

Chapter 5: Gases

Gases: fluid or rigid? *fast random motion of particles*
compressible or incompressible? *molecules spread apart.*

Pressure: force per unit area

$$P = \frac{F}{A}$$

psi can be relative pressure (1 atm = 0 psi)

Units of pressure:

absolute pressures: 0 = no pressure

atmosphere (atm): 1 atm = average atmospheric pressure at sea level

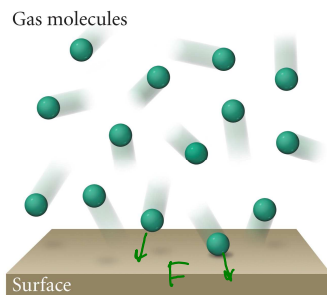
= 760 torr

millimeters mercury (mmHg): 1 atm = 760 mmHg

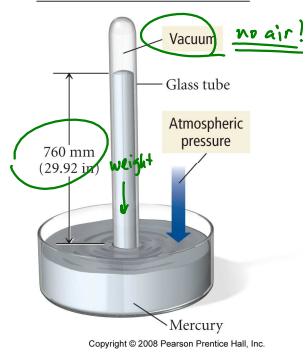
1 mmHg = 1 torr

pascal (Pa): SI unit of pressure
1 atm = 101,325 Pa

manometer: measures pressure of gas sample



The Mercury Barometer



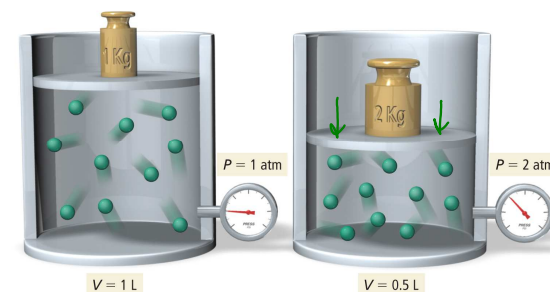
Simple gas laws

from experiment

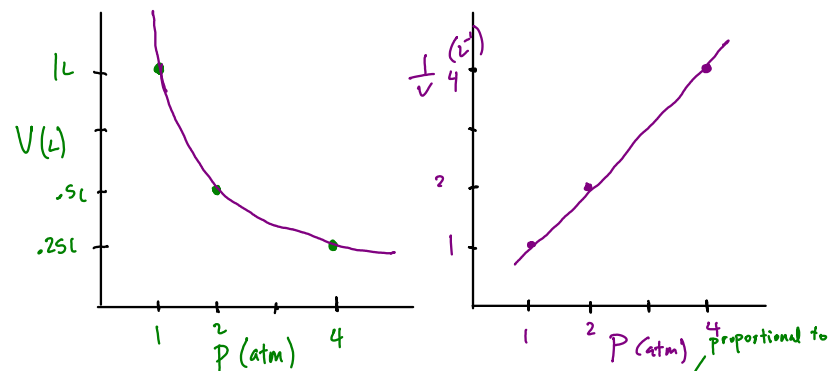
Simple gas laws (sometimes called empirical gas laws): relationships between pressure (P), volume (V), absolute temperature (T), and amount in moles (n)

Boyle's Law: relationship of P and V , and the compressibility of gases

Volume versus Pressure: A Molecular View



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Boyle's law: P and V are inversely proportional
 $P \times V = \text{constant}$
 $P \propto \frac{1}{V}$
 $V \propto \frac{1}{P}$

Boyle's law

$$V = \text{constant} \times \frac{1}{P}$$

$$\begin{array}{c} \uparrow \\ PV = \text{constant} \\ \downarrow \end{array}$$

$$\boxed{P_1 V_1 = P_2 V_2} (= \text{constant})$$

6.2 L air at 760 mmHg is compressed to 4.4 L. What is the new pressure (assuming constant temperature)?