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.

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

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

$V_1 = 7.67$ $\quad V_2 = ?$

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

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

Alternate solution: Since we know the number of moles of gas already, we could use it and the ideal gas equation to calculate the gas volume at STP.

Try it ... you should get the same answer reported above (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 smplify the problem, let's calculate the TOTAL MOLES OF GAS PRODUCED, since the volume of the gas doesn't depend on what KIND of gas molecules we have!

1 - Convert 15.0 g 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.

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

$$15.0\ g\ NH_4NO_3 \times \underset{(1)}{\frac{mol\ NH_4NO_3}{80.0434\ g\ NH_4NO_3}} \times \underset{(2)}{\frac{7\ mol\ gas}{2\ mol\ NH_4NO_3}} = 0.655894177 4\ mol\ gas$$

(3) $PV = nRT$

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

$n = 0.6558941774\ mol\ gas$

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

$T = 300.°C = 573\ K$

$P = 1.00\ atm$

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

The material increases in size by approximately 2000 times over the size of the solid!

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 molecülres

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

$2\bar{5}00$ 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 volume of 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) $PV = nRT$ $\quad n = \frac{PV}{RT}$

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

$V = 2\bar{5}00\text{ L}$ $\quad T = 25.0°C = 298.2\text{ K}$

$$n_{Cl_2} = \frac{(1.00\text{ atm})(2\bar{5}00\text{ L})}{\left(0.08206\ \frac{L\cdot atm}{mol\cdot K}\right)(298.2\text{ K})} = 102.1646983\text{ mol } Cl_2$$

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

H: 1×1.008
Cl: 1×35.45
$36.458\text{ g HCl} = \text{mol HCl}$

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

$$102.1646983\text{ mol } Cl_2 \times \frac{2\text{ mol HCl}}{\text{mol } Cl_2} \times \frac{36.458\text{ g HCl}}{\text{mol HCl}} \times \frac{\text{kg}}{10^3\text{ g}} = \boxed{7.45\text{ kg HCl}}$$

(2) (3)

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

↑ O_2

* Volume of a 10'x10'x8' room

1 - Convert 22650 L of oxygen gas to moles using ideal gas equation.
2 - Convert moles oxygen gas to mass using formula weight.

① $PV = nRT$

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

$P = 1.18\ atm$

$V = 22650\ L$

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

$T = 25.0°C = 298.2\ K$

$$n_{O_2} = \frac{(1.18\ atm)(22650\ L)}{\left(0.08206\ \frac{L \cdot atm}{mol \cdot K}\right)(298.2\ K)} = 1092.222357\ mol\ O_2$$

$32.00\ g\ O_2 = mol\ O_2$

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

35.0 kg

~77 lb