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 change the volume at the problem's conditions to an equivalent volume at STP.

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

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

$T_1 = 298.2\text{ K}$ $T_2 = 273.2\text{ K}$

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

$$V_2 = \frac{(0.950\text{ atm})(7.67\text{ L})(273.2\text{ K})}{(298.2\text{ K})(1\text{ atm})} = \boxed{6.68\text{ L gas at STP}}$$

Alternate solution (try it!): Since we know how many moles of carbon dioxide gas we have (it's in the solution to the previous problem), we could use the ideal gas equation to directly calculate the volume at STP. You'll get the same answer as we did above.

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

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

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 trying to get each gas volume separately.

1 - Convert 15.0 grams of 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 volume using ideal gas equation.

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

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

$V = \frac{nRT}{P}$ | $n = 0.6558941774\ mol\ gas$ $\quad R = 0.08206\ \frac{L\cdot atm}{mol\cdot K}$ $\quad T = 300.°C = 573\ K$

$P = 1.00\ atm$

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

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.

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

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