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!

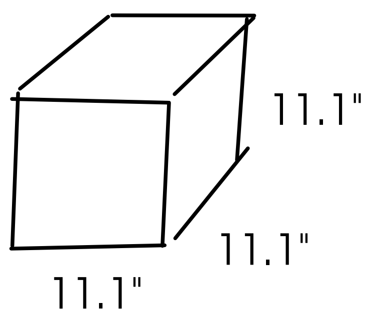
Avogadro's law:

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

1 mol gas molecules @ 0°C and 1 atm

volume = 22.4 L

"STP"
Standard Temperature and Pressure



= 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 ...

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

$$\boxed{PV = nRT}$$

P = pressure atm

V = volume L

T = ABSOLUTE temperature K

R = ideal gas constant

n = number of moles of gas molecules

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 grams 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 the IDEAL GAS EQUATION.

$84.007\ g\ NaHCO_3 = mol\ NaHCO_3$ | $2\ mol\ NaHCO_3 = 2\ mol\ CO_2$

$$25.0\ g\ NaHCO_3 \times \underbrace{\frac{mol\ NaHCO_3}{84.007\ g\ NaHCO_3}}_{1} \times \underbrace{\frac{2\ mol\ CO_2}{2\ mol\ NaHCO_3}}_{2} = 0.2975992481\ mol\ CO_2$$

$PV = nRT$

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

$n = 0.2975992481\ mol\ CO_2$ $\quad R = 0.08206 \frac{L \cdot atm}{mol \cdot K}$

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

$$V = \frac{(0.2975992481\ mol\ CO_2)(0.08206 \frac{L \cdot atm}{mol \cdot K})(298.2\ K)}{(0.950\ atm)}$$

$$= \boxed{7.67\ L\ CO_2}$$

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

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

Let's convert the volume of gas to volume at STP using the COMBINED GAS LAW...

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

$P_1 = .950 \text{ atm}$
$V_1 = 7.67 \text{ L}$
$T_1 = 298.2 \text{ K}$

$P_2 = 1 \text{ atm}$
$V_2 = ? \text{ L}$
$T_2 = 0°C = 273.2 \text{ K}$

$$V_2 = \frac{(.950 \text{ atm})(7.67 \text{ L})(273.2 \text{ K})}{(298.2 \text{ K})(1 \text{ atm})} = \boxed{6.67 \text{ L } CO_2 \text{ at STP}}$$

Alternate solution: Since we know the niumber of moles of gas in the previous problem, we COULD calculate the volume at STP using the ideal gas equation. You'll get the same answer...

$$\boxed{FW_{NH_4NO_3} = 80.0434\ g/mol}$$

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

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 this problem, let's calculate the TOTAL MOLES OF GAS instead of dealing with each gas individually.

1 - Convert 15 grams ammonium nitrate to moles using FORMULA WEIGHT
2- Convert moles ammonium nitrate to TOTAL MOLES OF GAS using CHEMICAL EQUATION
3 - Convert total moles of gas to total volume using IDEAL GAS EQUATION

$80.0434\ g\ NH_4NO_3 = mol\ NH_4NO_3$ 1

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

$$15.0\ g\ NH_4NO_3 \times \frac{mol\ NH_4NO_3}{80.0434\ g\ NH_4NO_3} \times \frac{7\ mol\ gas}{2\ mol\ NH_4NO_3} = 0.6558941774\ mol\ gas$$

3 $V = \frac{nRT}{P}$

$n = 0.6558941774\ mol\ gas$ $R = 0.08206\ \frac{L \cdot atm}{mol \cdot K}$

$T = 300°C = 573\ K$ $P = 1.00\ atm$

$$V = \frac{(0.6558941774\ mol\ gas)\left(0.08206\ \frac{L\cdot atm}{mol\cdot K}\right)(573\ K)}{(1.00\ atm)} = \boxed{30.8\ L\ gas}$$

REAL GASES

- The empirical gas laws (including the ideal gas equation) do not always apply.

    - The gas laws don't apply in situations where the assumptions made by kinetic theory are not valid.

        - When would it be FALSE that the space between gas molecules is much larger than the molecules themselves?

            - at high pressure, molecules would be much closer together!

        - When would it be FALSE that attractive and repulsive forces would be negligible?

            - at high pressure, attractions and repulsions should be stronger!

            - at low temperature, attractions and repulsions have a more significant affect on the paths of molecules

-The gas laws are highly inaccurate near the point where a gas changes to liquid!

- In general, the lower the pressure and the higher the temperature, the more IDEAL a gas behaves.