PURDUE UNIVERSITY DEPARTMENT OF PHYSICS

Physics 22000

General Physics

Lecture 17 – Gases

Fall 2016 Semester

Prof. Matthew Jones

SUPPLEMENTAL INSTRUCTION

Free Study Sessions!

Rachel Hoagburg

Come to SI for more help in **PHYS 220**

Tuesday and Thursday	7:30-8:30PM Shreve C113

Office Hour

Tuesday 1:30-2:30 4th floor of Krach

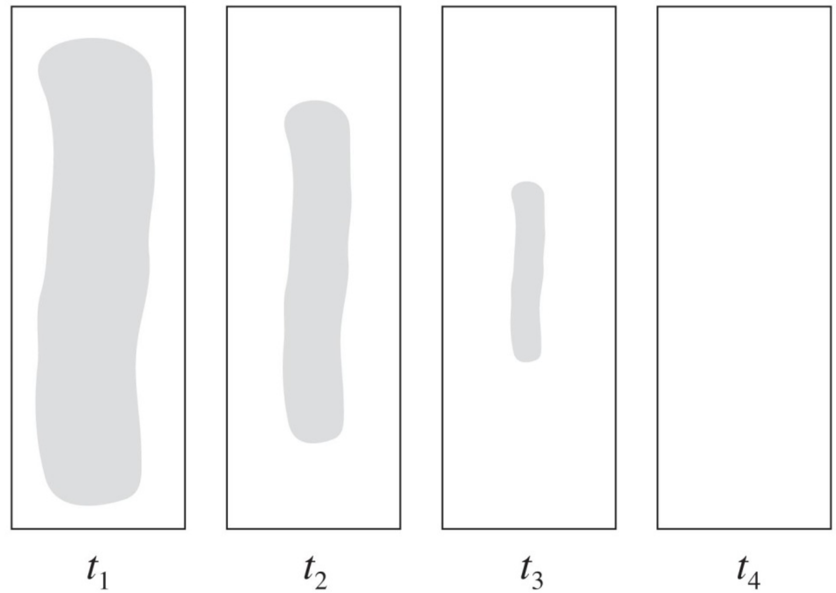
For other SI-linked courses and schedules, visit purdue.edu/si or purdue.edu/boilerguide

Structure of Matter

- Not everything around us is a rigid body…
- Do we need new laws of physics to describe things that aren't rigid bodies?
- What is "internal energy" anyway?
- What is the microscopic structure of matter?
- For things that are too small to see, we need to construct a model and see if it provides a satisfactory description of many different phenomena.

Particle Structure of Matter

- Put some alcohol on a piece of paper and watch the wet patch disappear…



- The alcohol disappears gradually, not all at once.
- You can smell the alcohol in the air from some distance away.

Particle Structure of Matter

- Based on various experiments, it is reasonable to assume that alcohol and other liquids are composed of smaller objects, called particles, that move randomly in all directions.
  - These particles need empty space between them so that particles of other materials can move between them.
  - This model of the internal structure of alcohol can be used to explain many other phenomena that we encounter.

Particle Structure of Matter

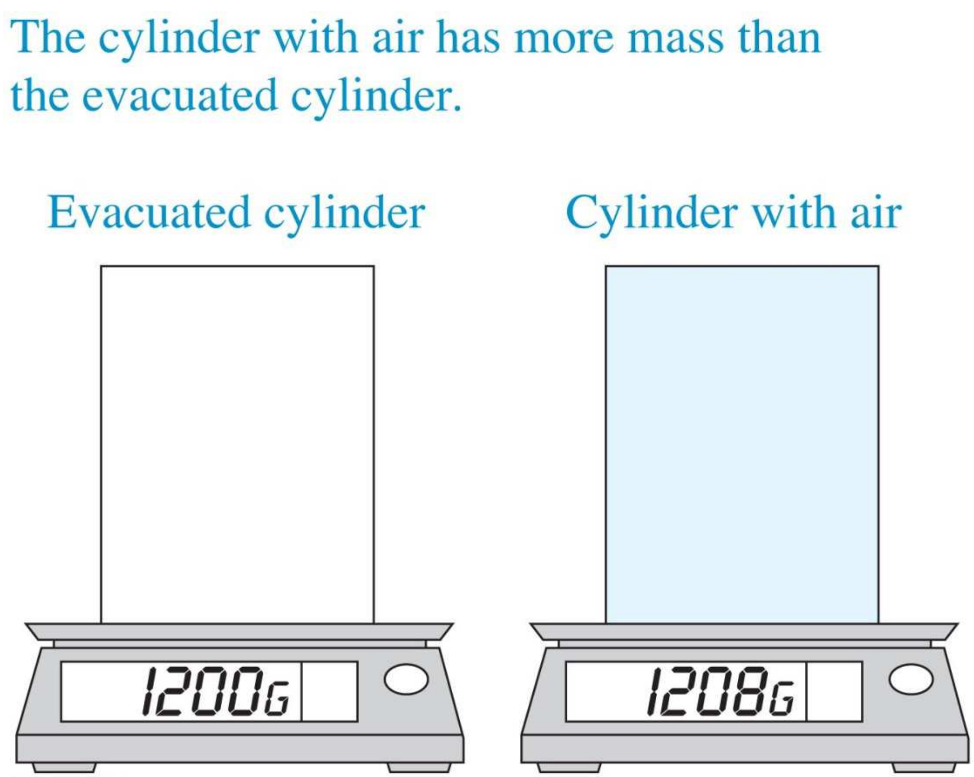
- Suppose we open a bottle of gasoline. Several minutes later, people all over the room can smell the gasoline.
- According to the particle model, the particles of gasoline leave the bottle and gradually disperse, eventually arriving at our nostrils.
- Because everyone in the room eventually smells the gasoline, its particles must move in all directions.

Air Has Mass

- We don’t usually notice that air has mass because we are surrounded by it which produces a buoyant force.
- Archimedes' principle: *the buoyant force is the same as the weight of the displaced material*.
- If we have a certain volume of air, the buoyant force is exactly equal to the weight of the air it displaces.
  - Net force is zero, so we can’t measure the mass of air this way.
  - We need to eliminate the buoyant force somehow.

Testing the Hypothesis that Air has Mass

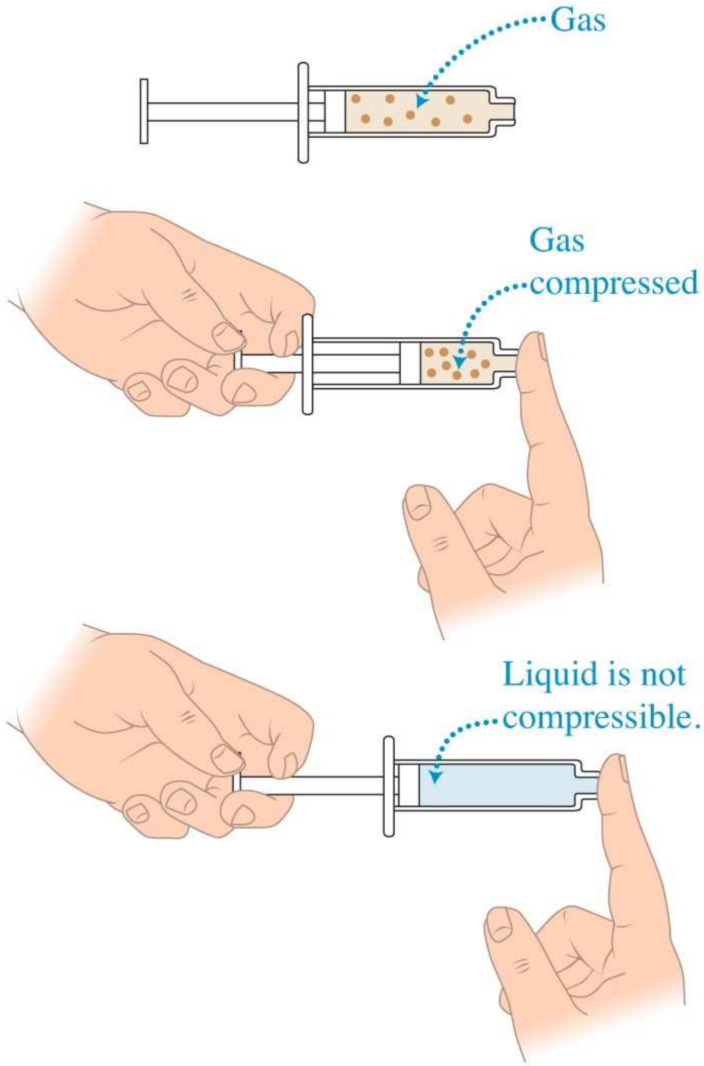
- Take two rigid metal cylinders—one from which the air has been evacuated with a vacuum pump, and the other filled with air. Weigh the cylinders.
- If the hypothesis that air has mass is correct, the evacuated jar should weigh less than the unevacuated jar.



Gases, Liquids and Solids

- Gases are easy to compress, whereas liquids and solids are almost incompressible.

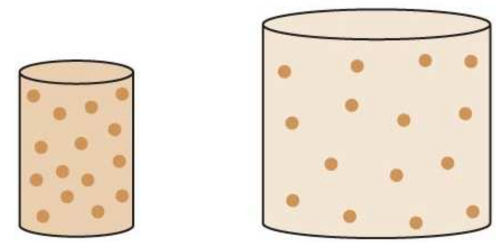
- The particle model helps us explain this difference: the amount of empty space between the particles is different in solids, liquids, and gases.



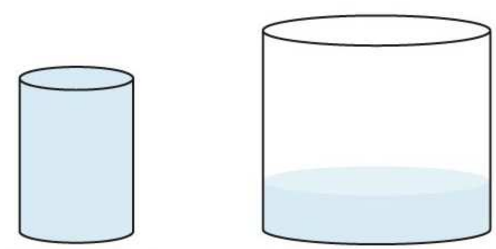
Gases, Liquids and Solids

- Gases tend to occupy whatever volume is available.
- In contrast, if we move the liquid filling a small container to a much larger container, the liquid volume remains the same independent of the container's shape.
- Solids maintain not only their volume but also their shape.

The same gas completely fills a different volume.

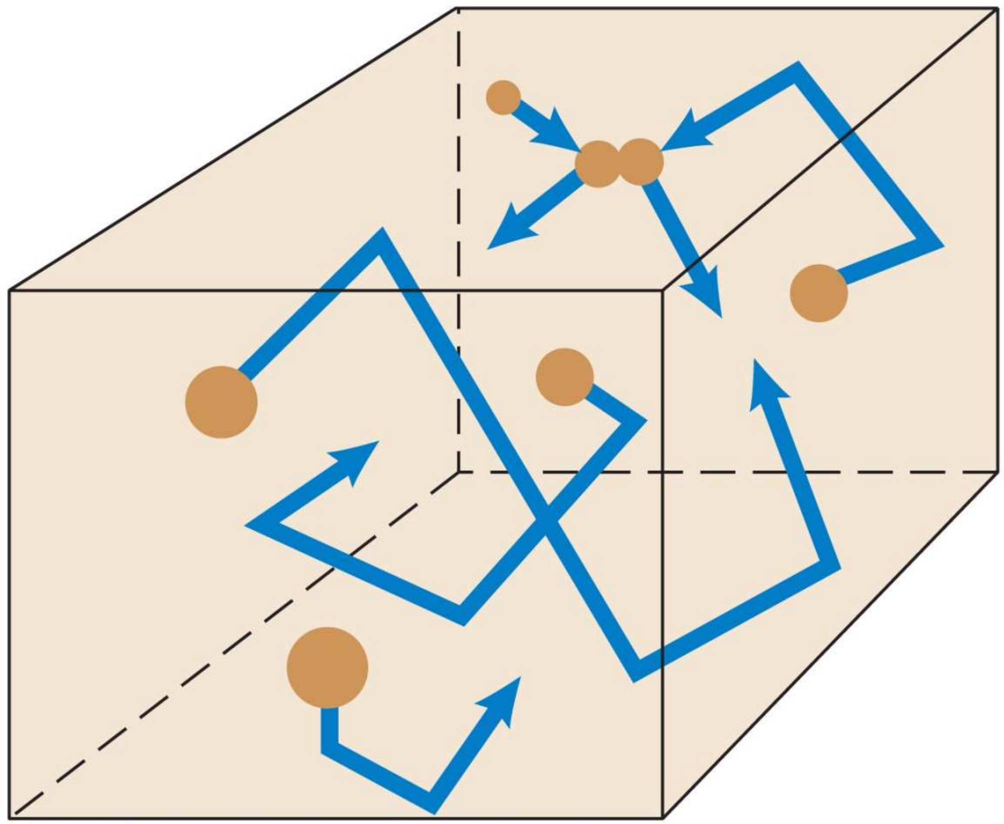


The liquid volume remains the same regardless of the container.



Ideal Gas Model

Ideal gas model A model of a system in which gas particles are considered point-like and only interact with each other and the walls of their container through collisions. This model also assumes that the particles and their interactions are accurately described using Newton's laws.

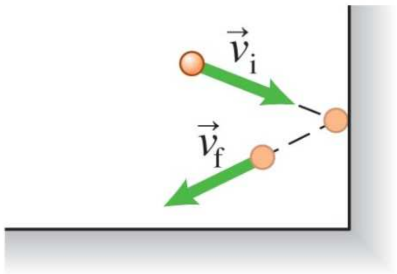


Ideal Gas Model

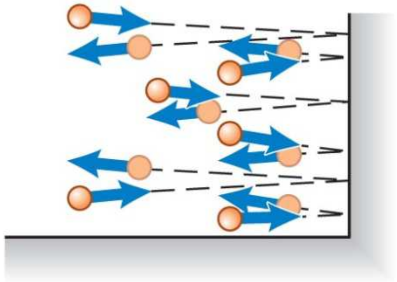
- *Ideal* in this context does not mean perfect, but rather simplified. The ideal gas model is a simplified model with certain assumptions.
  - Whether this model can be used to represent a real gas needs to be determined via testing experiments.
  - We need physical quantities to represent the features and behavior of the model. Then we can use it to develop descriptions, explanations, and predictions of new phenomena.

Pressure

- As air particles move randomly in space, they eventually collide with the solid surfaces of any objects in that space. In each of these collisions, the particle exerts an impulsive force on the object—like a tennis ball hitting a practice wall.



Each ball or particle exerts an impulsive force on the wall during a collision.

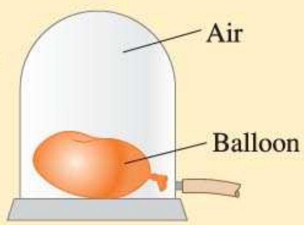
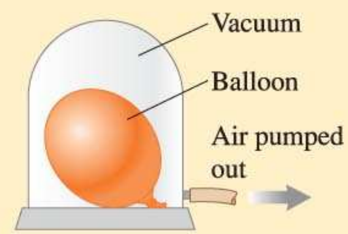


Many particles hitting the wall cause a near-constant force.

Testing the model of moving gas particles pushing on the surface

TESTING EXPERIMENT TABLE

9.2 Testing the model of moving gas particles pushing on the surface.

Testing experiment	Prediction based on a model of gas particle motion	Outcome
Place a partially inflated balloon inside a vacuum jar. Seal the jar.  What happens to the balloon's shape when you start pumping air out of the jar?	As we remove air particles from outside the balloon, the collisions of particles on the outside of the balloon are less frequent and exert less force on each part of the balloon's outer surface. The collision rate of the particles inside the balloon does not decrease. Therefore, the balloon should expand.	As air outside of the balloon is removed, the balloon expands. 

Conclusion
The model of air consisting of moving particles colliding with objects exposed to the air has not been disproved. The results of this experiment support the model.

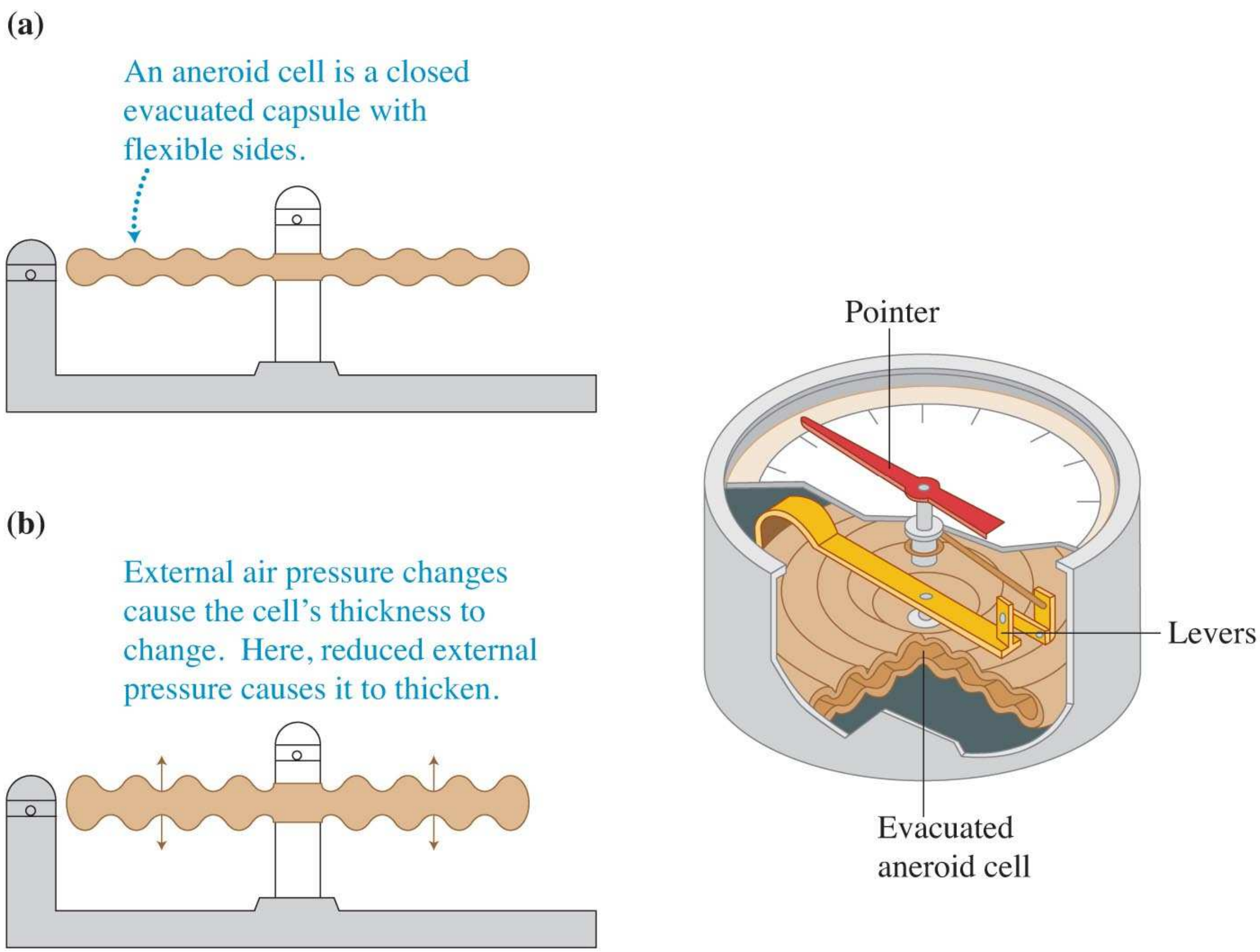
Pressure

Pressure P Pressure is a physical quantity equal to the magnitude of the perpendicular component of the force $F_\perp$ that a gas, liquid, or solid exerts on a surface divided by the contact area A over which that force is exerted:

$$P = \frac{F_\perp}{A} \qquad (9.1)$$

- The SI unit of pressure is the pascal (Pa), where 1 Pa = 1 N/m^2.
- Atmospheric pressure is 1.013 x 10^5 Pa
  – At least that's what it is at sea level…
  – It's less at high altitudes

Measuring Pressure



Gauge Pressure

Gauge pressure P_{gauge} Gauge pressure is the difference between the pressure in some container and the atmospheric pressure outside the container:

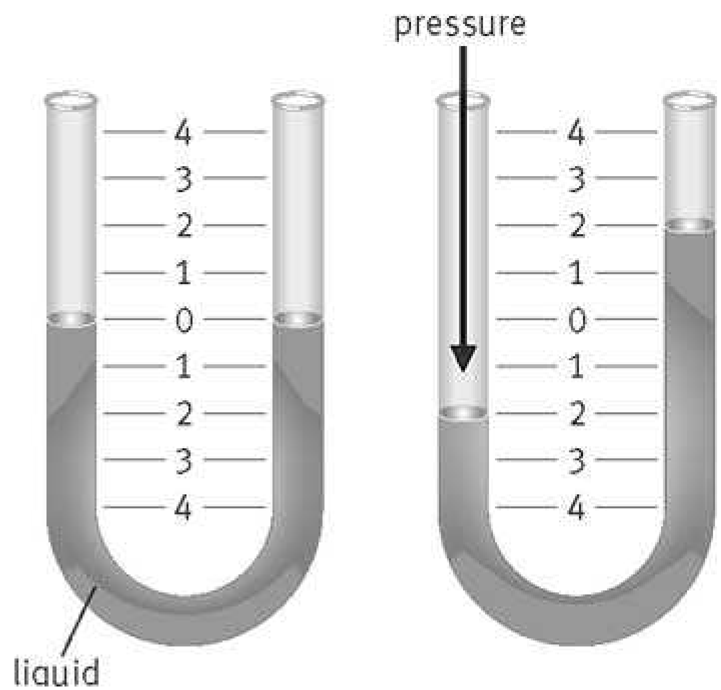
$$P_{gauge} = P - P_{atm} \tag{9.2}$$

where $P_{atm} = 1.0\,\text{atm} = 1.0 \times 10^5\,\text{N/m}^2$.

- When you use a tire gauge to measure the air pressure in a car tire, you are comparing the pressure inside the tire to the pressure of the atmosphere outside the tire.

Gauge and Absolute Pressure

- 14.7 pounds per square inch (psi) is about 101.3 kPa.
- Other ways to measure pressure:



Other Units for Pressure

- Mercury has a density of 13.56 grams/cm^3
- A column of mercury 760 mm tall exerts a pressure

$$
\begin{aligned}
P &= \rho g h \\
&= (13560\, kg/m^3)(9.8\, N/kg)(0.76\, m) \\
&= 1.01 \times 10^5\, N/m^2
\end{aligned}
$$

- Pressure is sometimes measured in mmHg
- You could also use water…

Density

- For gases, a much more useful physical quantity than mass of the individual particles is the mass of one unit of volume—density.

Density ρ The density ρ (lowercase Greek letter "rho") of a substance or of an object equals the ratio of the mass m of a volume V of the substance (for example, air or water) divided by that volume V:

$$\rho = \frac{m}{V} \tag{9.3}$$

The unit of density is kg/m^3.

TIP Density is different from mass. Air in a room has a particular mass and density. If you divide the room into two equal parts using a screen, the mass of air in each part will be half the total mass, but the density in each part will remain the same.

Mass and Sizes of Particles

- In 1811, an Italian scientist named Avogadro proposed that equal volumes of different types of gas, when at the same temperature and pressure, contain the same number of gas particles.
- The mass in grams of any substance that has exactly Avogadro's number of particles is equal to the atomic mass of that substance.
- This is actually the *definition* of “atomic mass”.
- One mole of gas particles at 0° C and at atmospheric pressure, occupies 22.4 Liters.

Avagadro's Number and the Mole

Avogadro's number and the mole Avogadro's number $N_A = 6.02 \times 10^{23}$ particles is called a mole. The number of particles in a mole is the same for all substances and is the number of particles whose total mass equals the atomic mass of that substance.

- Twelve eggs is a *dozen*.
- 144 eggs is a *gross*.
- 13 doughnuts is a *baker's dozen*.
- Twenty somethings is a *score*.
- **6.02×10^{23} particles is a *mole*.**

- English is weird.

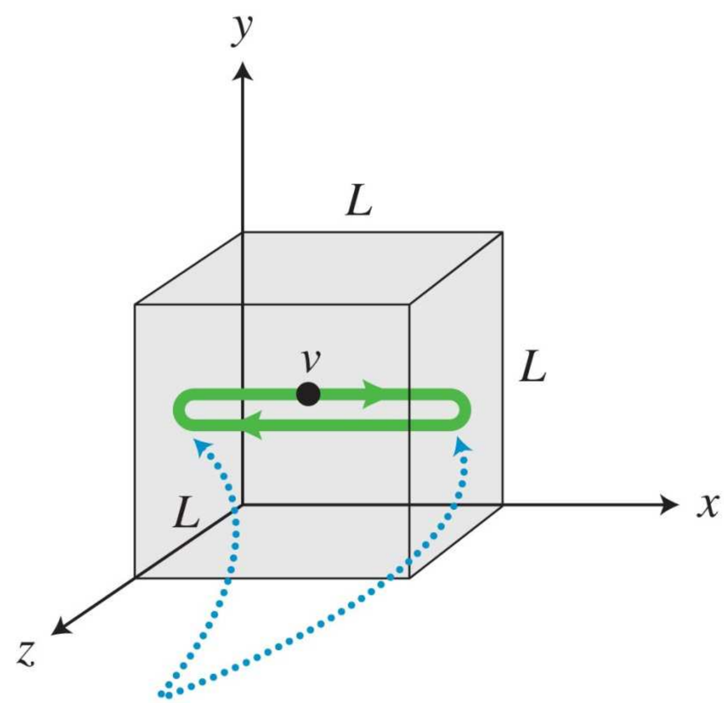
Quantitative Analysis of an Ideal Gas

- We need more simplifying assumptions.
  - Assume that the particles do not collide with each other—they collide only with the walls of the container, exerting pressure on the walls. This is a reasonable assumption for a gas of low density.
  - Assume that the collisions of particles with the walls are elastic. This makes sense, because the pressure of the gas in a closed container remains constant, which would not happen if the particles' kinetic energy decreased.

Quantitative Analysis of an Ideal Gas

- Imagine a gas particle moving inside a cubic container with sides of length L.

- The wall exerts a force on the particle, and the particle in turn exerts an equal-magnitude and oppositely directed force on the wall.



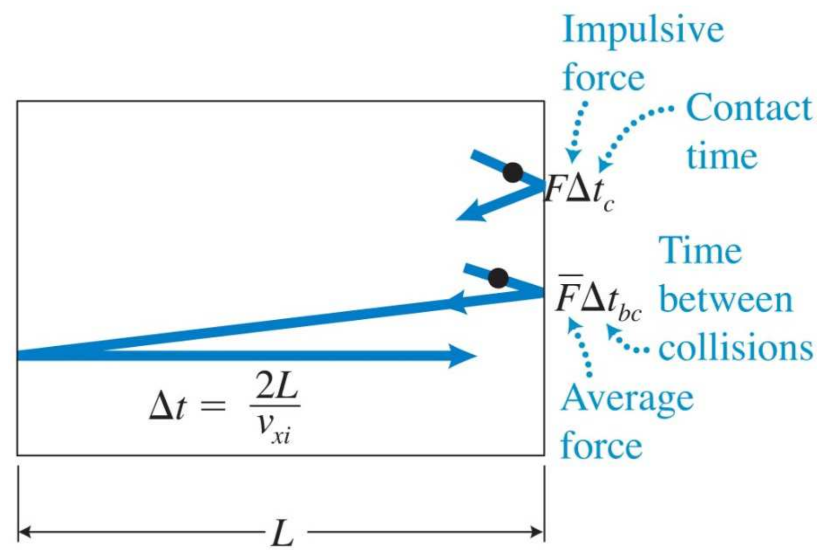
Particle speed is the same before and after these elastic collisions.

Quantitative Analysis of an Ideal Gas

- When particles have high speed they

1. Hit the sides of the container more frequently.
2. Exert a greater force during each collision.

- Both factors lead to a greater pressure.
- Thus it is the speed squared—not just the speed of the particles—that affects the pressure.



The impulsive force exerted by a particle against the wall during the short contact time interval equals the average force exerted by the particle against the wall during the long time interval between collisions: $F\Delta t_c = \bar{F}\Delta t_{bc}$.

Quantitative Analysis of an Ideal Gas

- In the text, it is shown that
$$Nm\overline{v_x^2} = PL^3 = PV$$
- The average velocity-squared in the x-direction should just be 1/3 of the total average velocity squared:
$$\overline{v^2} = \overline{v_x^2} + \overline{v_y^2} + \overline{v_z^2}$$
- Nothing special about the x-axis, so
$$Nm\overline{v^2} = 3PV$$

Boyle's Law

- This result is consistent with an experimental observation (Boyle, 1662).
- When temperature is constant, the volume is inversely proportional to the pressure.

$$V \propto \frac{1}{P}$$

$$PV = const.$$

Quantitative Exercise

- Estimate the average speed of air particles at normal conditions where the air is at a pressure of 10^5 N/m^2 and 1 mole of air particles occupies 22.4 L = 22.4x10^{-3} m^3.

$$\bar{v} \approx \sqrt{\frac{3PV}{Nm_p}}$$

Inserting the appropriate values gives

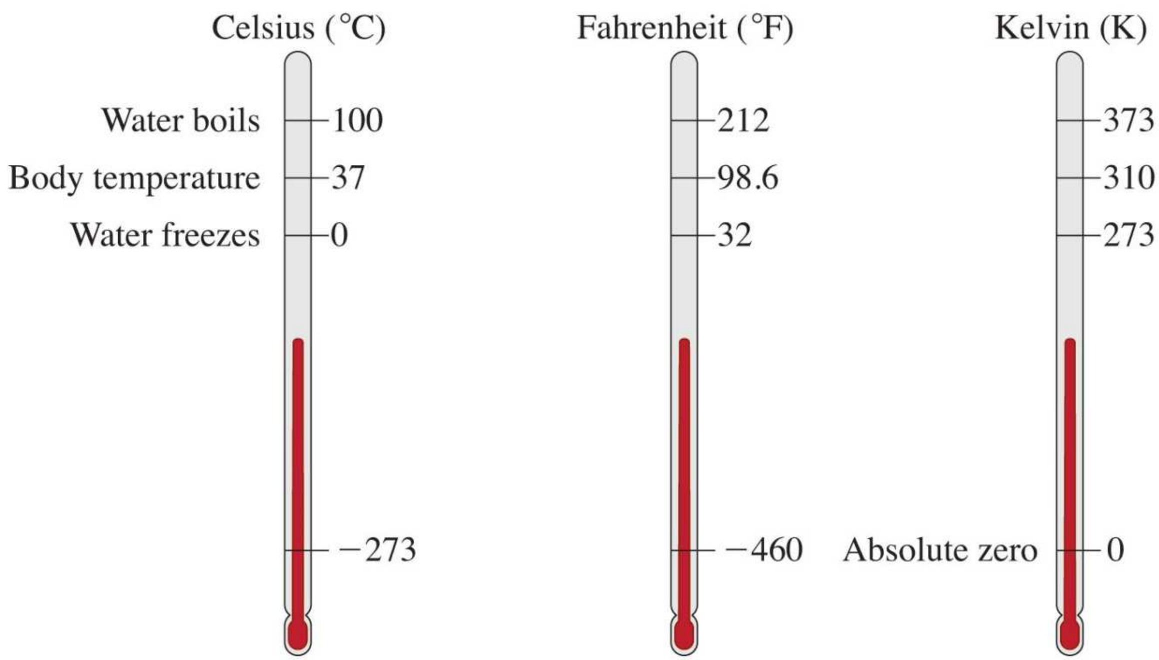
$$\bar{v} \approx \sqrt{\frac{3PV}{Nm_p}} = \sqrt{\frac{3(1.0 \times 10^5\,\mathrm{N/m^2})(22.4 \times 10^{-3}\,\mathrm{m^3})}{(6.02 \times 10^{23})(4.8 \times 10^{-26}\,\mathrm{kg})}}$$
$$= 480\,\mathrm{m/s}$$

Time Interval Between Collisions

- Why does it take 5 to 10 minutes for the smell of gasoline to travel across a room if the average speed of air particles is v = 480 m/s?
  - The average distance between particles in a gas is about $D = 3.3 \times 10^{-7}$ cm.
  - Perhaps we cannot assume that the gas particle collisions can be ignored—in fact, they collide about 109 times per second.
  - Even though the gas particles are moving fast, their migration from one place to another is slow.

Temperature

- Temperature can be measured using a liquid thermometer:



- The bulb and part of the tube are filled with a red liquid that expands predictably when heated and that shrinks when cooled.

Charles's Law

- When the pressure is constant, the volume of a gas is proportional to the absolute temperature.
- In an experiment, we can hypothesize that the volume of gas expands because the impulses of the particles against the inside walls are larger when the gas is warm.
  - This would happen if the particles were moving faster.
  - Based on this reasoning, we can hypothesize that the temperature of a gas is related to the speed of the random motion of its particles.

The Ideal Gas Law

- Boyle's law:

$$V \propto 1/P$$

- Charles's law:

$$V \propto T$$

- Law of the additive nature of stuff:

$$V \propto N$$

- Combining these three concepts:

$$\frac{PV}{N} = kT \quad \text{(Ideal gas law)}$$

Finding the constant "k"

- Measure P of one mole of gas in a fixed volume at two different temperatures:

Table 9.4 ***PV/N* for one mole of gas in a 22.4-L container at two different temperatures.**

Conditions in the bath	Pressure	Volume	$\frac{PV}{N} = kT$
Ice water (T)	1.013×10^5 N/m^2	22.42×10^{-3} m^3	3.773×10^{-21} J
Boiling water ($T + 100$)	1.384×10^5 N/m^2	22.42×10^{-3} m^3	5.154×10^{-21} J

- We have two equations with two unknowns: the constant *k* and the water temperature *T*.
- We subtract the first equation from the second to get

$$k = 1.38 \times 10^{-23}\ J/\text{degree}$$

Absolute Temperature Scale

- We need a scale on which the zero point is the lowest possible temperature. That way, all temperatures will be positive.
- The lowest possible temperature on the new scale is 0; on the Celsius scale, it would be –270 °C.
- This temperature scale is called the absolute temperature scale or the Kelvin scale. It was invented by William Thomson, Lord Kelvin, in 1848.

Ideal Gas Law

$$PV = NkT$$

- Because one mole has Avogadro's number of particles, $N = nN_A$, we can write:

$$PV = nN_AkT$$

where n is the number of moles of gas.

> **TIP** Note that Eq. (9.7) implies that when the absolute temperature of the ideal gas is zero, its pressure must be zero.

Ideal Gas Law

Ideal gas law For an ideal gas, the quantities pressure P, volume V, number of particles N, temperature T (in kelvins), and Boltzmann's constant $k = 1.38 \times 10^{-23}$ J/K are related in the following way:

$$PV = NkT \tag{9.7}$$

The law can also be written in terms of the number of moles of particles n, and the universal gas constant $R = 8.3 \frac{\text{J}}{\text{K} \cdot \text{mole}}$:

$$PV = nRT \tag{9.8}$$