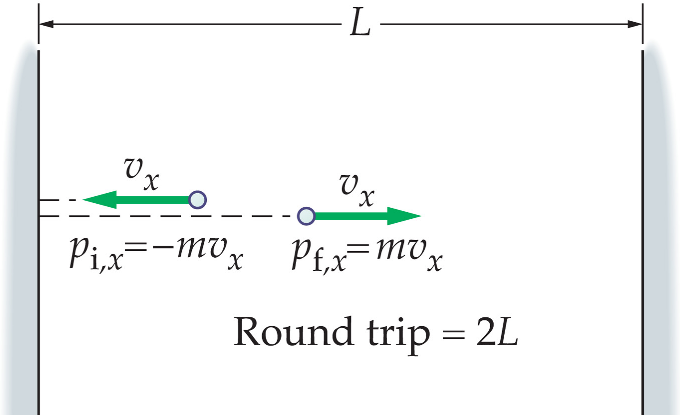
Physics 115
General Physics II

Session 9



Molecular motion and temperature
Phase equilibrium, evaporation

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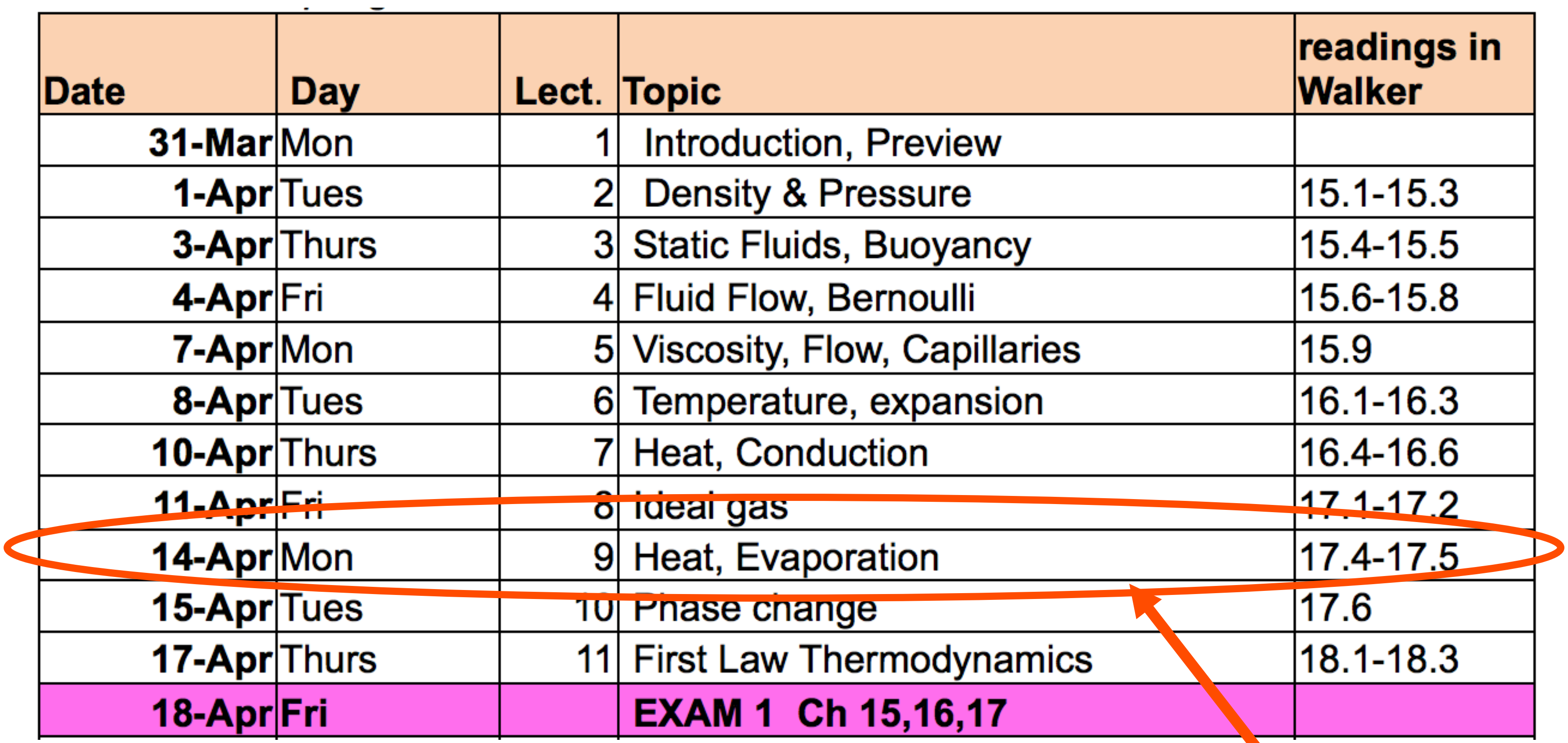
Lecture Schedule (up to exam 1)



Date	Day	Lect.	Topic	readings in Walker
31-Mar	Mon	1	Introduction, Preview	
1-Apr	Tues	2	Density & Pressure	15.1-15.3
3-Apr	Thurs	3	Static Fluids, Buoyancy	15.4-15.5
4-Apr	Fri	4	Fluid Flow, Bernoulli	15.6-15.8
7-Apr	Mon	5	Viscosity, Flow, Capillaries	15.9
8-Apr	Tues	6	Temperature, expansion	16.1-16.3
10-Apr	Thurs	7	Heat, Conduction	16.4-16.6
11-Apr	Fri	8	Ideal gas	17.1-17.2
14-Apr	Mon	9	Heat, Evaporation	17.4-17.5
15-Apr	Tues	10	Phase change	17.6
17-Apr	Thurs	11	First Law Thermodynamics	18.1-18.3
18-Apr	**Fri**		**EXAM 1 Ch 15,16,17**	

Today

Just joined the class? See course home page

courses.washington.edu/phy115a/

for course info, and slides from previous sessions

Announcements

- Prof. Jim Reid is standing in for RJW this week
- Exam 1 **this Friday 4/18**, in class, formula sheet provided
  - YOU bring a bubble sheet , pencil, calculator (NO laptops or phones; **NO personal notes allowed**.)
  - We will post sample questions tomorrow, and go over them in class Thursday
- Clicker responses from last week are posted, so you can check if your clicker is being detected. See link on class home page, http://courses.washington.edu/phy115a

Last time

Gas Law: Avogadro's number and R

Counting molecules to get N is difficult, so it is convenient to use Avagadro's number N_A, the number of carbon atoms in exactly 12 g (1 mole) of carbon. 1 mol = {molecular mass, A} grams of gas (For elements, what you see on the Periodic Table is A averaged over isotopes)

$N_A = 6.022 \times 10^{23}$ molecules/mole and $N = nN_A$, where n = number of moles of gas

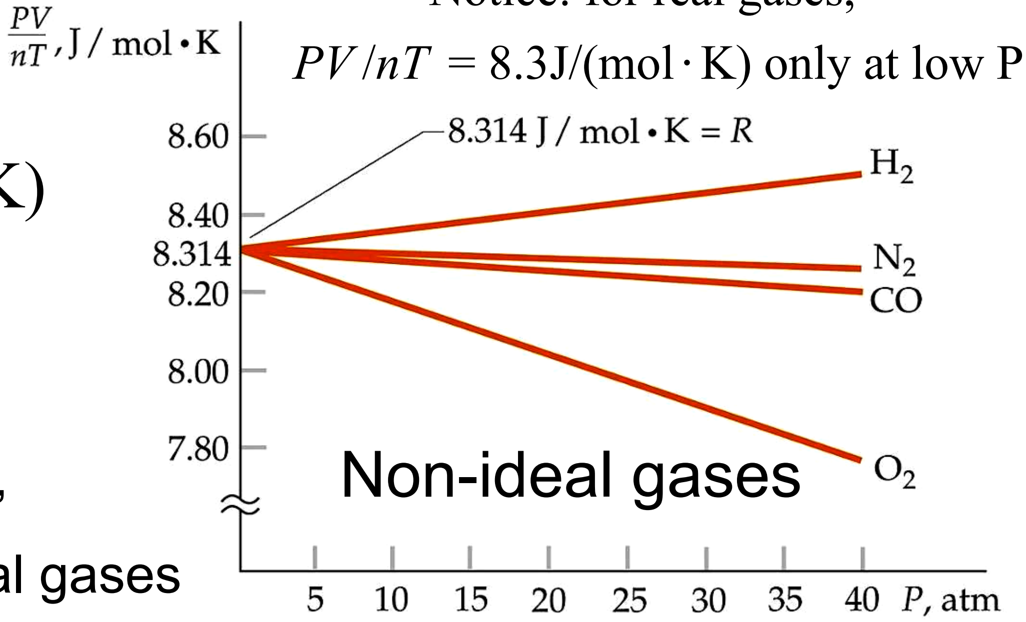
$$PV = nN_A kT = nRT$$

Notice PV = energy: N-m

$$R = N_A k = 8.314 \text{ J/(mol}\cdot\text{K)}$$

$$\boxed{PV = nRT}$$

Ideal Gas Law, in moles
R = "Universal gas constant"
Good approx at low P for real gases



Notice: for real gases, PV/nT = 8.3J/(mol·K) only at low P

Non-ideal gases

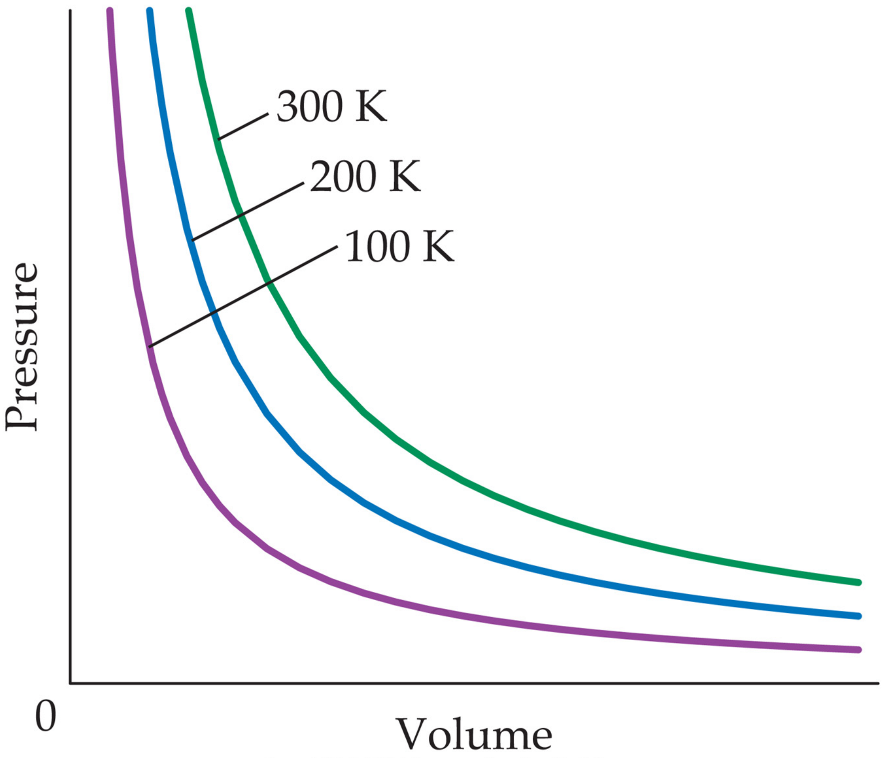
Isotherm plots

- PV=NkT results from many different observations:
  - Hold N, T constant and see how P, V vary: find

$$\boxed{\begin{array}{c} PV = const \rightarrow P = \dfrac{const}{V} \\ \text{(with T and N fixed)} \end{array}}$$

Boyle's Law

For different T's we get a set of (1/V)-shaped curves



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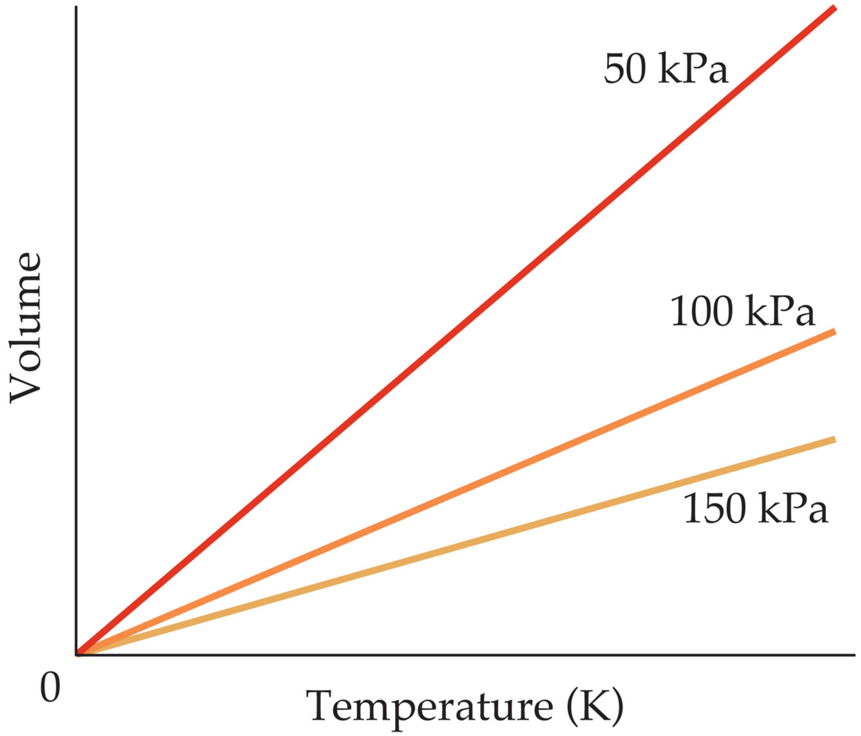
Isobar plots

- Hold N, P constant and see how V, T vary: find

$$\boxed{\begin{gathered}V=(const)T\\ \left(\text{constant } N,\ P\right)\end{gathered}}$$

Charles & Gay-Lussac Law

For different P's we get a set of linear plots



Example: Volume of an ideal gas

- What volume is occupied by 1.00 mol of an ideal gas if it is at T = 0.00°C and P = 1.00 atm?

$$PV = nRT \quad \text{so}$$

$$V = \frac{nRT}{P}$$

$$= \frac{(1.00\ \text{mol})\left[0.08206\ \text{L}\cdot\text{atm}/(\text{mol}\cdot\text{K})\right](273.15K)}{(1.00\ \text{atm})}$$

$$= 22.41\ \text{L}$$

  - If we increase the V available, with same T: P must drop
  - If we increase the T, with V kept the same: P must rise
- Standard Temperature and Pressure (STP) = 0°C, 1 atm
  - At STP, one mole of any ideal gas occupies 22.4 liters

Example: heating and compressing a gas

- An ideal gas initially has a volume = 2.00 L, temperature = 30.0°C, and pressure =1.00 atm.
- The gas is heated to 60.0°C and compressed to a volume of 1.50 L – what is its new pressure?

$$\frac{P_1V_1}{T_1} = \frac{P_2V_2}{T_2} \quad \text{so} \quad P_2 = P_1\frac{V_1T_2}{V_2T_1} = (1.00\ \text{atm})\frac{(2.0\ \text{L})(60.0°C + 273.15°C)}{(1.5\ \text{L})(30.0°C + 273.15°C)}$$

$$= 1.47\ \text{atm}$$

  - Notice: we **must** use Kelvin temperatures when applying ideal gas laws – what would result have been if we use the ratio (60/30)?

Quiz 5

- Two containers with equal V and P each hold samples of the same ideal gas. Container A has twice as many molecules as container B.
- Which is the correct statement about the absolute temperatures in containers A and B, respectively?

A. $T_A = T_B$
B. $T_A = 2\,T_B$
C. $T_A = (1/2)T_B$
D. $T_A = (1/4)\,T_B$
E. $T_A = (1/\sqrt{2})T_B$

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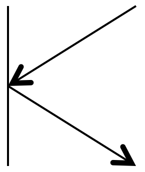
$$PV = nRT \quad \text{so} \quad T = \left(PV/nR\right) \propto \left(1/n\right)$$

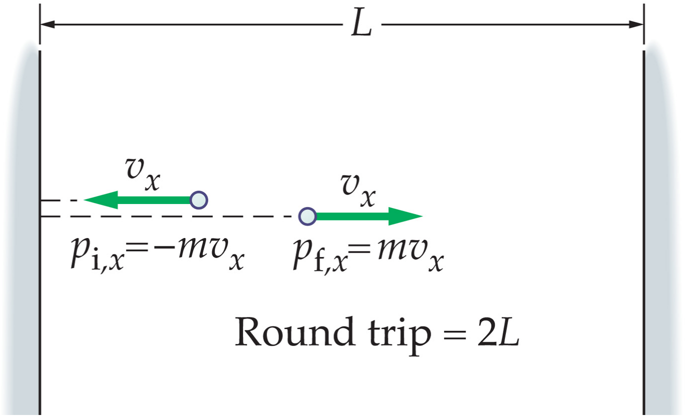
Relating gas laws to molecular motion

- P, V, T are macroscopic quantities
  - Human-scale quantities, measurable on a table-top
- Molecular motion (*x*, *v* vs *t*) = microscopic quantities
- Kinetic theory of gases: connect micro to macro
  - Model for ideal gas
    - N is large, molecules are identical point-particles
    - Molecules move randomly
    - No inelastic interactions: collisions are always elastic
      - Recall: elastic means no loss of KE due to collision

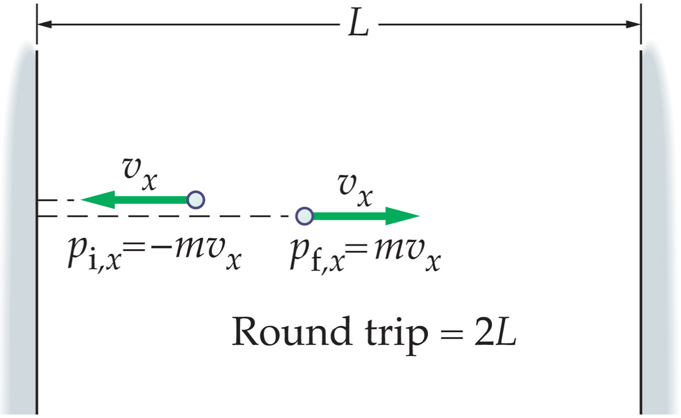
Elastic collision with wall means momentum (so, ***v***) component perpendicular to wall gets reversed



Speed unchanged
Vertical ***v*** unchanged
Horizontal ***v*** reversed



Calculate the pressure of a gas



Change in horizontal momentum of molecule

$$\Delta p_x = p_{x,f} - p_{x,i} = mv_x - (-mv_x) = 2mv_x$$

Change is due to force exerted by wall:

$$F\Delta t = \Delta p_x, \quad F_{ON\ WALL} = -F_{BY\ WALL}$$

Average force exerted on wall by one molecule $F_{AVG} = \dfrac{\Delta p_x}{\Delta t}$

where Δt = time between collisions = round-trip time

$$\Delta t = 2L/v_x \rightarrow F_{AVG} = \frac{2mv_x}{2L/v_x} = \frac{mv_x^2}{L}$$

Assume symmetrical container (LxLxL): (doesn't matter in the end)

$$P_{AVG} = \frac{F_{AVG}}{A} = \frac{1}{L^2}\left(\frac{mv_x^2}{L}\right) = \frac{mv_x^2}{V} \rightarrow \text{Adding up all molecules, } PV = N\,m\left(v_x^2\right)_{AVG}$$

Defining temperature (again): molecular scale

- Now we can connect macro to micro:

$$PV = NkT = 2N\left(\tfrac{1}{2}mv_x^2\right)_{\text{av}} \quad \rightarrow \quad \left(\tfrac{1}{2}mv_x^2\right)_{\text{av}} = \tfrac{1}{2}kT$$

Nothing special about the x-direction: random motion means

$$\left(v_x^2\right)_{\text{av}} = \left(v_y^2\right)_{\text{av}} = \left(v_z^2\right)_{\text{av}} \quad \text{and} \quad \left(v^2\right)_{\text{av}} = \left(v_x^2\right)_{\text{av}} + \left(v_y^2\right)_{\text{av}} + \left(v_z^2\right)_{\text{av}} = 3\left(v_x^2\right)_{\text{av}}$$

(because random ***v*** components are independent of one another)

The average translational kinetic energy of the molecules is:

$$\boxed{K_{\text{translational av}} = \left(\tfrac{1}{2}mv^2\right)_{\text{av}} = \tfrac{3}{2}kT \quad \text{per molecule}}$$

Deep and fundamental ! Avg KE of gas molecule is proportional to T, with Boltzmann constant as the factor

$$K_{\text{trans}} = N\left(\tfrac{1}{2}mv^2\right)_{\text{av}} = \tfrac{3}{2}NkT = \tfrac{3}{2}nRT$$

Root-mean-square (RMS) - useful avg where quantity-squared is what matters:

$$\left(v^2\right)_{\text{av}} = \frac{3kT}{m_{molecule}} = \frac{3N_AkT}{N_Am_{molecule}} = \frac{3RT}{M_{MOLE}} \quad \text{and} \quad v_{\text{RMS}} = \sqrt{\left(v^2\right)_{\text{av}}} = \sqrt{\frac{3RT}{M_{MOLE}}}$$

Example: RMS speed of gas molecules

RMS means: take each molecule's speed and square it, then find the average of those numbers, and THEN take the square root.

In practice: we find the statistical speed distribution of the molecules, and use that to estimate RMS speed

Oxygen gas (O_2) has a molar mass* M of about 32.0 g/mol, and hydrogen gas (H2) has a molar mass of about 2.00 g/mol. Assuming ideal-gas behavior, what is:

(a) the RMS speed of an oxygen molecule when the temperature is 300K (27°C), and

(b) RMS speed of a hydrogen molecule at the same temperature

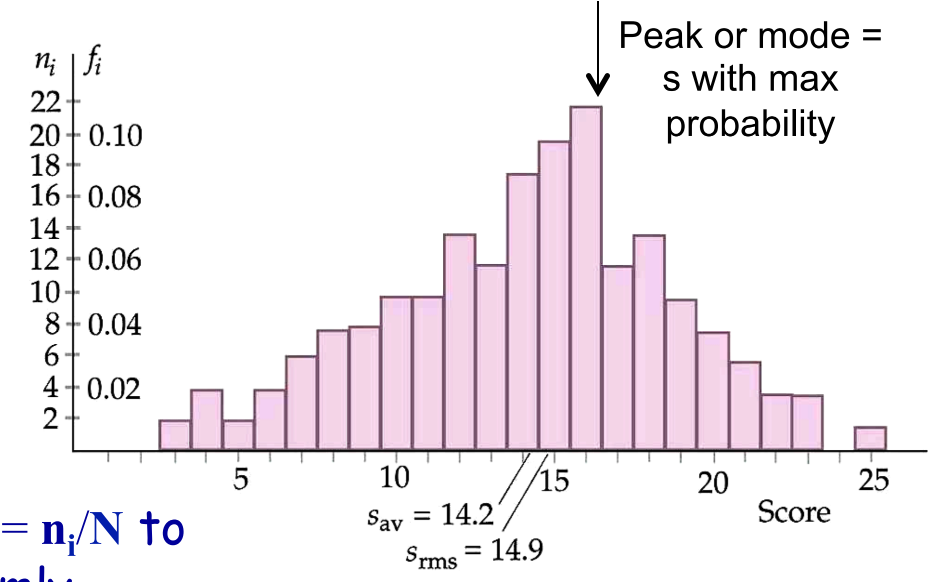
Note: Walker says "molecular mass" for molar mass – confusing.
M_X = grams in 1 mole of X,
m_X = mass (in kg) of one X molecule

$$v_{\text{O2 RMS}} = \sqrt{\frac{3RT}{M_O}} = \sqrt{\frac{3(8.314\ \text{J/mol}\cdot\text{K})(300\ \text{K})}{(0.0320\ \text{kg/mol})}} = 485\ \text{m/s}$$

$$v_{\text{H2 RMS}} = \sqrt{\frac{3RT}{M_H}} = \sqrt{\frac{3(8.314\ \text{J/mol}\cdot\text{K})(300\ \text{K})}{(0.0020\ \text{kg/mol})}} = 1{,}934\ \text{m/s}$$

Probability Distributions

We give a 25 point quiz to **N** students, and plot the results as a **histogram**, showing the number $\mathbf{n_i}$ of students, or fraction $\mathbf{f_i = n_i/N}$ of students, for each possible score vs. score, from 0 to 25.



Peak or mode = s with max probability

Such plots represent **distributions**.

For reasonably large N, we can use $\mathbf{f_i = n_i/N}$ to estimate the probability that a randomly selected student received a score $\mathbf{s_i}$.

Notice, the fractions will add to 1 for all possible scores, so that $\mathbf{\Sigma f_i = 1}$. In that case the histogram represents a **normalized distribution function**. We have the following relations:

It's not useful for class grades, but we could also calculate the average **squared** score:

$$\sum_i n_i = N \qquad \sum_i f_i = 1$$

$$\boxed{s_{\text{av}} = \frac{1}{N}\sum_i n_i\, s_i = \sum_i f_i\, s_i}$$

$$s_{\text{av}}^2 = \frac{1}{N}\sum_i n_i\, s_i^2 = \sum_i f_i\, s_i^2$$

$$\boxed{s_{RMS} = \sqrt{s_{\text{av}}^2} = \sqrt{\sum_i f_i\, s_i^2}}$$