delta p_x = (-mv_x) - (mv_x)$
  - $\Delta p_x = -2mv_x$ (collision with wall is elastic)
  - Δp_x delivered to the wall is positive
  - Δt = time between collisions is the time the molecule takes to travel to the opposite wall and back again ($2L$) at speed v_x
  - Average rate of change of momenta:
    - $$\frac{\Delta p_x}{\Delta t} = \frac{2mv_x}{2L/v_x} = \frac{mv_x^2}{L}$$

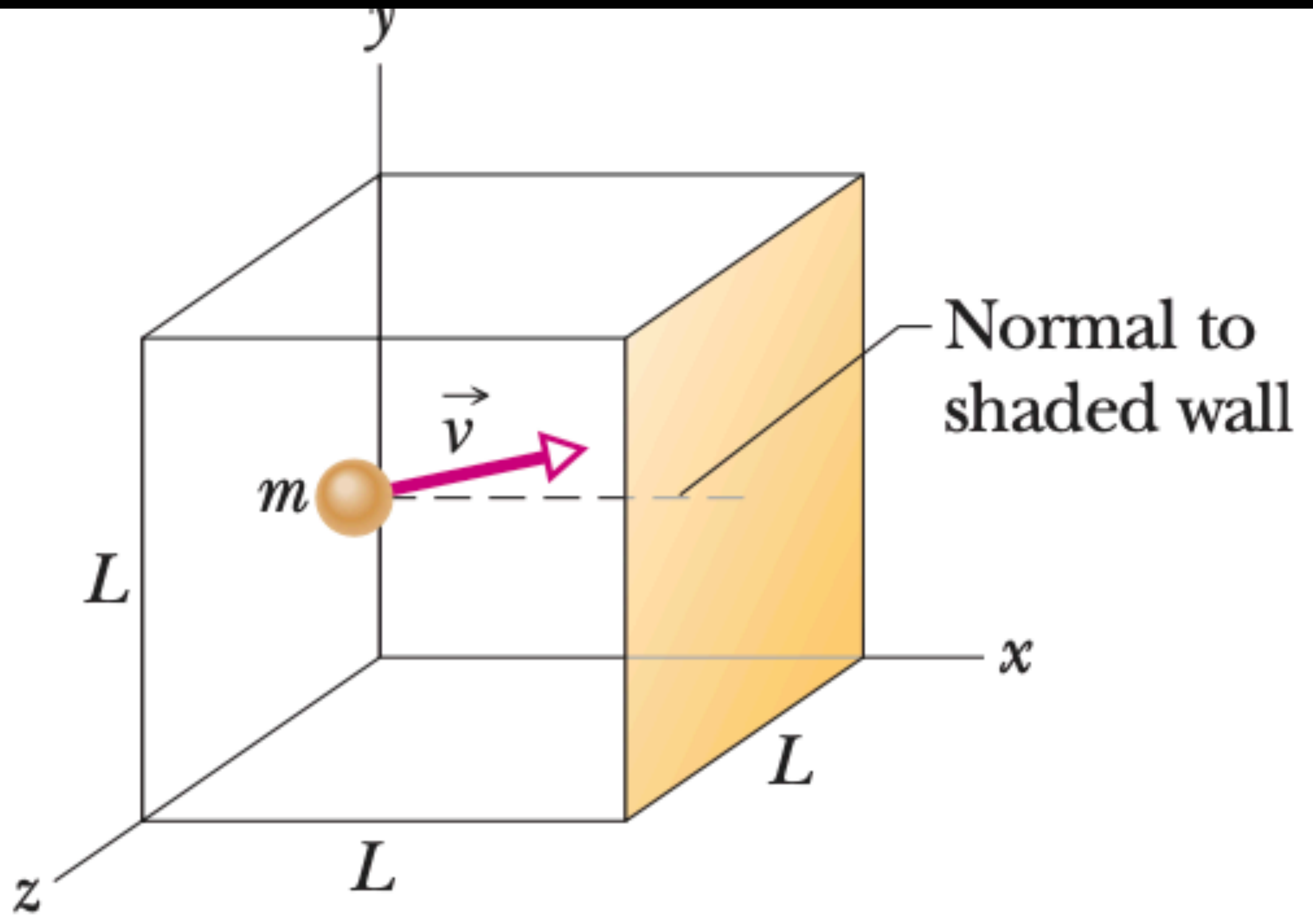
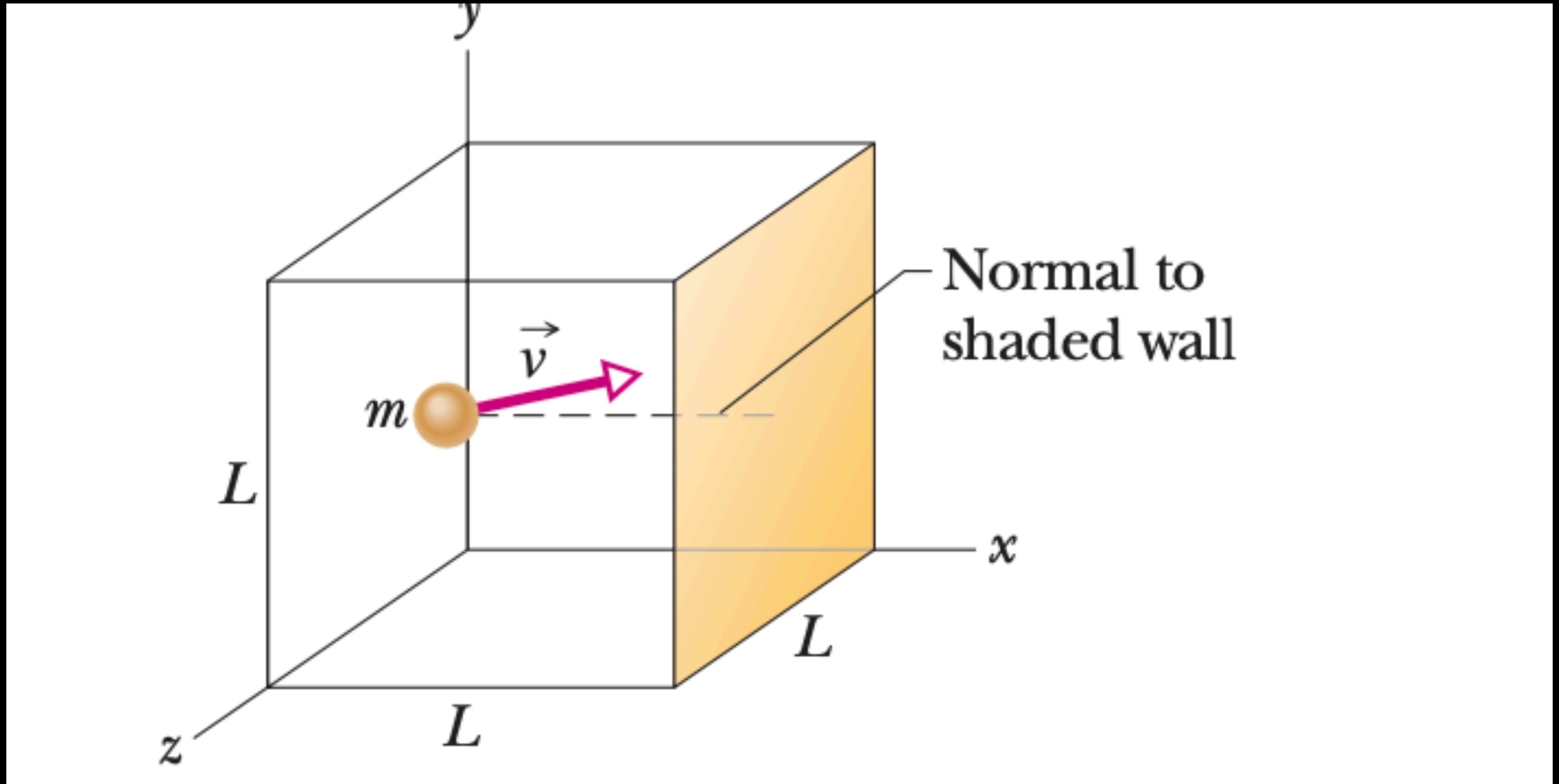


Figure 19-4 A cubical box of edge length L, containing n moles of an ideal gas. A molecule of mass m and velocity $\vec{v}$ is about to collide with the shaded wall of area L^2. A normal to that wall is shown.

Key concepts: Pressure, temperature and RMS speed

- Average rate of change of momenta:
- $\dfrac{\Delta p_x}{\Delta t} = \dfrac{2mv_x}{2L/v_x} = \dfrac{mv_x^2}{L}$
- This is force (F_x)
- $p = \dfrac{F_x}{L^2}$
- $p = \dfrac{mv_{x1}^2/L + mv_{x2}^2/L + \cdots + mv_{xN}^2/L}{L^2}$
- $p = \left(\dfrac{m}{L^3}\right)(v_{x1}^2 + v_{x2}^2 + \cdots + v_{xN}^2)$



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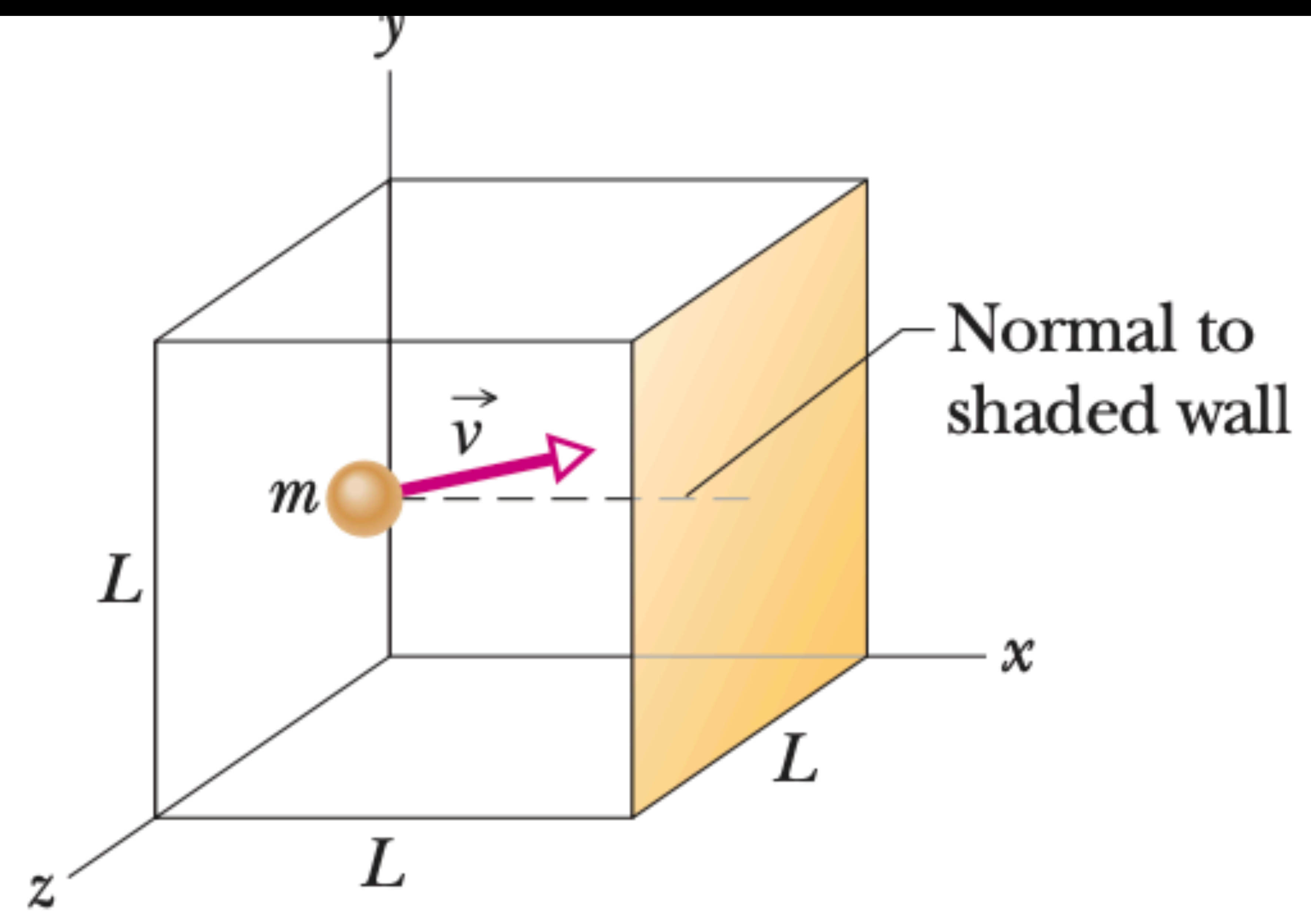
Key concepts: Pressure, temperature and RMS speed

- Average rate of change of momenta:
- $p = \left(\dfrac{m}{L^3}\right)(v_{x1}^2 + v_{x2}^2 + \cdots + v_{xN}^2)$
- $p = \dfrac{nmN_A}{L^3}(v_x^2)_{\text{avg}}$ $\longleftrightarrow$ $N = nN_A$
- $v_x^2 = \dfrac{1}{3}v^2$ on average
- $p = \dfrac{nMv_{\text{rms}}^2}{3V}$
- $v_{\text{rms}} = \sqrt{\dfrac{3RT}{M}}$

$mN_A = \text{molar mass (M)}$

$\sqrt{(v^2)_{\text{avg}}} = v_{\text{rms}}$



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Key concepts: Kinetic Theory

- Average translational kinetic energy:
  - $K_{\text{avg}} = \left(\frac{1}{2}mv^2\right)_{\text{avg}} = \frac{1}{2}m(v^2)_{\text{avg}} = \frac{1}{2}mv_{\text{rms}}^2$
  - $K_{\text{avg}} = \frac{1}{2}m\frac{3RT}{M} = \frac{3RT}{2N_A}$ (M/m: molar mass/mass of the molecule is Avogadro's number)
  - $K_{\text{avg}} = \frac{3}{2}kT\,(k = R/N_A)$
- At a given temperature T, all ideal gas molecules—no matter what their mass—have the same average translational kinetic energy—namely, $K_{\text{avg}} = \frac{3}{2}kT$. When we measure the temperature of a gas, we are also measuring the average translational kinetic energy of its molecules

Which one of the following statements is ***not consistent*** with the kinetic theory of gases?

(a)Individual gas molecules are relatively far apart.

(b)The actual volume of the gas molecules themselves is very small compared to the volume occupied by the gas at ordinary temperatures and pressures.

(c) The average kinetic energies of different gases are different at the same temperature.

(d) There is no net gain or loss of the total kinetic (translational) energy in collisions between gas molecules.

(e) The theory explains most of the observed behavior of gases at ordinary temperatures and pressures.

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$$K_{\text{avg}} = \frac{3}{2}kT$$

Temperature is the only factor determining average kinetic energy, **all gases at the same temperature have the exact same average kinetic energy**, regardless of their chemical identity or molecular mass.

A gas mixture in equilibrium consists of molecules with molecular masses $m_1 > m_2 > m_3$.

Rank the three types according to their average kinetic energy, greatest first.

a) $1 > 2 > 3$

b) $3 > 2 > 1$

c)$1 = 2 = 3$

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Thermal Equilibrium: The key phrase in the problem is "in equilibrium." For a gas mixture, being in thermal equilibrium means that all the different types of molecules in the container have reached the exact same temperature (T).

The Kinetic Energy Formula: As we touched on in the previous question, the average translational kinetic energy of a gas molecule depends *only* on its absolute temperature, determined by the formula $K_{\text{avg}} = \frac{3}{2}kT$

***Mass is Irrelevant to KE:** Notice that mass (m) does not appear anywhere in that equation. The lighter molecules (like m_3) will be zipping around at a much higher average *speed* to compensate for their lack of mass, but the average *kinetic energy* across all three types will be perfectly identical because they share the same temperature.

A gas mixture in equilibrium consists of molecules with molecular masses $m_1 > m_2 > m_3$.

Rank the three types according to their *rms* speed, greatest first.

a) 1 > 2 > 3

b) 3 > 2 > 1

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$$v_{\mathrm{rms}} = \sqrt{\frac{3RT}{M}} = \sqrt{\frac{3RT}{mN_A}}$$

Rank the gases A, B, C, for their *rms* speed, greatest first.

Gas	**Temperature (K)**	**Molar Mass (/kg mol)**
A	400	0.04
B	360	0.03
C	280	0.02

a) C > B > A

b) A > B > C

c) C > A > B

d) A = B = C

e) C > B = A

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$$v_{\text{rms}} = \sqrt{\frac{3RT}{M}} = \sqrt{\frac{3RT}{mN_A}}$$

Key concepts: Mean Free Path

- Mean free path:
  - Denoted by λ:
    - It is the average distance traversed by a molecule between collisions
  - $\lambda = \dfrac{1}{\sqrt{2}\pi d^2 N/V}$
  - N/V = number of molecules per unit volume
  - d = size of the molecules

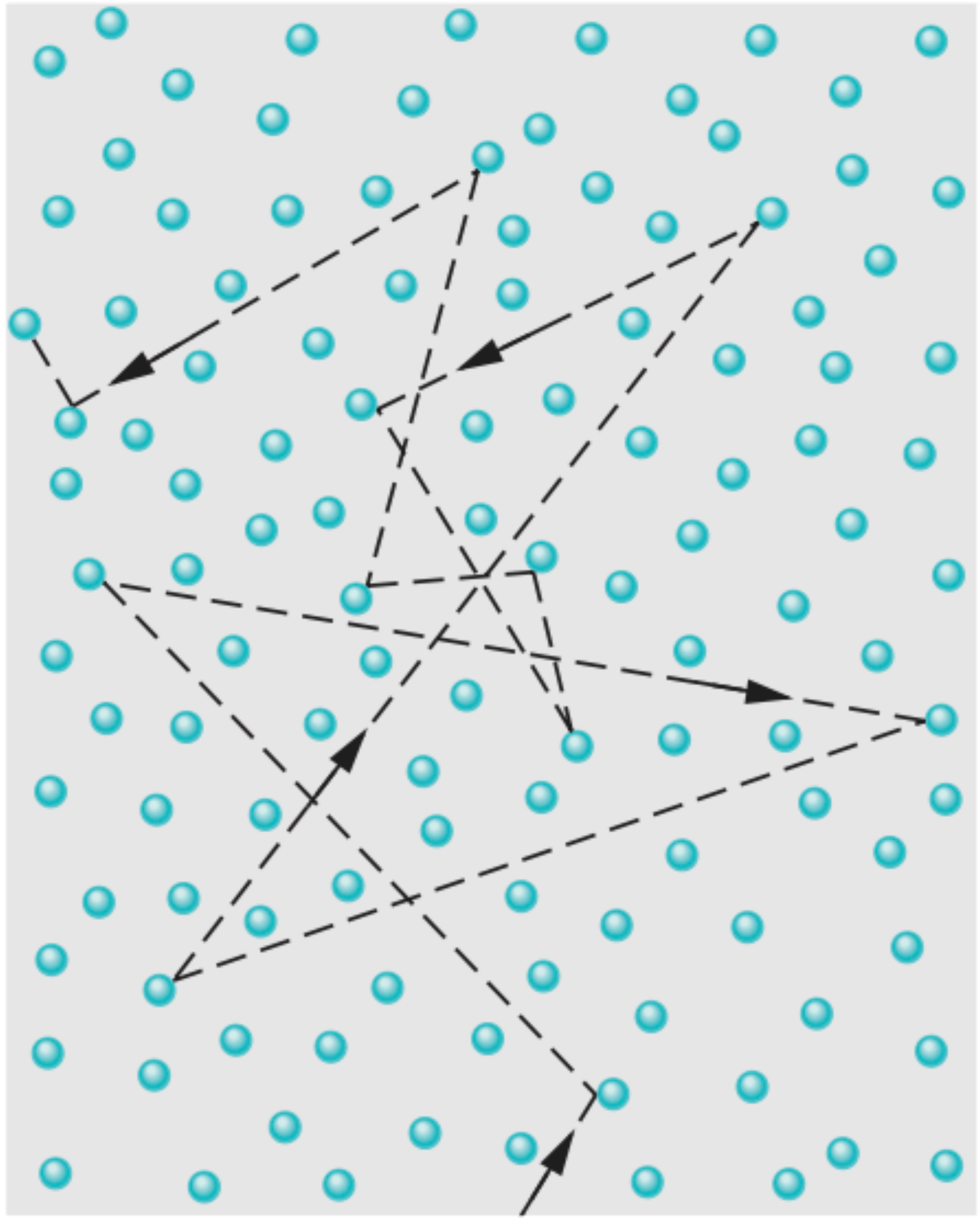

Figure 19-5 A molecule traveling through a gas, colliding with other gas molecules in its path. Although the other molecules are shown as stationary, they are also moving in a similar fashion.

Key concepts: Mean Free Path

- Mean free path: average distance traversed by a molecule between collisions
- $\lambda = \dfrac{\text{length of path during } \Delta \text{t}}{\text{number of collisons in } \Delta \text{t}}$
- $\lambda \approx \dfrac{v\Delta t}{N/V} \dfrac{1}{\pi d^2 v \Delta t}$ (differs from actual formula because this treatment assumes all other molecules are at rest, except one)

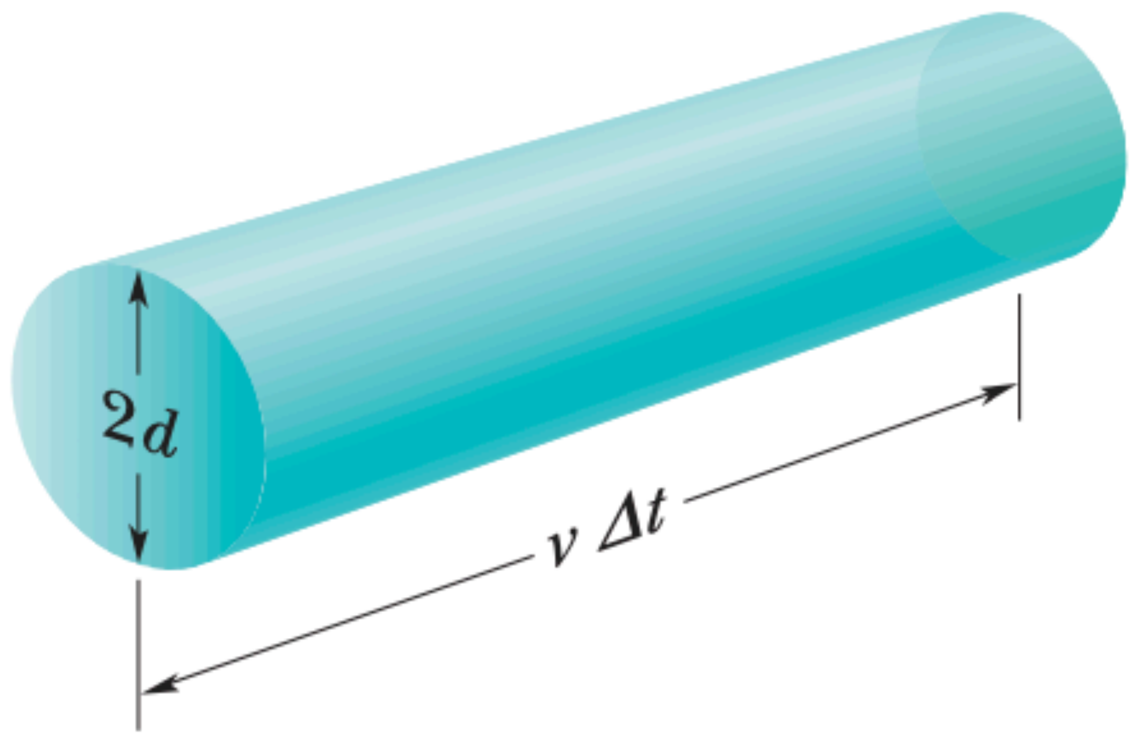

Figure 19-7 In time Δt the moving molecule effectively sweeps out a cylinder of length $v\,\Delta t$ and radius d.

One mole of gas A has molecular diameter $2d_0$ and average speed molecular speed v_0.
One mole of gas B has smaller molecules of diameter d_0 but higher average speed $2v_0$.

Within the same container, which has the greater mean free path?

a) A

b) B

c) they are the same

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Within the same container, which has the greater mean free path?

a) A

b) B

c) they are the same

Because Gas B has smaller molecules, they present a smaller target for collisions. Dividing by a smaller diameter squared results in a larger mean free path. Specifically, Gas B's mean free path will be four times greater than Gas A's.

One mole of gas A has molecular diameter $2d_0$ and average speed molecular speed v_0.
One mole of gas B has smaller molecules of diameter d_0 but higher average speed $2v_0$.

Within the same container, which has the greater average collision rate?

a) A

b) B

c) they are the same

One mole of gas A has molecular diameter $2d_0$ and average speed molecular speed v_0.
One mole of gas B has smaller molecules of diameter d_0 but higher average speed $2v_0$.

Within the same container, which has the greater average collision rate?

a) A

b) B

c) they are the same

Comparing the results, $4d_0^2v_0$ is greater than $2d_0^2v_0$. Even though Gas B's molecules are moving twice as fast, Gas A's molecules have twice the diameter, which gives them four times the cross-sectional area. This massive increase in target size more than makes up for their slower speed, causing Gas A to experience collisions much more frequently.