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

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

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

$P_1 = 0.950\text{ atm}$ $\quad P_2 = 1\text{ atm}$

$V_1 = 7.67\text{ L}$ $\quad V_2 = ?$

$T_1 = 298.2\text{ K}$ $\quad T_2 = 0°C = 273.2\text{ K}$

$$\frac{(0.950\text{ atm})(7.67\text{ L})}{(298.2\text{ K})} = \frac{(1\text{ atm})V_2}{(273.2\text{ K})}$$

$$\boxed{6.68\text{ L at STP}} = V_2$$

Alternate solution: Since we already know how many moles of carbon dioxide we have (we calculated it in the previous problem), we could also use the ideal gas equation with the STP conditions to find the new volume. (You'll get the same answer as we did here...)

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

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

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

$$15.0\text{ g } NH_4NO_3 \times \frac{\text{mol } NH_4NO_3}{80.052\text{ g } NH_4NO_3} \times \frac{7\text{ mol gas}}{2\text{ mol } NH_4NO_3} = 0.6558237146\text{ mol gas}$$

(3) $PV = nRT$

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

$P = 1.00$ atm $n = 0.6558237146$ mol gas

$R = 0.08206 \dfrac{L \cdot atm}{mol \cdot K}$ $T = 300.°C = 573$ K

$$V = \frac{(0.6558237146\text{ mol gas})\left(0.08206 \frac{L\cdot atm}{mol \cdot K}\right)(573\text{ K})}{(1.00\text{ atm})} = \boxed{30.8\text{ L}}$$

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.

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

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

nb: attempts to account for molecular size

$\frac{n^2 a}{V^2}$: 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

2500 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 2500L 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 HCl using FORMULA WEIGHT

① $PV = nRT$, $n = \frac{PV}{RT}$

$P = 1.00$ atm, $V = 25\bar{0}0$ L, $R = 0.08206 \frac{L \cdot atm}{mol \cdot K}$, $T = 25.0°C = 298.2$ K

$$n = \frac{(1.00\text{ atm})(25\bar{0}0\text{ L})}{\left(0.08206 \frac{L\cdot atm}{mol \cdot K}\right)(298.2\text{ K})} = 102.164983 \text{ mol } Cl_2$$

② mol Cl_2 = 2 mol HCl

③ HCl ~ H: 1 x 1.008
Cl: 1 x 35.45
36.458 g HCl = mol HCl

kg = 10^3 g

$$102.164983 \text{ mol } Cl_2 \times \frac{2 \text{ mol HCl}}{\text{mol } Cl_2} \times \frac{36.458 \text{ g HCl}}{\text{mol HCl}} \times \frac{\text{kg}}{10^3 \text{ g}} = 7.45 \text{ kg HCl}$$

② ③

$$2\,HCl + Na_2CO_3 \rightarrow CO_2 + H_2O + 2\,NaCl$$

If 48.90 mL of 0.250 M HCl solution reacts with sodium carbonate to produce 50.0 mL of carbon dioxide gas at 290.2 K, what is the pressure of the carbon dioxide gas?

1 - Convert 48.90 mL of 0.250 M HCl to moles using MOLARITY.
2 - Convert moles HCl to moles carbon dioxide gas using CHEMICAL EQUATION
3 - Convert moles carbon dioxide gas to pressure using IDEAL GAS EQUATION.

(1) 0.250 mol HCl = L   mL = 10^{-3} L

(2) 2 mol HCl = mol CO_2

$$48.90\text{ mL} \times \frac{10^{-3}\text{ L}}{\text{mL}} \times \frac{0.250\text{ mol HCl}}{\text{L}} \times \frac{\text{mol } CO_2}{2\text{ mol HCl}} = 0.006\,1125\text{ mol } CO_2$$

(3) $PV = nRT$

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

$n = 0.006\,1125$ mol CO_2 $R = 0.08206\ \frac{\text{L}\cdot\text{atm}}{\text{mol}\cdot\text{K}}$

$T = 290.2$ K $V = 50.0$ mL

$$50.0\text{ mL} \times \frac{10^{-3}\text{ L}}{\text{mL}} = 0.0500\text{ L}$$

$$P = \frac{(0.006\,1125\text{ mol } CO_2)\left(0.08206\,\frac{\text{L}\cdot\text{atm}}{\text{mol}\cdot\text{K}}\right)(290.2\text{ K})}{(0.0500\text{ L})} = \boxed{2.91\text{ atm}}$$