Combined gas law:

$$\frac{PV}{T} = \text{constant}$$

Must use ABSOLUTE temperature units!

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

Must use ABSOLUTE temperature units!

↑ amount (moles) of gas must be constant!

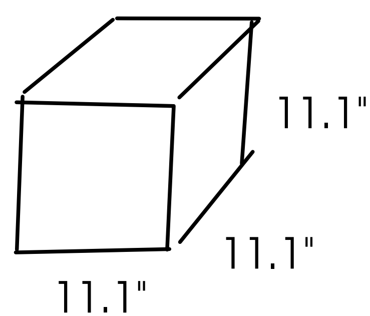
Avogadro's law:

- a mole of any gas at the same conditions has the same volume.

1 mol gas molecules @ 0°C and 1 atm

"STP"
Standard Temperature and Pressure

volume = 22.4 L



= 22.4 L

Ideal gas law:

$$\frac{PV}{T} = \text{constant}$$

... but this constant actually depends on the amount of gas!

$$= n \times \text{"}R\text{"}$$

The ideal gas constant.

$$0.08206 \frac{L \cdot atm}{mol \cdot K}$$

... combining these together ...

$$\frac{PV}{T} = nR$$

$$PV = nRT$$

P = pressure atm

V = volume L

T = ABSOLUTE temperature K

R = ideal gas constant

n = number of moles of gas molecules

A balloon is taken from a room where the temperature is 27.0 C to a freezer where the temperature is -5.0 C. If the balloon has a volume of 3.5 L in the 27.0 C room, what is the volume of the balloon in the freezer. Assume pressure is constant.

$$\frac{P_1 V_1}{T_1} = \frac{P_2 V_2}{T_2}; \text{ since } P = \text{constant}, \quad \frac{V_1}{T_1} = \frac{V_2}{T_2}$$

$V_1 = 3.5\ L$

$T_1 = 27.0°C = 300.2\ K$

$V_2 = ?$

$T_2 = -5.0°C = 268.2\ K$

$$\frac{3.5\ L}{300.2\ K} = \frac{V_2}{268.2\ K}$$

$$\boxed{3.1\ L = V_2}$$

Volume in the freezer..

2.25 L of nitrogen gas is trapped in a piston at 25.0 C and 1.00 atm pressure. If the piston is pushed in so that the gas's volume is 1.00 L while the temperature increases to 31.0 C, what is the pressure of the gas in the piston?

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

$P_1 = 1.00\ atm$

$V_1 = 2.25\ L$

$T_1 = 25.0°C = 298.2\ K$

$P_2 = ?$

$V_2 = 1.00\ L$

$T_2 = 31.0°C = 304.2\ K$

$$\frac{(1.00\ atm)(2.25\ L)}{(298.2\ K)} = \frac{P_2(1.00\ L)}{(304.2\ K)}; \quad P_2 = \boxed{2.30\ atm}$$

Calculate the mass of 22650* L of oxygen gas at 25.0 C and 1.18 atm pressure.

O_2

O_2: 32.00 g O_2 = mol O_2

*Volume of a 10'x10'x8' room

1 - Calculate moles of oxygen using ideal gas equation
2 - Convert moles oxygen gas to mass using formula weight.

$$PV = nRT$$

$$\frac{PV}{RT} = n$$

$$n = \frac{PV}{RT}$$

$P = 1.18$ atm $\quad R = 0.08206 \frac{L \cdot atm}{mol \cdot K}$

$V = 22650$ L $\quad T = 25.0°C = 298.2$ K

(1) $$n_{O_2} = \frac{(1.18\ atm)(22650\ L)}{\left(0.08206 \frac{L \cdot atm}{mol \cdot K}\right)(298.2\ K)} = 1092.222357\ mol\ O_2$$

(2) $$1092.222357\ mol\ O_2 \times \frac{32.00\ g\ O_2}{mol\ O_2} = 35\bar{0}00\ g\ O_2$$

35.0 kg
~77 lb

CHEMICAL CALCULATIONS WITH THE GAS LAWS

$FW_{NaHCO_3} = 84.007\ g/mol$

$H_2SO_4(aq) + 2NaHCO_3(s) \rightarrow 2H_2O(l) + 2CO_2(g) + Na_2SO_4(aq)$

Given 25.0 g of sodium bicarbonate and sufficient sulfuric acid, what volume of carbon dioxide gas would be produced at 25.0 C and 0.950 atm pressure?

1 - Convert 25.0 g sodium bicarbonate to moles. Use FORMULA WEIGHT.
2 - Convert moles sodium bicarbonate to moles carbon dioxide gas. Use CHEMICAL EQUATION.
3 - Convert moles carbon dioxide gas to volume. Use IDEAL GAS EQUATION.

(1) $84.007\ g\ NaHCO_3 = mol\ NaHCO_3$ (2) $2\ mol\ NaHCO_3 = 2\ mol\ CO_2$

$$25.0\ g\ NaHCO_3 \times \frac{mol\ NaHCO_3}{84.007\ g\ NaHCO_3} \times \frac{2\ mol\ CO_2}{2\ mol\ NaHCO_3} = 0.297594248\ mol\ CO_2$$

(1) (2)

$PV = nRT$

$V = \frac{nRT}{P}$

$n = 0.297594248\ mol\ CO_2$ $T = 25.0°C = 298.2\ K$

$R = 0.08206\ \frac{L \cdot atm}{mol \cdot K}$ $P = 0.950\ atm$

$$V = \frac{(0.297594248\ mol\ CO_2)\left(0.08206\ \frac{L \cdot atm}{mol \cdot K}\right)(298.2\ K)}{(0.950\ atm)} =$$

7.67 L of CO_2 at 25.0°C, 0.950 atm

What volume would the gas in the last example problem have at STP?

STP: "Standard Temperature and Pressure" (0 C and 1 atm)

$$\frac{P_1V_1}{T_1} = \frac{P_2V_2}{T_2}$$

$P_1 = 0.950$ atm $P_2 = 1$ atm

$V_1 = 7.67$ L $V_2 = ?$

$T_1 = 298.2$ K $T_2 = 273.2$ K

$$\frac{(0.950 \text{ atm})(7.67 \text{ L})}{(298.2 \text{ K})} = \frac{(1 \text{ atm})V_2}{(273.2 \text{ K})}; \quad V_2 = \boxed{6.67 \text{ L at STP}}$$

Alternate solution: Use PV=nRT to find the volume of the gas at STP. You can do this because we already calculated the moles of gas in the previous example. If you do it correctly, you should get the same answer (6.67 L at STP)!

$$2\,NH_4NO_3(s) \longrightarrow 2\,N_2(g) + O_2(g) + 4\,H_2O(g)$$

At 300°C, ammonium nitrate violently decomposes to produce nitrogen gas, oxygen gas, and water vapor. What is the total volume of gas that would be produced at 1.00 atm by the decomposition of 15.0 grams of ammonium nitrate?

To simplify the problem, let's calculate the TOTAL MOLES OF GAS instead of doing each gas individually!

$FW_{NH_4NO_3}$: 80.0434 g/mol

1 - Convert 15.0 g ammonium nitrate to moles. Use FORMULA WEIGHT.
2 - Convert moles ammonium nitrate to TOTAL MOLES GAS. Use CHEMICAL EQUATION.
3 - Convert moles gas to volume. Use IDEAL GAS EQUATION

(1) 80.0434 g NH_4NO_3 = mol NH_4NO_3 (2) 2 mol NH_4NO_3 = 7 mol gas (2+1+4=7)

$$15.0\text{ g } NH_4NO_3 \times \frac{\text{mol } NH_4NO_3}{80.0434\text{ g } NH_4NO_3} \times \frac{7\text{ mol gas}}{2\text{ mol } NH_4NO_3} = 0.655894 1774\text{ mol gas}$$

(3) $V = \frac{nRT}{P}$ | n = 0.6558941774 mol gas, R = 0.08206 $\frac{L\cdot atm}{mol\cdot K}$, T = 300.°C = 573 K, P = 1.00 atm

$$V = \frac{(0.6558941774\text{ mol gas})(0.08206\ \frac{L\cdot atm}{mol\cdot K})(573\text{ K})}{(1.00\text{ atm})} = 30.8\text{ L at } 300°C,\ 1\text{ atm}$$