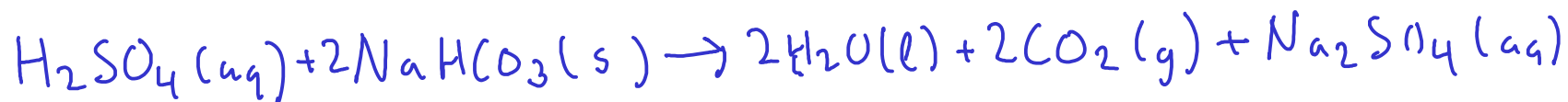


$$FW_{\text{NaHCO}_3} = 84.007 \text{ g/mol}$$



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 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

$$84.007 \text{ g NaHCO}_3 = 1 \text{ mol NaHCO}_3 \quad | \quad 2 \text{ mol NaHCO}_3 = 2 \text{ mol CO}_2$$

$$25.0 \text{ g NaHCO}_3 \times \frac{1 \text{ mol NaHCO}_3}{84.007 \text{ g NaHCO}_3} \times \frac{2 \text{ mol CO}_2}{2 \text{ mol NaHCO}_3} = 0.2975942481 \text{ mol CO}_2$$

$$PV = nRT \quad | \quad n = 0.2975942481 \text{ mol CO}_2 \quad R = 0.08206 \frac{\text{L}\cdot\text{atm}}{\text{mol}\cdot\text{K}}$$

$$V = \frac{nRT}{P} \quad | \quad T = 25.0^\circ\text{C} = 298.2 \text{ K} \quad P = 0.950 \text{ atm}$$

$$V = \frac{(0.2975942481 \text{ mol CO}_2)(0.08206 \frac{\text{L}\cdot\text{atm}}{\text{mol}\cdot\text{K}})(298.2 \text{ K})}{(0.950 \text{ atm})} = 7.67 \text{ L of CO}_2 \text{ @ } 25.0^\circ\text{C}, 0.950 \text{ atm}$$

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 solve this using the combined gas law.

$$\frac{P_1 V_1}{T_1} = \frac{P_2 V_2}{T_2} ; \quad \frac{P_1 V_1 T_2}{T_1 P_2} = V_2$$

$$\begin{array}{l|l} P_1 = 0.950 \text{ atm} & P_2 = 1 \text{ atm} \\ V_1 = 7.67 \text{ L} & V_2 = ??? \text{ L} \\ T_1 = 298.2 \text{ K} & T_2 = 273.2 \text{ K} \end{array}$$

Note: P and T in "STP" are exact, since STP is a defined condition!

$$V_2 = \frac{(0.950 \text{ atm})(7.67 \text{ L})(273.2 \text{ K})}{(298.2 \text{ K})(1 \text{ atm})} = \boxed{6.67 \text{ L CO}_2 \text{ @ STP}}$$

Alternate solution: Use $PV=nRT$ to find the volume at the new conditions. We can use the 'n' calculated in the previous problem. You should get the same answer as we have here.



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, let's calculate the TOTAL MOLES OF GAS instead of the individual moles of each gas.

$$\text{FW}_{\text{NH}_4\text{NO}_3} = 80.0434 \text{ g/mol}$$

- 1 - Convert 15.0 grams 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.

$$80.0434 \text{ g NH}_4\text{NO}_3 = 1 \text{ mol NH}_4\text{NO}_3 \quad | \quad 2 \text{ mol NH}_4\text{NO}_3 = 7 \text{ mol gas} \quad (2+1+4)$$

$$15.0 \text{ g NH}_4\text{NO}_3 \times \frac{1 \text{ mol NH}_4\text{NO}_3}{80.0434 \text{ g NH}_4\text{NO}_3} \times \frac{7 \text{ mol gas}}{2 \text{ mol NH}_4\text{NO}_3} = 0.6558941774 \text{ mol gas}$$

$$\textcircled{3} \quad V = \frac{nRT}{P} \quad \left| \quad \begin{array}{l} n = 0.6558941774 \text{ mol gas} \quad R = 0.08206 \frac{\text{L} \cdot \text{atm}}{\text{mol} \cdot \text{K}} \\ T = 300.^\circ\text{C} = 573 \text{ K} \quad P = 1.00 \text{ atm} \end{array} \right.$$

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

@ 300°C, 1 atm

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

$$PV = nRT \quad] \text{ Ideal gas equation}$$

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

attempts to account for attractive / repulsive forces

attempts to account for molecular size

* "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

H₂O: $a = 5,537$, $b = 0,03049$ small, but strong attractions between molecules

CH₃CH₂OH: $a = 12,56$ $b = 0,08710$ 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?



- 1 - Convert volume 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 (and g \rightarrow kg conversion)

$$\textcircled{1} \quad PV = nRT \quad \left| \quad P = 1.00 \text{ atm} \quad V = 2500 \text{ L} \quad R = 0.08206 \frac{\text{L}\cdot\text{atm}}{\text{mol}\cdot\text{K}} \right.$$

$$n = \frac{PV}{RT} \quad \left| \quad T = 25.0^\circ\text{C} = 298.2 \text{ K} \right.$$

$$n_{\text{Cl}_2} = \frac{(1.00 \text{ atm})(2500 \text{ L})}{(0.08206 \frac{\text{L}\cdot\text{atm}}{\text{mol}\cdot\text{K}})(298.2 \text{ K})} = 102.1646983 \text{ mol Cl}_2$$

$$\text{mol Cl}_2 = 2 \text{ mol HCl}$$

$$\text{HCl} : \text{H} : 1 \times 1.008$$

$$\text{Cl} : 1 \times 35.45$$

$$\underline{36.458 \text{ g HCl} = \text{mol HCl}}$$

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

$$102.1646983 \text{ mol Cl}_2 \times \frac{2 \text{ mol HCl}}{\text{mol Cl}_2} \times \frac{36.458 \text{ g HCl}}{\text{mol HCl}} \times \frac{\text{kg}}{10^3 \text{ g}} = \boxed{7.45 \text{ kg HCl}}$$

$\textcircled{2}$
 $\textcircled{3}$