Key!

Chapter 10: Gases

What Gas law has these conditions?

Boyles' Law: For a given mass of gas at a constant temperature, the volume of a gas varies inversely with pressure.

$$P_1 V_1 = P_2 V_2$$

Constants: moles (n) + temperature (T)

Gay-Lussacs Law: The volume of a fixed mass of gas is directly proportional to its Kelvin temperature if the pressure is kept constant.

$$\frac{P_1}{T_1} = \frac{P_2}{T_2}$$

Constants: moles (n) + volume (V)

Charles' Law: The volume of a fixed mass of a gas is directly proportional to its absolute temperature.

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

Constants: moles (n) + Pressure (P)

Avogadro's law: The volume of a gas at constant temperature and pressure is directly proportional to the number of moles of the gas.

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

Constants: temperature (T) + Pressure (P)

What is the Combined Gas Law formula?

Boyles + Charles + Gay-Lussacs

$$\frac{P_1 V_1}{n_1 T_1} = \frac{P_2 V_2}{n_2 T_2}$$

What's the mole fraction of F_2? Given: 0.35 mol F_2 + 0.25 mol H_2 + 0.65 mol CO_2 = 1.25 mol

$\frac{\text{Part}}{\text{Whole}}$

$$\underbrace{X}_{\text{symbol for mole fraction}} = \frac{0.35 \text{ mol } F_2}{1.25 \text{ mol}} = \boxed{0.28}$$

no units!

Always convert °C to K!

A submarine with a volume of 1.2×10^5 L has an internal pressure of 1.0 atm and an internal temperature of 15° C. If the submarine descends to a depth where the pressure is 150 atm and the temperature is 3.0° C, what will the volume of the gas inside be if the hull of the submarine breaks?

$V_1 = 1.2 \times 10^5\ L$
$P_1 = 1.0$ atm
$T_1 = 15°C$
$P_2 = 150$ atm
$T_2 = 3.0°C$
$V_2 = ?$

$C + 273.15 = K$

$T_1 = 15°C + 273.15 = 288.15\ K$

$T_2 = 3.0°C + 273.15 = 276.15\ K$

$$\frac{T_2}{P_2} \cdot \frac{P_1 V_1}{T_1} = \frac{\cancel{P_2} V_2}{\cancel{T_2}} \cdot \frac{\cancel{T_2}}{\cancel{P_2}}$$

$$V_2 = \frac{P_1 V_1 T_2}{T_1 P_2} = \frac{(1.0\ atm)(1.2 \times 10^5\ L)(276.15\ K)}{(288.15\ K)(150\ atm)}$$

$$= 766.6840187\ L$$

$$\boxed{V_2 = 7.7 \times 10^2\ L}$$

sig figs!

At what temperature will 5.00 g of Cl_2 exert s pressure of 900. Torr at a volume of 750. mL? 3 sig figs

$T = ?$
$V = 750$ mL
$P = 900$ torr
$n = 5.00\ g\ Cl_2$
$R = \left(0.0821 \frac{atm \cdot L}{mol \cdot K}\right)$

$750\ ml \cdot \frac{1\ L}{1000\ ml} = 0.750\ L$

$900\ torr \cdot \frac{1.0\ atm}{760\ torr} = 1.18\ atm$

$5.00\ g\ Cl_2 \cdot \frac{1\ mol}{70.9\ g\ Cl_2} = 0.0705\ mol$

Ideal Gas Law

$PV = nRT$

R ↓ constant

$$T = \frac{PV}{nR} = \frac{(1.18\ \cancel{atm})(0.750\ \cancel{L})}{(0.0705\ \cancel{mol})\left(0.0821 \frac{\cancel{atm} \cdot \cancel{L}}{\cancel{mol} \cdot K}\right)}$$

$$= 152.901$$

$$\boxed{T = 152\ K}$$

Ammonia is often formed by reacting nitrogen and hydrogen gases. How many liters of ammonia gas can be formed from 24.5 L of hydrogen gas at 93.0° C and a pressure of 33.9 kPa?

NH_3 →

$\underbrace{N_2 + 3H_2 \rightarrow 2NH_3}_{\text{Equation}}$

$PV = nRT$

* rearrange to find moles
* Plug in values with units

$V_{NH_3} = ?$

$V_{\text{Hydrogen gas}} = 24.5\ L$

$T = 93.0°C + 273.15 = 366.15\ K$

$P = 33.9\ kPa \cdot \frac{1\ atm}{101.3\ kPa} = 0.335\ atm$

$$n = \frac{PV}{RT} = \frac{(0.335\ \cancel{atm})(24.5\ \cancel{L})}{\left(0.0821 \frac{\cancel{atm} \cdot \cancel{L}}{mol \cdot \cancel{K}}\right)(366.15\ \cancel{K})}$$

$$n = 0.273\ mol\ H_2$$

(STP) * got from Mole roadmap

$$0.273\ mol\ H_2 \cdot \frac{2\ mol\ NH_3}{3\ mol\ H_2} \cdot \frac{22.4\ L}{1\ mol\ NH_3} = 4.0768 \rightarrow \boxed{4.08\ L\ NH_3}$$

Empirical formula: % changes to decimal

A container with two gases, helium and argon, is 30.0% helium. Calculate the partial pressure of helium and argon if the total pressure inside the container is 4.00 atm

$P_{total} = P_{gas} + P_{gas} + P_{gas} \ldots$

$4.00 \text{ atm} \times 0.30 = 1.2 \text{ He}$

$\boxed{\text{He} = 1.2 \text{ partial pressure}}$

$\text{Ar} = (4.00 \text{ atm} - 1.2 \text{ atm})$

$\boxed{\text{Ar} = 2.8 \text{ partial pressure}}$

solve this way OR solve this way

% = out of hundred

$n_{He} = 30.0\% \rightarrow 0.30 \text{ mol He}$

$n_{Ar} = 70.0\% \rightarrow 0.70 \text{ mol Ar}$

$P_{He} = \left(\frac{0.30 \text{ mol}}{1 \text{ mol}}\right)(4.00 \text{ atm}) = 1.20 \text{ atm He}$

$P_{Ar} = \left(\frac{0.70 \text{ mol}}{1 \text{ mol}}\right)(4.00 \text{ atm}) = 2.8 \text{ atm Ar}$

A sample of nitrogen gas (N_2) is contained in a 5.00 L container at a temperature of 300 K and a pressure of 2.00 atm. Calculate the density of the nitrogen gas.

$mm = \frac{dRT}{P}$

*rearrange formula

$d = \frac{P \cdot mm}{RT}$

$mm \cdot N_{\boxed{2}} = 14.007_g \times 2 = 28.014 \text{ g/mol}$

$d = \frac{(2.00 \text{ atm})(28.014 \text{ g/mol})}{(0.0821 \frac{\text{atm}\cdot\text{L}}{\text{mol}\cdot\text{K}})(300 \text{ K})}$

$\boxed{d = 2.27 \text{ g/L}}$

Automobile air bags are inflated by nitrogen gas generated by the rapid decomposition of sodium azide, NaN_3:

$$2 NaN_3 (s) \rightarrow 2 Na (s) + 3 N_2 (g)$$

If an air bag has a volume of 36 L and is to be filled with nitrogen gas at 1.15 atm and 26.0° C, how many grams of NaN_3 must be decomposed?

$PV = nRT$

$n_{N_2} = \frac{PV}{RT} = \frac{(1.15 \text{ atm})(36 \text{ L})}{(0.0821 \frac{\text{atm}\cdot\text{L}}{\text{mol}\cdot\text{K}})(299.15 \text{ K})}$

$= 1.69 \text{ mol } N_2$ *sig figs

sig figs

$= 1.7 \text{ mol } N_2 \cdot \frac{2 \text{ mol } NaN_3}{3 \text{ mol } N_2} \cdot \frac{65.011 \text{ g } NaN_3}{1 \text{ mol } NaN_3} = 73.67913 \rightarrow \boxed{74 \text{ g } NaN_3}$