Physical Chemistry

The First Course

SYLLABUS

-Gases [ideal gas laws, Kinetic theory of gases, Maxwells distribution of molecular velocities, collision properties, mean free path, deviation from ideal behavior-vander waals equation

-Thermochemistry

- Second law of thermodynamics and its applications

- Chemical equilibrium of homogenous and heterogeneous reaction

- Colligative properties [Osmotic pressuer,boiling point elevation, freezing point depression]

Reference

-Physical Chemistry   by AtKins

- Physical Chemistry   by Barrow

-كيمياء فيزيائية مسائل وحلول   د.انيس النجار

Common unit

Volume

$1m = 1000mm$

$= 100cm$

$= 10dm$

$1ml = 1cm^3$

$1L = 1000ml = 1000cm^3$

$1L = dm^3$

$1m^3 = (100)^3 cm^3$

$1m^3 = 1000L$

Pressure

$$P = F/A = N/m^2 = Pa$$

$$\frac{J/m}{m^2} = \frac{J}{m^3}$$

$$J = Kg\,\frac{m^2}{S^2}$$

$$N = \frac{J}{m}$$

$$P = \frac{KgS^{-2}m^2}{m} = Kg.m.\,S^{-2}$$

$$= kg/mS^2$$

1atm=760mmHg

=760 torr

1torr=1mmHg

1atm=101325Pa

=1.01325x10^{5}Pa

=101.325 kPa

1bar=10^{5}Pa

1bar=0.989 atm

1 torr=101325/760

=133.3 Pa

1mmHg =133Pa.

The gas Law

Four variables are needed to define the physical condition, or *state*, of a gas: temperature, pressure, volume, and amount of gas, usually expressed as number of moles. The equations that express the relationships among these four variables are known as the *gas laws*

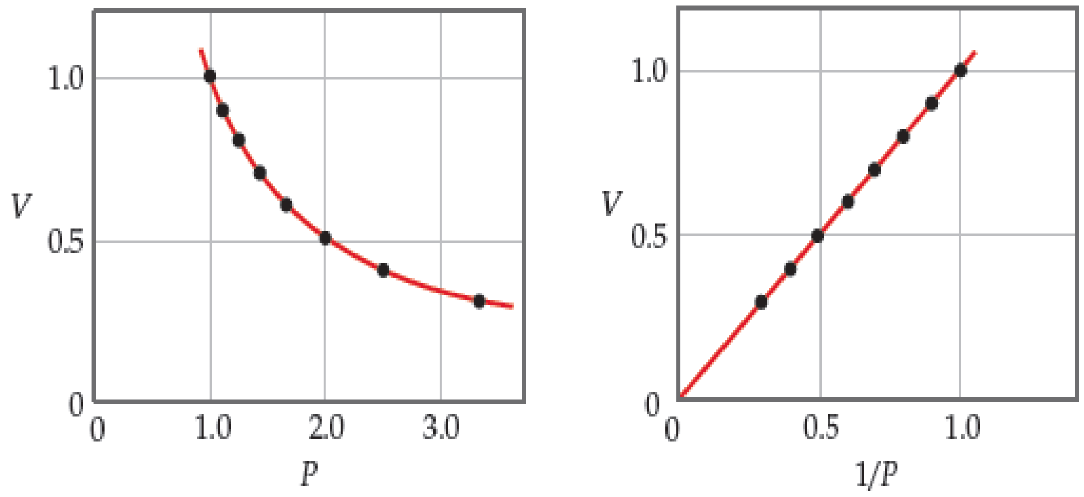
The Pressure-Volume Relationship: Boyel s Law

Boyle’s law can be expressed mathematically as

$V = \text{constant} \times \frac{1}{p}$ or $PV = \text{constant}$

$$\mathbf{P_1V_1=P_2V_2 \text{ at constant(T, n)}}$$

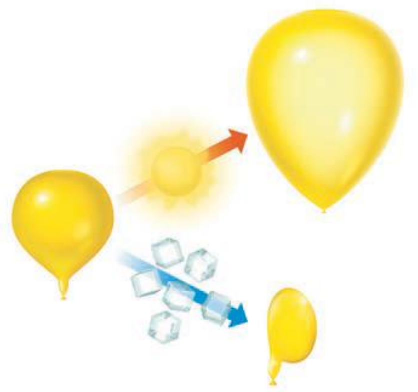
What would a plot of *P* versus 1/V look like for a fixed quantity of gas at a fixed temperature?

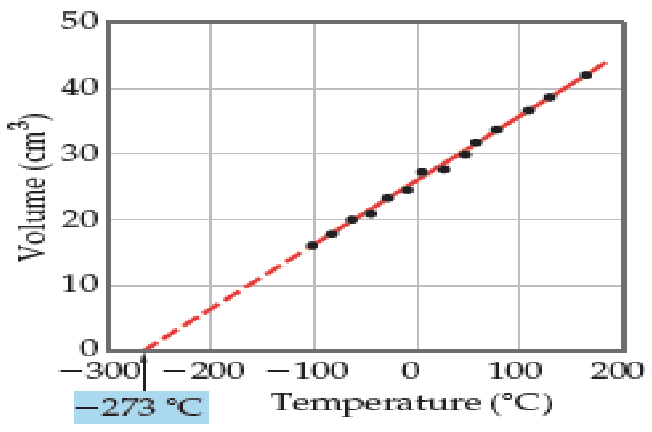


▲ Figure 10.6 **Boyle’s Law.** For a fixed quantity of gas at constant temperature, the volume of the gas is inversely proportional to its pressure.

The Temperatuer-Volume Relationship: Charles s Law

The relationship between gas volume and temperature—volume increases as temperature increases and decreases as temperature decreases—was discovered in 1787 by French scientist Jacques Charles





Mathematically, Charles’s law takes the form

$V = \text{constant} \times T$ or $\frac{V}{T} = \text{constant}$

$$\frac{V1}{T1} = \frac{V2}{T2}$$ at constant (p . n)

The Quantity–Volume Relationship :Avogadro's Law

Avogadro's law follows from Avogadro's hypothesis: The volume of a gas maintained at constant temperature and pressure is directly proportional to the number of moles of the gas. That is,

V = constant x n   or V. n = constant $\frac{V1}{n1} = \frac{V2}{n2}$ **at constant (T,P)**

Practice Exercise

A helium balloon is filled to a volume of 5.60 liters at 25 °C. What will the volume of the balloon become if it is put into liquid nitrogen to lower the temperature of the helium to 77 K?

(a) 17 L   (b) 22 L   (c) 1.4 L   (d) 0.046 L   (e) 3.7 L

The Ideal-Gas Equation

PV = nRT

Gas Densities and Molar Mass

n=m/M, d=m/V   P=m/V. RT/M

$$d = \frac{n\mathcal{M}}{V} = \frac{P\mathcal{M}}{RT}$$

Exercise:- What is the density of carbon tetrachloride vapor at 714 torr and 125 °C?

$$\mathcal{M} = \frac{dRT}{P}$$

$$d = \frac{(0.939\ \text{atm})(153.8\ \text{g/mol})}{(0.08206\ \text{L-atm/mol-K})(398\ \text{K})} = 4.42\ \text{g/L}$$

<u>Gas Mixtures and Partial pressuer</u>

The total pressure of a mixture of gases equals the sum of the pressures that each would exert if it were present alone

$P_t = P_1 + P_2 + P_3 + c$

$P_1 = n_1 (RT/V)$

$P_2 = n_2 (RT/V)$

$P_3 = n_3 (RT/V)$

$P_T = (n_1 + n_2 + n_3 + \ldots) (RT/V) = n_T (RT/V)$

Exercise :- mixture of 6.00 g of O_2 and 9.00 g of CH_4 is placed in a 15.0-L vessel at 0 °C. What is the partial pressure of each gas, and what is the total pressure in the vessel?

Solve

We first convert the mass of each gas to moles:

$nO_2 = (6.00 \text{ g } O_2)(1 \text{ mol } O_2/32.0 \text{ g } O_2) = 0.188 \text{ mol } O_2$

$nCH_4 = (9.00 \text{ g } CH_4)(1 \text{ mol } CH_4/16.0 \text{ g } CH_4) = 0.563 \text{ mol } CH_4$

We use the ideal-gas equation to calculate the partial pressure of each gas:

$PO_2 = n_{O2}RT/V = (0.188 \text{ mol})(0.08206 \text{ L-atm/mol-K})(273 \text{ K})/15.0 \text{ L} = 0.281 \text{ atm}$

$PCH_4 = n_{CH4}RT/V = (0.563 \text{ mol})(0.08206 \text{ L-atm/mol-K})(273 \text{ K})/15.0 \text{ L} = 0.841 \text{ atm}$

According to Dalton's law of partial pressures (Equation 10.12), the total pressure in the vessel is the sum of the partial pressures:

$Pt = P_{O2} + P_{CH4} = 0.281 \text{ atm} + 0.841 \text{ atm} = 1.122 \text{ atm}$