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

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

We can solve this problem using the combined gas law, since we know all properties of the gas at the original conditions.

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

$P_1 = 0.950\,atm$ $P_2 = 1.00\,atm$

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

$T_1 = 298.2\,K$ $T_2 = 0°C = 273.2\,K$

$$V_2 = \frac{(0.950\,atm)(7.67\,L)(273.2\,K)}{(298.2\,K)(1.00\,atm)} = \boxed{6.68\,L \text{ of } CO_2 \text{ gas at STP}}$$

Alternate solution: Since we knew the number of moles of carbon dioxide in the previous problem, we could have also used the ideal gas law to find the new volume. (The answer will, of course, be the same number)

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

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

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?

Shorter solution: Calculate the TOTAL MOLES of gas produced, then use ideal gas equation to find volume rather than finding moles of each gas (nitrogen, oxygen, water vapor)

1 - Convert 15.0 g of ammonium nitrate to moles. Use formula weight.
2 - Convert moles ammonium nitrate to TOTAL MOLES OF GAS. Use chemical equation
3 - Convert TOTAL MOLES OF GAS to VOLUME using ideal gas equation.

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

$$15.0\,g\ NH_4NO_3 \times \underbrace{\frac{mol\ NH_4NO_3}{80.0434\ g\ NH_4NO_3}}_{1} \times \underbrace{\frac{7\ mol\ gas}{2\ mol\ NH_4NO_3}}_{2} = 0.6558941774\ mol\ gas$$

$PV = nRT$

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

$n = 0.6558941774\ mol\ gas$

$P = 1.00\ atm$

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

$V = ?\ L$

$T = 300\ °C = 573\ K$

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

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.

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

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

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