11-2 Ideal Gas Law

Ideal Gas Constant

Moles can be included in the Combined Gas Law

$$\frac{P_1V_1}{n_1T_1} = \frac{P_2V_2}{n_2T_2} = \text{Constant}$$

1 mole of a gas at STP conditions can be substituted in:

$$\frac{PV}{nT} = \frac{(1\ \text{atm})(22.4\ \text{L})}{(1\ \text{mol})(273\ \text{K})} = 0.0821\ \frac{\text{L} \cdot \text{atm}}{\text{mol} \cdot \text{K}}$$

This gives us the ideal gas constant "R"

$$R = 0.0821\ \frac{\text{L} \cdot \text{atm}}{\text{mol} \cdot \text{K}}$$

11-2 Ideal Gas Law

Ideal gas Law

$\frac{PV}{nT}$ **= R can be have both sides multiplied by nT to give:**

Pressure **Volume** **Temperature (K)**

$$PV = nRT$$

Moles **Ideal Gas Constant**

$$R = 0.0821\ \frac{\text{L} \cdot \text{atm}}{\text{Mol} \cdot \text{K}}$$

This can find the pressure, volume, temperature or moles of a gas

11-2 Ideal Gas Law

What is the pressure in atmospheres exerted by a 0.500 mol sample of nitrogen gas in a 10.0 L container at 298 K?

$$PV = nRT \quad \rightarrow \quad P = \frac{nRT}{V}$$

$$P = \frac{(0.500\ \cancel{\text{mol}})(0.0821\ \frac{\cancel{\text{L}} \cdot \text{atm}}{\cancel{\text{mol}} \cdot \cancel{\text{K}}})(298\ \cancel{\text{K}})}{(10.0\ \cancel{\text{L}})}$$

$$P = 1.22\ \text{atm}$$

11-2 Ideal Gas Law

What mass of chlorine gas, Cl_2, in grams, is contained in a 10.0 L tank at 27°C and 3.50 atm of pressure?

$$PV = nRT \quad \rightarrow \quad n = \frac{PV}{RT}$$

$$27°\text{C} + 273 = 300\text{K}$$

$$n = \frac{(3.5\ \cancel{\text{atm}})(10.0\ \cancel{\text{L}})}{(0.0821\ \frac{\cancel{\text{L}} \cdot \cancel{\text{atm}}}{\text{mol} \cdot \cancel{\text{K}}})(300\ \cancel{\text{K}})} = 1.42\ \text{mol}\ Cl_2$$

$$1.42\ \text{mol}\ Cl_2 \left(\frac{71.0\ \text{g}}{\cancel{\text{mol}}}\right) = 101\ \text{g}\ Cl_2$$

11-2 Ideal Gas Law

Molar Mass of a Gas

Molar Mass can be found using substitution

$$\frac{\textbf{Mass}}{\textbf{Molar Mass}} = \textbf{Moles (n)}$$

With the substitution of moles, we can relate Molar Mass and Mass

$$PV = \frac{\text{Mass} \cdot RT}{\text{Molar Mass}}$$

Solving for Molar Mass we get:

$$\boxed{\text{Molar Mass} = \frac{\text{Mass} \cdot RT}{PV}}$$

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At 28°C and 0.974 atm, 1.00 L of gas has a mass of 5.16 g. What is the molar mass of this gas?

$$\text{Molar Mass} = \frac{\text{Mass} \cdot RT}{PV}$$

$$28°C + 273 = 301K$$

$$\frac{(5.16\ g)(0.0821\ \frac{\cancel{L} \cdot \cancel{atm}}{mol \cdot \cancel{K}})(301\ \cancel{K})}{(0.974\ \cancel{atm})(1.00\ \cancel{L})} = 131\ \frac{g}{mol}$$

11-2 Ideal Gas Law

Gas Density

Gas density can be found using substitution

$$\frac{\text{Mass}}{\text{Molar Mass}} = \text{Moles (n)} \qquad \frac{\text{Mass}}{V} = \text{Density}$$

With the substitution of moles, then mass per volume, we can relate density

$$PV = \frac{\text{Mass} \cdot RT}{\text{Molar Mass}} \qquad P = \frac{\text{Mass} \cdot RT}{V \cdot \text{Molar Mass}} \qquad P = \frac{\text{Density} \cdot RT}{\text{Molar Mass}}$$

Solving for density we get:

$$\boxed{D = \frac{\text{Molar Mass} \cdot P}{RT}}$$

11-2 Ideal Gas Law

What is the density of a sample of ammonia gas, NH3, if the pressure is 0.928 atm and the temperature is 63.0°C?

$$D = \frac{\text{Molar Mass} \cdot P}{RT}$$

$$63°C + 273 = 336\ K$$

$$\text{Molar Mass} = 14.0 + 3 \times 1.0 = 17.0\ \frac{g}{mol}$$

$$\frac{(17.0\ \frac{g}{\cancel{mol}})(0.928\ \cancel{atm})}{(0.0821\ \frac{L \cdot \cancel{atm}}{\cancel{mol} \cdot \cancel{K}})(336\ \cancel{K})} = 0.572\ \frac{g}{L}$$