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

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

$$\boxed{P_1V_1 = P_2V_2}$$

True at constant temperature

Charles's Law:

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

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!

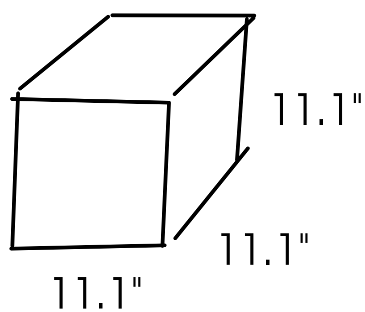
Avogadro's law:

- 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 ...

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

$$\downarrow$$

$$\boxed{PV = nRT}$$

P = pressure atm

V = volume L

T = ABSOLUTE temperature K

R = ideal gas constant

n = number of moles of gas molecules

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 of 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.

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

1 2

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

$PV = nRT$

$\downarrow$

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

$n = 0.297594248l\ mol\ CO_2$ $\quad P = 0.950\ atm$

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

$T = 25.0\ °C = 298.2\ K$

3

$$V = \frac{(0.297594248l\ mol\ CO_2)\left(0.08206\ \frac{L \cdot atm}{mol \cdot K}\right)(298.2\ K)}{(0.950\ atm)} = \boxed{7.67\ L\ CO_2\ gas}$$

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 use the COMBINED GAS LAW to solve this problem.

$$\frac{P_1V_1}{T_1} = \frac{P_2V_2}{T_2} \; ; \; \frac{P_1V_1T_2}{T_1P_2} = V_2$$

$P_1 = 0.950$ atm
$V_1 = 7.67$ L
$T_1 = 298.2$ K
$P_2 = 1$ atm
$V_2 = ?$
$T_2 = 0°C = 273.15$ K

$$V_2 = \frac{(0.950 \text{ atm})(7.67 \text{ L})(273.15 \text{ K})}{(298.2 \text{ K})(1 \text{ atm})} = \boxed{6.67 \text{ L } CO_2 \text{ at STP}}$$

Alternate solution (try this at home): Use the moles of gas calculated in the previous problem and the ideal gas law to find the volume at STP. You should get the same answer we got here (within roundoff error...)

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

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

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 calculation, we'll calculate the TOTAL MOLES OF GAS instead of trying to calculate each gas separately.

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.

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

$$15.0\,g\ NH_4NO_3 \times \frac{mol\ NH_4NO_3}{80.0434\ g\ NH_4NO_3} \times \frac{7\ mol\ gas}{2\ mol\ NH_4NO_3} = 0.655694\,1774\ mol\ gas$$

(3) $V = \frac{nRT}{P}$ | $n = 0.6556941774\ mol\ gas$ $\quad T = 300.°C = 573\ K$

$R = 0.08206\ \frac{L \cdot atm}{mol \cdot K}$ $\quad P = 1.00\ atm$

$$V = \frac{(0.6556941774\ mol\ gas)\left(0.08206\ \frac{L\cdot atm}{mol\cdot K}\right)(573\ K)}{(1.00\ atm)} = \boxed{30.8\ L}$$

As an exercise, assume volume is constant (enclosed container), and calculate the final pressure in the container. (Try a volume of 15 mL, as this amount of ammonium nitrate would fit easily into a 15 mL container...)

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.

146 van der Waals equation

- an attempt to modify PV = nRT to account for several facts.
  - gas molecules actually have SIZE (they take up space)
  - attractive and repulsive forces

$$PV = nRT$$ Ideal gas equation

$$\Big(\underbrace{P + \frac{n^2 a}{V^2}}\Big)\Big(\underbrace{V - nb}\Big) = nRT$$ van der Waals equation

attempts to account for molecular size

attempts to account for attractive / repulsive forces

* "a" and "b" are experimentally determined parameters that are different for each gas. p 208

He: a = 0,0346, b = 0,0238 tiny, no special attractive forces

H_2O: a = 5.537, b = 0,03049 small, but strong attractions between moleculres

CH_3CH_2OH: a = 12.56 b = 0,08710 larger, and strong attractions between molecules

147

$25\bar{0}0$ L of chlorine gas at 25.0 C and 1.00 atm are used to make hydrochloric acid. How many kilograms of hydrochloric acid could be produced if all the chlorine reacts?

$$H_2 + Cl_2 \rightarrow 2\,HCl$$

---

1 - Convert 2500 L chlorine gas to moles using IDEAL GAS EQUATION.
2 - Convert moles chlorine gas to moles HCl using chemical equation.
3 - Convert moles HCl to mass using formula weight.

---

1 $PV = nRT$

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

$P = 1.00 \text{ atm}$ $V = 25\bar{0}0 \text{ L}$ $R = 0.08206 \dfrac{\text{L}\cdot\text{atm}}{\text{mol}\cdot\text{K}}$

$T = 25.0^{\circ}\text{C} = 298.2 \text{ K}$

$$n_{Cl_2} = \frac{(1.00 \text{ atm})(25\bar{0}0 \text{ L})}{\left(0.08206 \frac{\text{L}\cdot\text{atm}}{\text{mol}\cdot\text{K}}\right)(298.2 \text{ K})} = 102.1646983 \text{ mol } Cl_2$$

$\text{mol } Cl_2 = 2 \text{ mol } HCl$

HCl: H: 1 x 1.008
Cl: 1 x 35.45

35.458 g HCl = mol HCl

$\text{kg} = 10^3 \text{ g}$

2

3

$$102.1646983 \text{ mol } Cl_2 \times \frac{2 \text{ mol } HCl}{\text{mol } Cl_2} \times \frac{35.458 \text{ g } HCl}{\text{mol } HCl} \times \frac{\text{kg}}{10^3 \text{ g}} = \boxed{7.245 \text{ kg } HCl}$$

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

$\uparrow O_2$

* Volume of a 10'x10'x8' room

1 - Convert 22650 L oxygen gas to moles using ideal gas equation.

2 - Convert moles gas to mass using formula weight.

① $PV = nRT$

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

$P = 1.18\ \text{atm}$

$V = 22680\ \text{L}$

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

$T = 25.0\ ^\circ\text{C} = 298.2\ \text{K}$

$$n_{O_2} = \frac{(1.18\ \text{atm})(22680\ \text{L})}{\left(0.08206\ \frac{\text{L}\cdot\text{atm}}{\text{mol}\cdot\text{K}}\right)(298.2\ \text{K})} = 1092.222357\ \text{mol}\ O_2$$

$32.00\ \text{g}\ O_2 = \text{mol}\ O_2$

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

35.0 kg

~77 lb