

Boyle's Law	Charles' Law	Gay-Lussac Law	Combined Law
For a given mass of gas at constant temperature, the volume of a gas varies inversely with pressure.	The volume of a fixed mass of gas is directly proportional to its Kelvin temperature if the pressure is kept constant.	The pressure of a gas is directly proportional to its Kelvin temperature if the volume is kept constant.	Boyle's, Charles' and Gay-Lussac Laws combine the relationship between pressure, volume and temperature.
$PV = k$ $P_1V_1 = P_2V_2$	$\frac{V}{T} = k$ $V_1T_1 = V_2T_2$ $\frac{V_1}{T_1} = \frac{V_2}{T_2}$	$\frac{P}{T} = k$ $P_1T_1 = P_2T_2$ $\frac{P_1}{T_1} = \frac{P_2}{T_2}$	$\frac{PV}{T} = k$ $P_1V_1T_1 = P_2V_2T_2$ $\frac{P_1V_1}{T_1} = \frac{P_2V_2}{T_2}$

Dalton's Law	Ideal Gas Law	Graham's Law
At constant volume and temperature, the total pressure exerted by a mixture of gases is equal to the sum of the pressures of each gas.	The Ideal Gas Law relates the pressure, temperature, volume and moles of a gas through the gas constant "R". The Ideal Gas Law can also relate density and molar mass with the gas constant "R".	The rate of effusion of two gases is inversely proportional to the square root of the ratio of the molar masses of each gas.
$P_{\text{total}} = nP_1 + nP_2 + nP_3 \dots$ $P_{\text{total}} = P_{\text{gas}} + P_{\text{vap}}$	$PV = nRT$ $PM = dRT$	$\frac{\text{Rate}_1}{\text{Rate}_2} = \sqrt{\frac{\text{Molar Mass}_2}{\text{Molar Mass}_1}}$

Abbreviations	Standard Conditions
atm = atmosphere mmHg = millimeters of Hg bar = barometer kPa = kilopascals dm ³ = decimeters cubed cm ³ = centimeters cubed	°C = Celsius K = Kelvin mol = moles L = milliliters mL = milliliters
Conversions	Gas Law Conversion Symbols
K = °C + 273 °F = 1.8°C + 32 °C = $\frac{F - 32}{1.8}$ 1 atm = 101.3 kPa 1 L = 1000 mL = 1000cm ³ = 1 dm ³	Subscript 1 = original or initial condition. Subscript 2 = new or final condition. Temperatures must be in Kelvin. Always!! n = number of moles. R = 0.0821 $\frac{\text{atm} \cdot \text{L}}{\text{mol} \cdot \text{K}}$

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