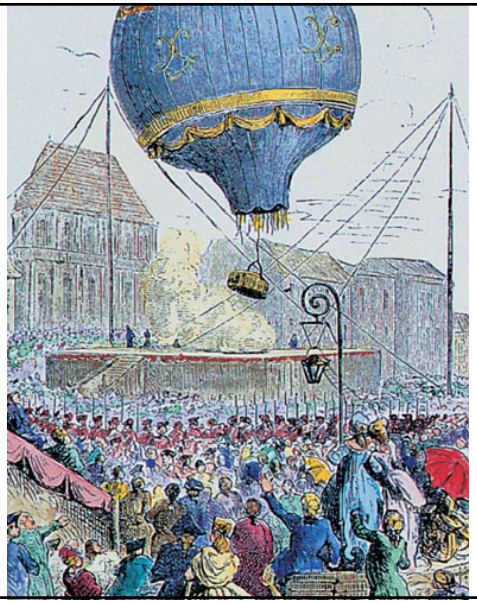


Gases and the Kinetic Molecular Theory

– Importance in atmospheric phenomena, gas phase reactions, combustion engines, etc.

5.1 The Physical States of Matter

- The condensed states – liquid and solid
- The gaseous state
  - – Gas volume changes greatly with pressure
  - – Gas volume changes greatly with temperature
  - – Gases have low viscosity (flow easily)
  - – Gases have low density (~1000 times lower than liquids and solids)
  - – Gases are miscible in all proportions

- Molecular model of the gaseous state
  - – Molecules are in constant, rapid, random motion (explains the absence of definite shape, miscibility, low viscosity)
  - – Molecules are widely separated (explains the absence of definite volume, low density, compressibility)

5.2 Pressure

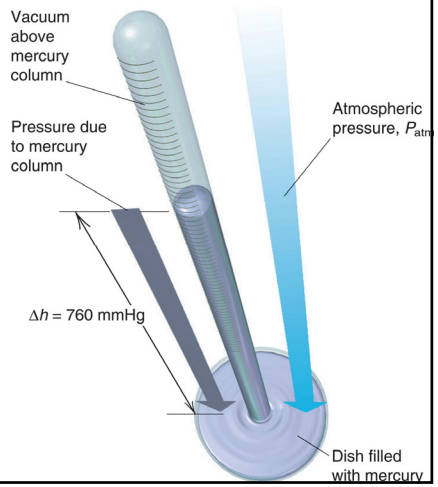
- Gas molecules collide with each other an the walls of the container → molecules exert force on the walls

$$\text{Pressure} = \frac{\text{Force}}{\text{Area}}$$

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

- **Atmospheric pressure** (P_{atm}) – caused by the gravitational pull of the Earth → molecules exert force on all objects

➢ **Barometers** - measure the atmospheric pressure (Torricelli)
  - – no pressure above the mercury column (vacuum)
  - – the weight of the Hg column balances P_{atm}
  - – the height of the Hg column is proportional to P_{atm}



g - acceleration of free fall (**9.81 m/s^2**)
d - density of Hg (**13546 kg/m^3**)
V - volume of Hg column
h - height of Hg column (**0.760 m** at sea level)
A - area of Hg column base

$$P_{atm} = P_{Hg} = \frac{F}{A} \qquad F = m_{Hg} g$$

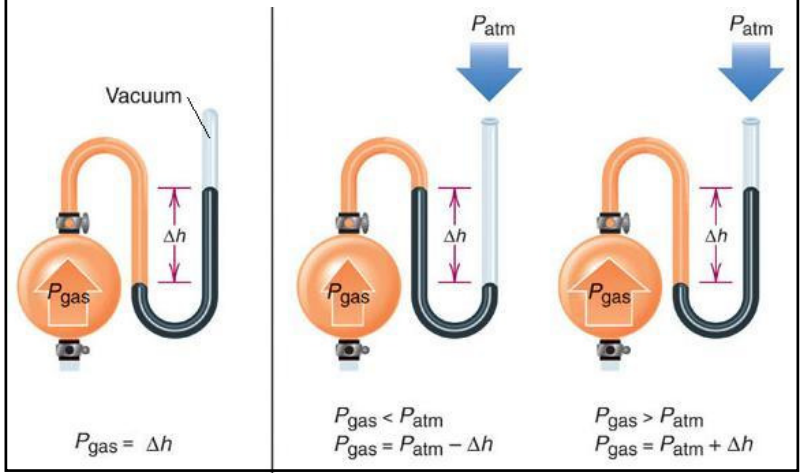
$$m_{Hg} = dV = dhA \qquad \Rightarrow \qquad F = dhAg$$

$$\Rightarrow P_{atm} = \frac{dhAg}{A} = dhg$$

$$P_{atm} = 13546 \times 0.760 \times 9.81 = 1.01 \times 10^5 \,\text{kg/m}\cdot\text{s}^2$$

➢ **Manometers** – measure the pressure of gases in containers

- **Close-end** and **open-end** manometers



Units of Pressure

- SI unit → pascal (Pa)

  1 Pa = 1 N/m^2 = 1 kg/m·s^2

  P$_{atm}$ **= 1.01×10^5 Pa** (at sea level)

- Conventional units:

  atm → **1 atm = 101325 Pa** (exactly)
  bar → **1 bar = 100000 Pa** (exactly)
  torr → **760 Torr = 1 atm** (exactly)
  mm Hg → **1 mmHg = 1 Torr**
  lb/in^2 → **14.7 lb/in^2 = 1 atm**

Example:

Convert **630.0 Torr** to atmospheres and kilopascals.

$$\textbf{630.0 Torr} \times \left(\frac{\textbf{1 atm}}{\textbf{760 Torr}} \right) = \textbf{0.8289 atm}$$

$$\textbf{630.0 Torr} \times \left(\frac{\textbf{1 atm}}{\textbf{760 Torr}} \right) \times \left(\frac{\textbf{101325 Pa}}{\textbf{1 atm}} \right)$$

$$= \textbf{8.399} \times \textbf{10}^{\textbf{4}} \ \textbf{Pa} = \textbf{83.99 kPa}$$

5.3 The Gas Laws

- Relate the parameters of the gaseous state – pressure, volume, temperature, and number of moles

Boyle's Law

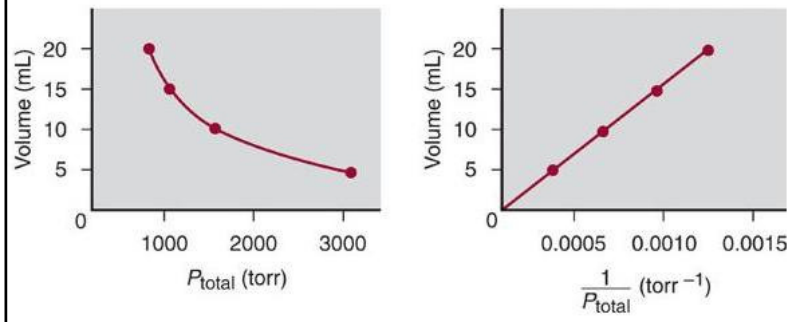
- At constant temperature (T) the pressure (P) of a fixed amount of gas is inversely proportional to its volume (V)

  $\Rightarrow$At constant T and n:

$$P = \frac{k}{V} \qquad k \rightarrow \text{constant (depends on } T \text{ and } n\text{)}$$

$$PV = k = \text{constant} \qquad V \downarrow \Leftrightarrow P \uparrow$$

V (mL)	P (torr)	$\frac{1}{P}$	PV (torr • mL)
20.0	780	0.00128	1.56×10^4
15.0	1038	0.000963	1.56×10^4
10.0	1560	0.000641	1.56×10^4
5.0	3112	0.000321	1.56×10^4



- Assume two states of a gas at constant T
  - state 1 $\rightarrow P_1, V_1$
  - state 2 $\rightarrow P_2, V_2$

$$P_1V_1 = k \qquad P_2V_2 = k$$

$$P_1V_1 = P_2V_2$$

Example: A 2.0 L sample of oxygen at 10 atm is transferred to a 15.0 L container at constant temperature. What is the new pressure?

$V_1 = 2.0$ L $\quad P_1 = 10$ atm $\quad V_2 = 15.0$ L $\quad P_2 = ?$

$$P_2 = \frac{P_1V_1}{V_2} = \frac{10 \text{ atm} \times 2.0 \text{ L}}{15.0 \text{ L}} = 1.3 \text{ atm}$$