

Gas Laws Synopsis

Chapter Terms and Definitions:

Pressure: force exerted per unit area of surface

Acceleration: change of speed per unit time

Pascal (Pa): SI unit of pressure; $1 \text{ Pa} = 1 \text{ kg/m}\cdot\text{s}^2$

Barometer: device for measure the pressure of the atmosphere

Manometer: device that measures the pressure of a gas or liquid in a sealed vessel

Millimeters of mercury (mmHg or torr): traditional unit of pressure equal to that exerted by a 1 mm column of mercury at 0.00°C in a barometer or manometer

Atmosphere(atm): traditional unit of pressure equal to exactly 760 mm Hg; $1 \text{ atm} = 101.325 \text{ kPa}$ exact.

Compressibility: ability to be squeezed into a smaller volume by the application of pressure

Boyle's Law: the volume occupied by any given amount of gas at a fixed temperature varies inversely with its pressure.

Kelvin scale: absolute temperature scale on which the units (Kelvins, K) are given by $K = ^\circ\text{C} + 273.15$.

Charles's Law: The volume occupied by any given sample of gas at constant pressure is directly proportional to the absolute temperature

Law of combining volumes: Volumes of reactant gases at a given pressure and temperature are in ratios of small whole number.

Avogadro's Law: Equal volumes of any two gases at the same temperature and pressure contain the same number of molecules.

Molar gas volume (V_M): The volume occupied by one mole of any gas at a given temperature or pressure

Standard Temperature and Pressure (STP): reference conditions chosen by convention to be 0°C and 1 atm pressure

Molar Gas Volume: Invariant (unchanging) factor relating the molar volume of a gas to T/P .

Ideal Gas Law: Mathematical expression relating the volume (V), pressure (P), Kelvin temperature (T), and moles (n) of a gas to the molar gas constant R ; $PV = nRT$

Amonton's Law or Gay-Lussac's Law: the pressure of a given amount of gas at a fixed volume is proportional to the absolute temperature

Partial Pressure: Pressure exerted by a particular gas in a gas mixture

Dalton's Law of Partial Pressures: the sum of the partial pressures of all the different gases in a mixture is equal to the total pressure of the mixture

Mole Fraction: fraction of moles of a component gas in the total moles of a gas mixture

Vapor Pressure: Partial pressure of the molecules of a substance in the gaseous state over the liquid (or solid)

Kinetic Molecular Theory of Gases: idea that a gas consists of molecules in constant random motion

Postulates: basic statements from which all conclusion or predictions of a theory are deduced

Ideal gas: gas that follows the ideal gas law; its molecules have essentially no volume of their own, and no attraction for each other.

Intermolecular forces: forces of attraction or repulsion between molecules.

Root-Mean-Square (rms) molecular speed (u) type of average molecular speed, or the speed of a molecule that has the average molecular kinetic energy; can be shown to equal $u = (3RT/M_m)^{1/2}$ Where R is the molar gas constant, T , is the Kelvin temperature, and M_m is the molar mass for the gas.

Diffusion: process whereby a gas spreads out through another gas to occupy a space wherein the partial pressure is uniform.

Effusion: escape of a gas through a small hole into a vacuum at the same velocity it had in the container

Graham's Law of Effusion: the rate of effusion of gas molecules from a particular hole is inversely proportional to the square root of the molecular weight of the gas at constant T and P .

Van der Waals Equation: mathematical expression relating P , T , V and n for nonideal gases at moderately high pressures and low temperatures.

I. The Ideal Gas Law

a. Four Variables: P, V, n, T

Relationship between variables:

$$PV = nRT$$

where the constant $R = 0.821 \text{ L}\cdot\text{mol}\cdot\text{K}$

P = pressure (atmospheres, mmHg)

V = volume (liters, cm^3 , m^3)

n = amount in moles

T = temperature (Kelvin)

b. R can be calculated from experimental values in P, V, n and T . One mole of gas occupies a volume of 22.4 L at 0°C and 1.00 atm. One atm and $0^\circ\text{C} = 273 \text{ K}$ are **Standard Temperature and Pressure** units.

Note that other R 's can be calculated by substituting into the Ideal Gas Law Equation utilizing other pressure units such as mmHg and solving for R . $R = PV/nT$

$$R = 760\text{mmHg}\cdot 22.4 \text{ L}/1 \text{ mol}\cdot 273 \text{ K} = 62.4 \text{ mmHg}\cdot\text{L}/\text{Mol}\cdot\text{K}$$

c. Alterations/Manipulations

How do we calculate a mole? Mole = mass (in grams)/molecular weight

The symbol for mole = n

By Definition: $n = \text{mass}/\text{Mwt}$ mass is grams (g)

If you substitute this into the Ideal Gas Law for n then: $PV = \frac{g \cdot R \cdot T}{\text{Mwt}}$

Rearrangement of this equation enables one to solve for Density and Molecular Weight

- What is Density? $D = \text{mass}/\text{Volume}$ or g/V
Rearranging the Ideal Gas Law and solving for Density
 $g/V = P \cdot \text{Mwt}/RT = \text{Density}$ $D = P \cdot \text{Mwt}/RT$
**note that Density will calculate to the units of grams/Liter.*
- To solve for Molecular Weight, rearrange the equation:
 $\text{Mwt} = g \cdot R \cdot T / P \cdot V$

d. Real Gases vs. Ideal Gases

Real Gases deviate slightly from ideal because of two important factors:

- Gas molecules do attract each other
- Gas molecules do occupy a finite volume
- Both of these factors are neglected in the Ideal Gas Law.
Both increase in importance when the molecules are close together (**high pressure, low temperature**)
- Van der Waal's equation for real gases correct for the force of attraction and for particle volume
 $(P + n^2\mathbf{a}/V^2)(V - n\mathbf{b}) = n \cdot R \cdot T$ [\mathbf{a} is the attraction correction, \mathbf{b} is the volume correction]