Chapter 5
Gases

Outline

- Introduction
- Pressure: definition, measurement and units
- Gas laws: Boyle, Charles, and Avogadro
- The state of the gas and the Ideal gas equation
- Gas stoichiometry, molar volume and molar mass
- Dalton's law of partial pressure
- Kinetic molecular theory
- Effusion and diffusion

Sections 5.8 and 5.9 are omitted

A Gas

- Uniformly fills any container.
- Mixes completely with any other gas
- Exerts pressure on its surroundings.

Pressure is force/unit area

SI units = Newton/meter^2 = 1 Pascal (Pa)

1 standard atmosphere = 101,325 Pa

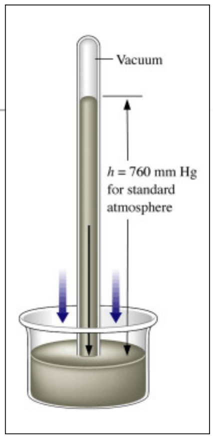
1 standard atmosphere = 1 atm =760 mm Hg = 760 torr

Other units: bar, lb/square inch, kg/meter square

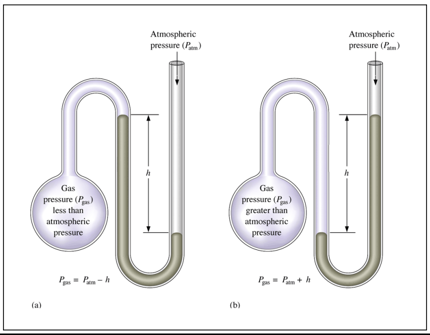
Figure 5.01b:
As the can cools, the water vapor condenses, lowering the gas pressure inside the can. This causes the can to crumple (b). (cont'd)

Measurement:

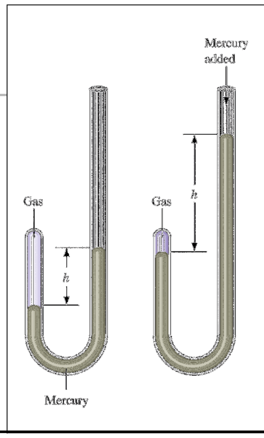
A- Barometer (Torricelli) . The tube, completely filled with mercury, is inverted in a dish of mercury.



b- open end manometer.



C- closed end manometer (J-tube)
D- Pressure gauges



Gas laws

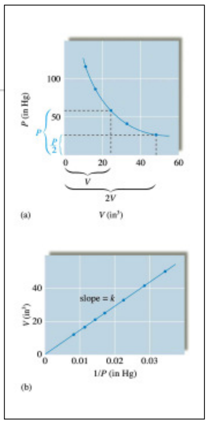
- Boyle's law: At constant temperature and amount of gas, the volume is inversely proportional to pressure
  $V \propto 1/P \quad (T = \text{constant})$

Pressure × Volume = Constant $\quad (T = \text{constant})$

$$P_1V_1 = P_2V_2 \quad (T = \text{constant})$$

(*Holds *precisely* only at very low pressures.)
The gas that obeys Boyles law is called ideal gas

Figure 5.5:
Plotting Boyle's data from Table 5.1. (a) A plot of P versus V shows that the volume doubles as the pressure is halved. (b) A plot of V versus $1/P$ gives a straight line. The slope of this line equals the value of the constant k.



Gas laws - continued

- Charles's law:

At constant pressure and amount, the volume of a gas is directly proportional to temperature.

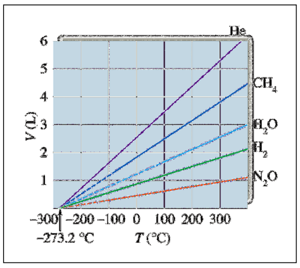
Extrapolates to zero at zero Kelvin (-273.15 ^{o}C).

$$V = bT \quad (P = \text{constant})$$

$$V_i / T_i = V_f / T_f$$

b = a proportionality constant

Figure 5.8: Plots of V versus T (ºC) for several gases.



Gas laws - continues

- Avogadro's law:

For a gas at constant temperature and pressure, the volume is directly proportional to the number of moles of gas (at low pressures).

$$V = an$$

a = proportionality constant
V = volume of the gas
n = number of moles of gas

Ideal gas law

- An equation of state for a gas.
- "state" is the condition of the gas at a given time.

$$PV = nRT$$

R = proportionality constant
= 0.08206 L atm K^{-1} mol^{-1}
P = pressure in atm
V = volume in liters
n = moles
T = temperature in Kelvin
Holds closely at P < 1 atm (Solve 5.6 – 5.10)

Gas stoichiometry

- Standard Temperature and Pressure "STP"

  P = 1 atmosphere

  $T \simeq$ □C

  The molar volume of an ideal gas is 22.42 liters at STP

We can use this relation with the chemical equation to calculate the volume of the gas produced and consumed at STP

Solve 5.11, 5.12, 5.13

Molar mass and gas density (M = dRT/P)

Solve 5.14

Mixtures of gases
Dalton's law of partial pressures

For a mixture of gases in a container,

$$P_{\text{Total}} = P_1 + P_2 + P_3 + \ldots$$

P_i is called partial pressure of gas "I"

Define mole fraction (x_i) = $n_i / n_{tot} = p_i / p_{tot}$

Solve 5.15, 5.16, 5.17

Application: Collection of gases over water

Solve 5.18

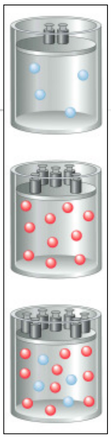
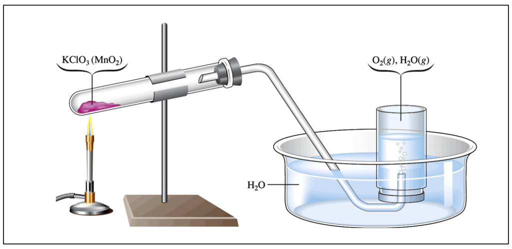


Figure 5.12:
The partial pressure of each gas in a mixture of gases in a container depends on the number of moles of that gas.

Figure 5.13: The production of oxygen by thermal decomposition of $KClO_3$. The MnO_2 is mixed with the $KClO_3$ to make the reaction faster.



Kinetic Molecular Theory

1. Volume of individual particles is ≈ zero.
2. Collisions of particles with container walls cause pressure exerted by gas.
3. Particles exert no forces on each other.
4. Average kinetic energy ∝ Kelvin temperature of a gas.

The Meaning of Temperature

$$(\mathrm{KE})_{\mathrm{avg}} = \frac{3}{2} RT$$

Kelvin temperature is an index of the random motions of gas particles (higher T means greater motion.)

Root mean square velocity (U_{rms}) = $\sqrt{3RT/M}$

Solve 5.18

Effusion and diffusion

- Diffusion: describes the mixing of gases. The rate of diffusion is the rate of gas mixing.

$$\frac{\text{Distance traveled by gas 1}}{\text{Distance traveled by gas 2}} = \frac{\sqrt{M_2}}{\sqrt{M_1}}$$

Effusion: describes the passage of gas into an evacuated chamber.

$$\frac{\text{Rate of effusion for gas 1}}{\text{Rate of effusion for gas 2}} = \frac{\sqrt{M_2}}{\sqrt{M_1}}$$

Figure 5.24: (top) When $HCl(g)$ and $NH_3(g)$ meet in the tube, a white ring of $NH_4Cl(s)$ forms. (bottom) A demonstration of the relative diffusion rates of NH_3 and HCl molecules through air.

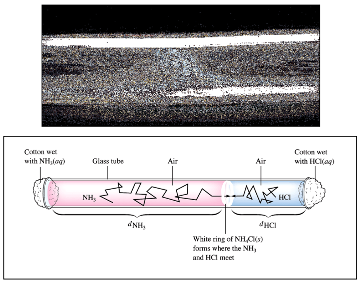


Figure 5.22: The effusion of a gas into an evacuated chamber.