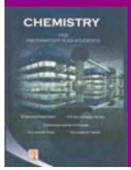
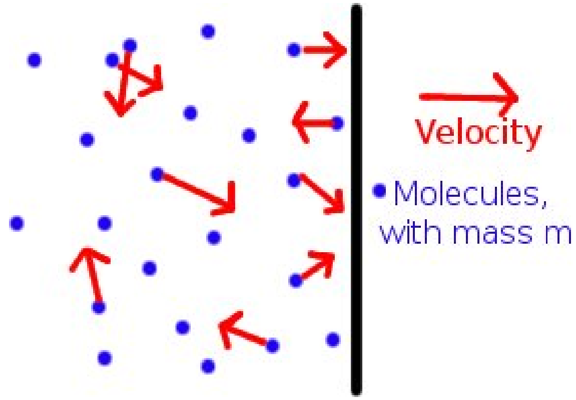
Chapter 3
Gases



Properties of Gases

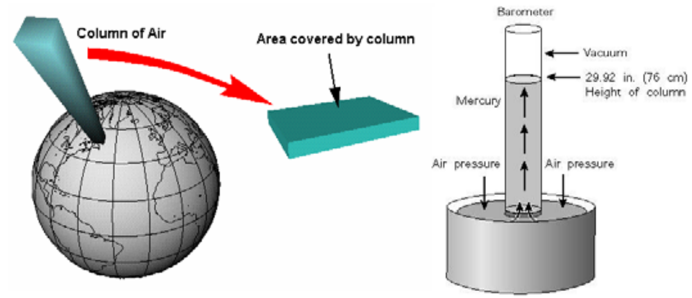
- Gases have four main characteristics compared with solids and liquids:
- Gases take the volume and shape of their containers.
- Mix completely (homogeneously) with any other gas.
- Compressible: their volume decrease when pressure applied.
- Exert pressure on its surroundings.



Pressure of Gases and its Units

Pressure is defined as the force applied per unit area.

Standard atmospheric pressure: the pressure that supports a column of mercury exactly 760 mm high at 0oC at sea level. A **barometer** is used to measure the pressure as shown.



$$\text{Pressure (P)} = \frac{\text{force}}{\text{area}} \quad \frac{\text{N}}{\text{m}^2}$$

$1\ \text{Pa} = 1\ \text{N m}^{-2}$

Common Units of Pressure

SI units of pressure is kg/m. s^2

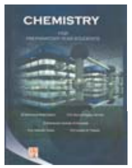
Pascal, Pa = Newton/meter^2 = N/m^2 (N = kg m/s^2)

Standard Atmospheric Pressure = 101.3 kPa

=1 atmosphere (atm)

= 760 torr

= 760 mm Hg



Gas Laws

Boyle's Law:

(the relationship between volume and pressure):

"*The volume of a sample of gas is inversely proportional to its pressure, if temperature remains constant*"

Pressure x Volume = Constant (T = constant)

P_1 and V_1 are initial conditions

P_2 and V_2 are final conditions

$P_1V_1 = P_2V_2$ (T = constant)

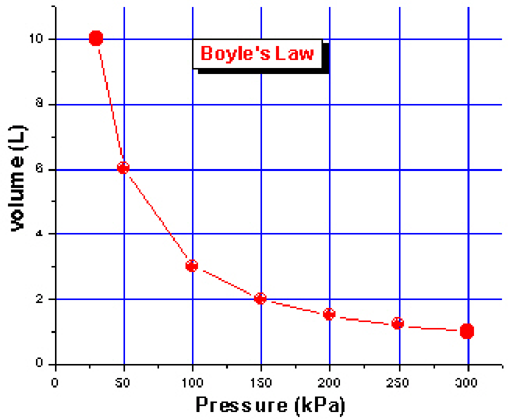


Fig 6.2 Boyle's Law and the relationship between volume and pressure.

Example

A cylinder of oxygen has V = 50.0 L and P = 10 atm at 20°C. What will be the volume of the gas at atmospheric pressure (1 atm) and 20°C?

According to Boyle's law $P_1V_1 = P_2V_2$

P_1= 10 atm, V_1 = 50.0 L, P_2= 1 atm , V_2 = ??

(10 atm x 50 L) = (1 atm x V_2), and V2 = 500 L

Gas Laws

Charles's Law:

(the relationship between volume and temperature):

"The volume of a fixed amount of gas is directly proportional to the absolute temperature of the gas at constant pressure."

Volume = Temperature x Constant   (P = constant)

$$\frac{V_1}{T_1} = \frac{V_2}{T_2}$$

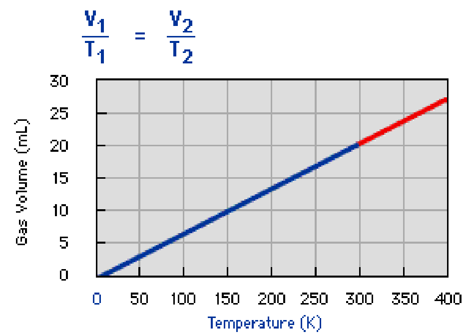


Fig. 6.3 Charles's Law and the relationship between volume and temperature.

Example

A sample of gas at 1 atm pressure and 25°C has V = 1 L, The gas is cooled to -196 °C. What is the new volume?

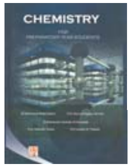
According to Charles' law $\mathbf{V_1/T_1 = V_2/T_2}$

V_1= 1.0 L, T_1 = 25°C = 25+ 273.15 = 298.15 K,

V_2= ?? L, T_2 = -196°C = -196 + 273.15 = 76.15 K

$\mathbf{V_1/T_1 = V_2/T_2}$ (**1.0** L / 298.15 K) = (V_2 L / 76.15 K)

V_2= 0.26 L   (temperature decreased, volume decreased)



Gas Laws

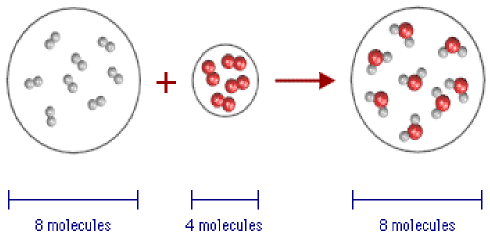
Avogadro's Law:

(the relationship between volume and amount):

"The volume of a gas is directly proportional to the number of moles (n) of gas at constant temperature and pressure."

Volume = # moles x Constant

$$\frac{V_1}{n_1} = \frac{V_2}{n_2}$$



Example

In the following reaction: $3\ H_2\ (g) + N_2\ (g) \rightarrow 2\ NH_3\ (g)$

Before the reaction begins, the reactants occupy a volume of 1.32 L. If a constant pressure is maintained in the reaction, what is the final volume occupied by the product after the completion of the reaction (assume all reactants are consumed)?

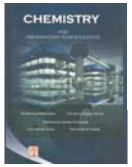
According to Avogadro's law **$V_1/n_1 = V_2/n_2$**

V_1= 1.320 L, n_1 = 3 mole (**H_2**) + 1 mole (**N_2**) = 4 moles,

V_2= ?? L, n_2 = 2 moles (**NH_3**)

$V_1/n_1 = V_2/n_2$ (**1.32** L / 4 moles K) = (V2 L / 2 moles)

V_2= 0.660 L (number of moles decreased, volume decreased)



Ideal Gas Law

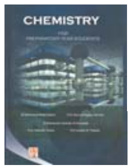
Boyle's Law: $V \propto 1/P$ (T = constant)
Charles's Law: $V \propto T$ (P = constant)
Avogadro's Law: $V \propto n$ T and P are constant

Therefore: $V \propto nT/P$ or $PV \propto nT$

So, the Ideal Gas Law: PV = nRT

Where: R = proportionality constant called the universal gas constant and it is equal = 0.08206 L atm K^{-1} mol^{-1}, or 8.3145 J mol^{-1}. K^{-1}.

P = pressure in atm **V =** volume in Litres
n = moles **T =** temperature in Kelvin



Example

What is the lung capacity (volume) of an average adult lung, if the number of moles of the air in the lungs is 0.15 moles? Assume that the person is at 1.00 atm pressure and has a normal body temperature of 37^{o}C.

Answer

According to the ideal gas law **PV = nRT**

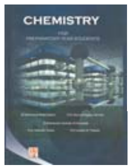
P = 1.00 atm, V= ??, n = 0.15 moles, R = 0.08206 L atm K^{-1} mol^{-1},

T = 37 ^{o}C = 37 + 273.15 = 310.15 K

V = nRT/P

= (0.15 mole x 0.08206 L atm K^{-1} mol^{-1} x 310.15 K)/1.00 atm

The lung capacity V = 3.8 L



Standard Temperature and Pressue "STP"

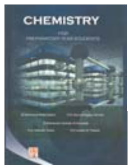
The STP are considered when the pressure is equal to 1 atm and the temperature = 0 ^{o}C (273 K).

Example

What is the molar volume (volume of 1 mole) of an ideal gas at STP?

According to the ideal gas law: PV = nRT

- V = (nRT)/P = (1 mole x 0.08206 L atm K^{-1} mol^{-1} x 273 K)/ (1.00 atm)
- Molar volume Vm = 22.4 L
- ***The molar volume of ANY ideal gas is 22.4 liters at STP.***



Gas Density

$d = m/V$

According to the ideal gas law: $PV = nRT$

Number of moles n = (mass / molar mass)
$= m/M$

$$PV = nRT = (m/M)RT = (mRT/M)$$

$$P \times M = (m/V)\ (RT)$$

$$P \times M = d\ (RT)$$

$$\boxed{d = \frac{PxM}{RT}}$$

Calculate the density of Cl_2 gas at STP

According to the formula $d = P(M)/ (RT)$

Molar mass (M) of Cl_2 gas $= 2x\ 35.5$

$= 71.0$ g/mole

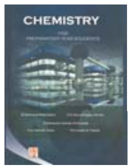
At STP

$P = 1$ atm $\qquad\qquad T= 273$ K

$R = 0.08216$ L atm/mol

$d = (1 \text{ atm} \times 71.0 \text{ g/mole})/ (0.08216 \text{ L atm/mol K} \times 273 \text{ K})$

$= 2.91$ g/L



Dalton's Law of Partial Pressure:

For a mixture of gases in a container, the overall pressure is the sum of all the partial pressures of the individual components.

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

Example

A mixture of oxygen, hydrogen and nitrogen gases exerts a total pressure of 278 kPa. If the partial pressures of the oxygen and the hydrogen are 112 kPa and 101 kPa respectively, what would be the partial pressure exerted by the nitrogen.

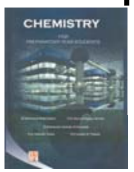
According to Dalton's law of partial pressure

$P_{total} = P_1 + P_2 + \ldots P_n$

$278\ kPa = 112\ kPa + 101\ kPa + P_{nitrogen}$

$P_{nitrogen} = 278\ kPa - (112\ kPa + 101\ kPa)$

$P_{nitrogen} = 65\ kPa$



Mole Fraction (X)

- If a mixture contains two components, A and B, then:

$$X_A = \frac{\text{moles A}}{\text{moles A} + \text{moles B}} \qquad X_B = \frac{\text{moles B}}{\text{moles A} + \text{moles B}}$$

$$X_A + X_B = 1$$

Example

A gaseous mixture contains 0.38 moles of nitrogen gas and 0.45 moles of oxygen gas. Determine the mole fractions of oxygen and nitrogen.

Total number of moles = 0.38 mol nitrogen + 0.45 mole oxygen = 0.83 moles

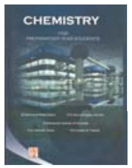
X_{Oxygen} = (moles of oxygen / total number of moles)

= (0.45 mole/ 0.83 moles) = 0.54

$X_{nitrogen}$ = (moles of nitrogen / total number of moles)

= (0.38 mole/ 0.83 moles) = 0.46

Check: $X_{Oxygen} + X_{nitrogen}$ = 0.54 + 0.46 =1.00



Determining Partial Pressure from Mole Fraction

$$X_A = \frac{\text{moles A}}{\text{moles of all species in sample}}$$

$$P_A = \frac{n_A RT}{V} \qquad P_{total} = \frac{n_{total} RT}{V}$$

$$\frac{P_A}{P_{total}} = \left(\frac{\frac{n_A RT}{V}}{\frac{n_{total} RT}{V}} \right) = \frac{n_A}{n_{total}} = X_A$$

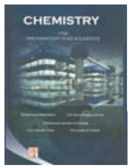
$$P_A = P_{total} X_A$$

Example

The mole fraction of nitrogen in air is 0.7808. Calculate the partial pressure of N2 in air when the atmospheric pressure is 760 torr.

Use the partial pressure formula:

$$P_{N_2} = P_{total} X_{N_2} = (760 \text{ torr})(0.7808) = 593 \text{ torr}$$



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