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

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 calculation, we'll calculate the TOTAL MOLES OF GAS instead of the individual moles of each gas! $F_W\ NH_4NO_3 = 80.052\ g/mol$

1 - Convert 15.0 g ammonium nitrate to moles. Use FORMULA WEIGHT.
2 - Convert moles ammonium nitrate to TOTAL MOLES GAS. Use CHEMICAL EQUATION.
3 - Convert TOTAL MOLES GAS to volume. Use IDEAL GAS EQUATION.

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

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

3 $PV = nRT$

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

$n = 0.6558237146\ mol\ gas$ $T = 300.°C = 573\ K$

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

$$V = \frac{(0.6558237146\ mol\ gas)\times\left(0.08206\ \frac{L\cdot atm}{mol\cdot K}\right)\times(573\ K)}{(1.00\ atm)} = \boxed{30.8\ L \text{ at } 300°C, 1\ atm}$$

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

$$PV = nRT$$ Ideal gas equation

$$\left(\underbrace{P + \frac{n^2 a}{V^2}}_{\text{attempts to account for attractive / repulsive forces}}\right)\left(V - \underbrace{nb}_{\text{attempts to account for molecular size}}\right) = nRT$$ van der Waals equation

* "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

$25\bar{0}0$ 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 2500 L chlorine gas to moles. Use IDEAL GAS EQUATION.
2 - Convert moles chlorine gas to molesa HCl. Use CHEMICAL EQUATION.
3 - Convert moles HCl to mass. Use FORMULA WEIGHT.

1 $PV = nRT$

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

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

$V = 25\bar{0}0$ L

$$n_{Cl_2} = \frac{(1.00\text{ atm}) \times (25\bar{0}0\text{ L})}{(0.08206 \frac{L \cdot atm}{mol \cdot K}) \times (298.2\text{ K})} = 102.1646983 \text{ mol } Cl_2$$

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

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

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

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 HCl solution to moles. Use MOLARITY (0.250 M)
2 - Convert moles HCl to moles carbon dioxide gas. Use CHEMICAL EQUATION.
3 - Convert moles carbon dioxide gas to pressure, Use IDEAL GAS EQUATION

1 $0.250 \text{ mol HCl} = \text{L}$ $\quad \text{mL} = 10^{-3}\text{L}$ $\quad$ 2 $2 \text{ mol HCl} = \text{mol } CO_2$

$$48.90 \text{ mL} \times \frac{10^{-3}\text{L}}{\text{mL}} \times \frac{0.250 \text{ mol HCl}}{\text{L}} \times \frac{\text{mol } CO_2}{2 \text{ mol HCl}} = 0.0061125 \text{ mol } CO_2$$

3 $PV = nRT$ $\quad n = 0.0061125 \text{ mol } CO_2$ $\quad T = 290.2\text{ K}$

$P = \dfrac{nRT}{V}$ $\quad R = 0.08206 \dfrac{\text{L}\cdot\text{atm}}{\text{mol}\cdot\text{K}}$ $\quad V = 50.0 \text{ mL} = 0.0500\text{L}$

$$P = \frac{(0.0061125 \text{ mol } CO_2) \times \left(0.08206 \frac{\text{L}\cdot\text{atm}}{\text{mol}\cdot\text{K}}\right) \times (290.2\text{ K})}{(0.0500\text{L})} =$$

$$= \boxed{2.91 \text{ atm } CO_2 \text{ at } 290.2\text{ K}}$$