GAS LAWS

- were derived by experiment long before kinetic theory, but agree with the kinetic picture!

Boyle's Law: $PV = \text{constant}$ True at constant temperature

$P_1V_1 = \text{constant}$ $P_2V_2 = \text{constant}$

$$\boxed{P_1V_1 = P_2V_2}$$

True at constant temperature

Charles's Law: $\dfrac{V}{T} = \text{constant}$ True at constant pressure, and using ABSOLUTE temperature

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

True at constant pressure, and using ABSOLUTE temperature

Combined gas law:

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

Must use ABSOLUTE temperature units!

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

Must use ABSOLUTE temperature units!

amount (moles) of gas must be constant!

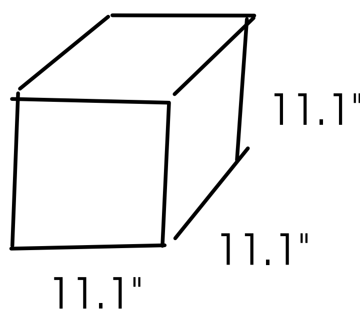
Avogadro's law:

- 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:

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

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

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

The ideal gas constant.

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

... combining these together ...

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

$$\downarrow$$

$$\boxed{PV = nRT}$$

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.

$$\frac{P_1V_1}{T_1} = \frac{P_2V_2}{T_2} \text{ ; constant } P \text{ , so } \frac{V_1}{T_1} = \frac{V_2}{T_2}$$

$V_1 = 3.5\,L$ $\quad$ $V_2 = ?$

$T_1 = 27.0°C = 300.2\,K$ $\quad$ $T_2 = -5.0°C = 268.2\,K$

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

$V_2 = \boxed{3.1\,L \text{ in freezer}}$

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?

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

$P_1 = 1.00\,atm$ $\quad$ $P_2 = ?$

$V_1 = 2.25\,L$ $\quad$ $V_2 = 1.00\,L$

$T_1 = 25.0°C = 298.2\,K$ $\quad$ $T_2 = 31.0°C = 304.2\,K$

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

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

Use the ideal gas equation to find MOLES GAS, then use oxygen's FORMULA WEIGHT to convert from moles to mass.

$$PV = nRT$$

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

P = 1.18 atm

V = 22650 L

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

T = 25.0 °C = 298.2 K

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

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

(35.0 kg)

(~77 lb)

CHEMICAL CALCULATIONS WITH THE GAS LAWS

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

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

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 using FORMULA WEIGHT.
2 - Convert moles sodium bicarbonate to moles carbon dioxide using CHEMICAL EQUATION
3 - Convert moles carbon dioxide to VOLUME using IDEAL GAS EQUATION

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

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

$PV = nRT$ (3)

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

$n = 0.297594248\ mol\ CO_2$ $R = 0.08206 \frac{L \cdot atm}{mol \cdot K}$

$T = 25.0°C = 298.2\ K$ $P = 0.950\ atm$

$$V = \frac{(0.297594248\ mol\ CO_2)\left(0.08206 \frac{L \cdot atm}{mol \cdot K}\right)(298.2\ K)}{(0.950\ atm)} = \boxed{7.67\ L \text{ at } 25.0°C,\ 0.950\ atm}$$

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 $\quad P_2 = 1$ atm

$V_1 = 7.67$ L $\quad V_2 = ?$

$T_1 = 298.2$ K $\quad T_2 = 273.15$ K

We can find the volume using the combined gas law ...

$$\frac{(0.950\text{ atm})(7.67\text{ L})}{(298.2\text{ K})} = \frac{(1\text{ atm})V_2}{(273.15\text{ K})};\ V_2 = \boxed{6.67\text{ L at STP}}$$

... or we can use the ideal gas equation!

$PV = nRT$

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

$P = 1$ atm $\quad R = 0.08206\ \frac{\text{L}\cdot\text{atm}}{\text{mol}\cdot\text{K}}$

$T = 273.15$ K $\quad n = 0.2975942481$ mol CO_2

$$V = \frac{(0.2975942481\text{ mol } CO_2)\left(0.08206\ \frac{\text{L}\cdot\text{atm}}{\text{mol}\cdot\text{K}}\right)(273.15\text{ K})}{(1\text{ atm})}$$

$$= \boxed{6.67\text{ L at STP}}$$

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

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, we will calculate TOTAL MOLES OF GAS instead of each gas separately.

$F_w\ NH_4NO_3 = 80.052\ g/mol$

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

$80.052\ g\ NH_4NO_3 = mol\ NH_4NO_3$ | $2\ mol\ NH_4NO_3 = 7\ mol\ gas\ (2+1+4)$

1 2

$$15.0\,g\ NH_4NO_3 \times \frac{mol\ NH_4NO_3}{80.052\ g\ NH_4NO_3} \times \frac{7\ mol\ gas}{2\ mol\ NH_4NO_3} = 0.655823714\ mol\ gas$$

3 $V = \dfrac{nRT}{P}$ | $n = 0.655823714\ mol\ gas$ $\quad R = 0.08206\ \dfrac{L\cdot atm}{mol\cdot K}$

$T = 300.°C = 573\ K$ $\quad P = 1.00\ atm$

$$V = \frac{(0.655823714\ mol\ gas)\left(0.08206\ \frac{L\cdot atm}{mol\cdot K}\right)(573\ K)}{(1.00\ atm)} = \boxed{30.8\ L\ at\ 300°C,\ 1.00\ atm}$$