3. A Closer Look at Molar Volume and a Review of Stoichiometry

In 1808, Gay Lussac published experiments on combining gas volumes. Three years later Avogadro proposed an explanation for Gay-Lussac's results by proposing his famous hypothesis. That in turn paved the way for atomic masses, which were tediously worked out by Stanislao Cannizzaro in 1860.

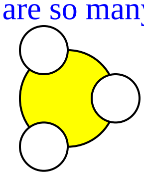
Amadeo Avogadro(1776 – 1856)

<u>Example</u>

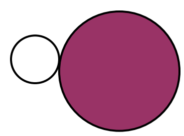
How did Avogadro's hypothesis make sense of the following results from Gay Lussac's experiments?

- 22L of $HBr_{(g)}$ reacted with 11 L of $NH_{3(g)}$ to produce 11L of $NH_4Br_{(g)}$ and leftover HBr
- 11L of $HBr_{(g)}$ reacted with 11 L of $NH_{3(g)}$ to produce 11L of $NH_4Br_{(g)}$
- 11L of $H_{2(g)}$ reacted with 11L of $Cl_{2(g)}$ to produce 22L of $HCl_{(g)}$

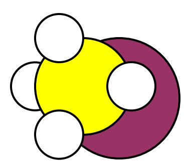
Since equal volumes contain an equal number of particles, we will represent 11 L of NH_3 with the following molecule (of course there are so many more, but our argument will hold because it's all proportional):

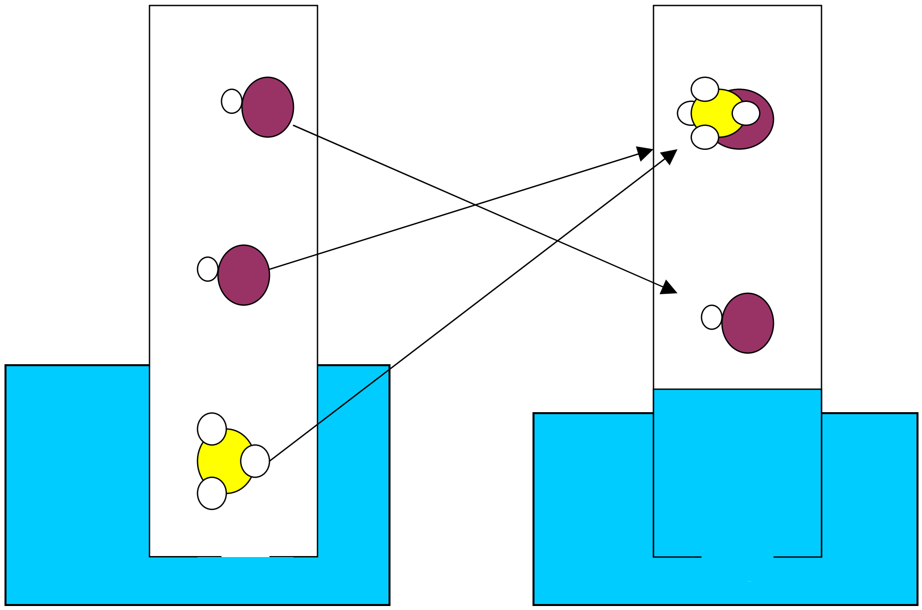


According to Avogadro, equal volumes of different gases have the same number of particles, so 11 L HBr will be represented by



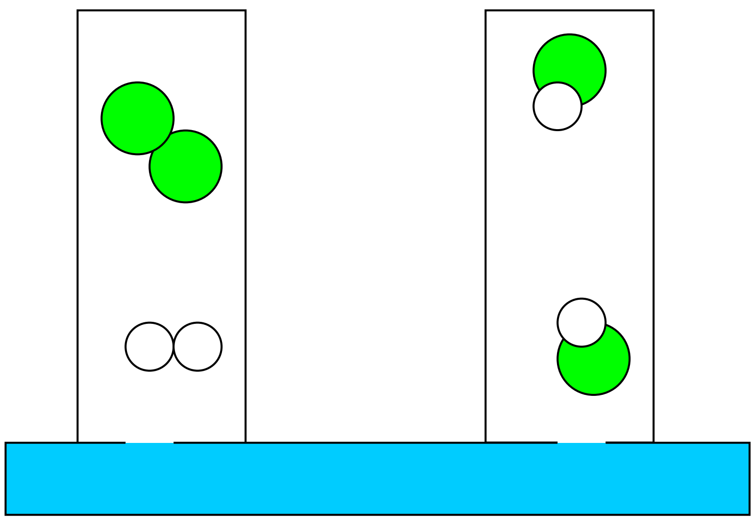
Notice that to make NH_4Br, you have to bond the two above molecules and make one single molecule. And it's consitent with the above that 11L of NH_4Br will be represented by





22 L of HBr are mixed with 11 L of NH_3. 11 L of NH_4 Br form, and since there is no more NH_3, there is 11 L of HBr leftover. If we had used 11L of each reactant, then we would have no leftover HBr. Since two molecules combine to produce only one, the total volume of reactants is reduced to half of the original.

- When you have diatomic molecules, each molecule splits into two to create two new molecules of product so that 11L of $H_{2(g)}$ reacting with 11L of $Cl_{2(g)}$ will produce 22L of $HCl_{(g)}$. Observe:

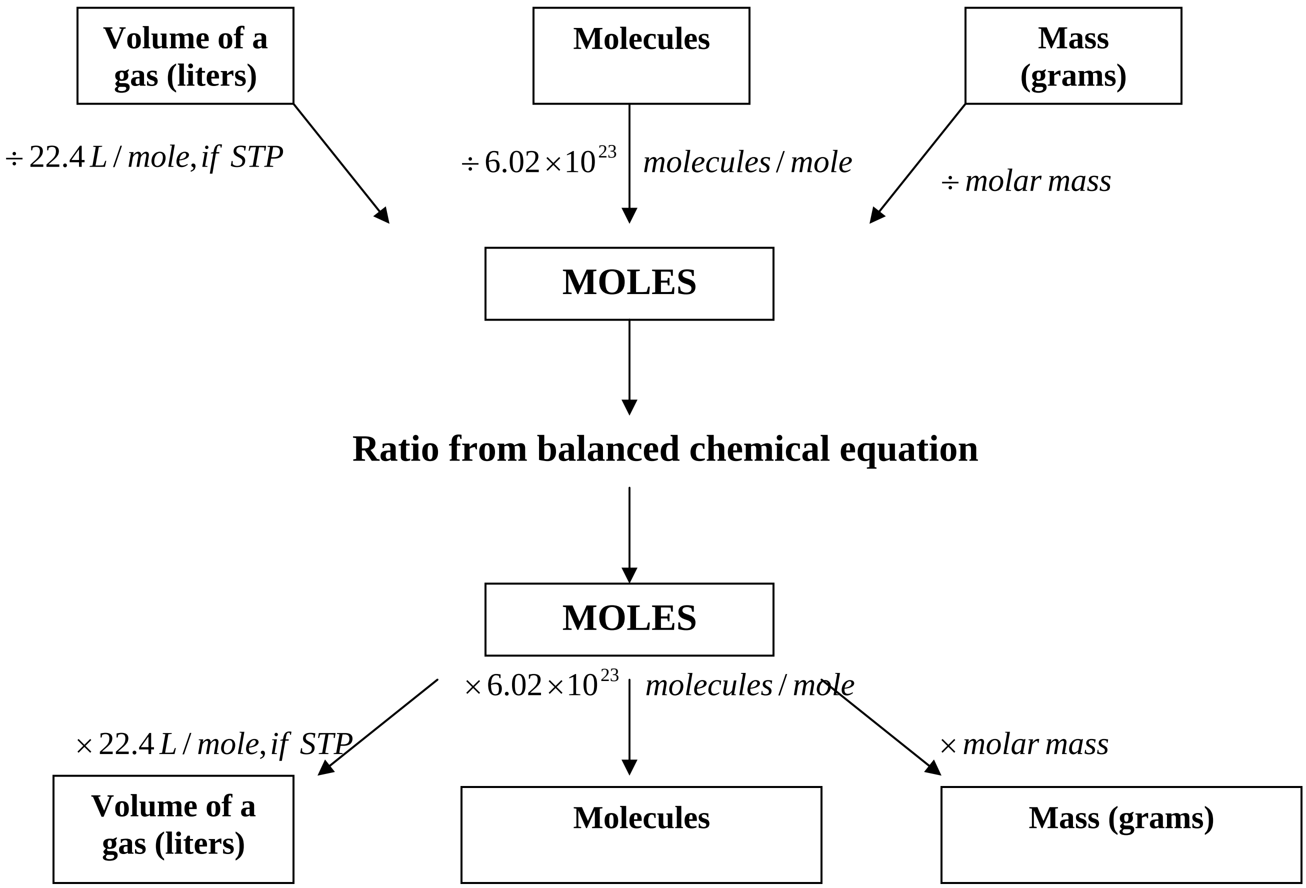


Once Avogadro's Hypothesis was accepted, scientists such as Cannizzaro had a way of getting the relative weights of different gas molecules.

Why?

By weighing the same volume of different gases (example oxygen and hydrogen) you could find out how much heavier an oxygen molecule is compared to hydrogen because you know you are comparing an equal number of molecules.

Because of Avogadro's Law, we can extend our stoichiometry chart from last year to the following:



Example 1

At STP, how many litres of hydrogen gas would you collect by reacting 0.10 g of Na?

$$2\ Na_{(s)} + 2\ HCl_{(aq)} \rightarrow 2\ NaCl_{(aq)} + 1\ H_{2(g)}$$

Answer: You need moles to compare Na to hydrogen:

0.10 g Na/ (23 g/mole) =0.00434 moles of Na

Use ratio above: 0.00434 moles of Na (1 H_2/2 $Na_{(s)}$) = 0.00217 moles of $H_{2(g)}$

At STP each mole = 22.4 L, so

0.00217 moles of $H_{2(g)}$ * 22.4 L/mole = 0.0486 L of hydrogen

Example 2

a. Find the density of Ne at S.T.P.
b. Repeat for -20 ^{o}C and 101.3 kPa.

a) d = m/V Take 1 mole of Ne = 20 g and 22.4 L at STP
= 20g/22.4 L= 0.89 g/L

b) Use Charles Law to find the volume at -20^{o}C. Then divide 20g by that calculated volume to give you the new density = 0.96 g/L. Notice it's understandably higher at a lower temperature.

Example 3 If carbon dioxide gas was allowed to reach STP before we measured its volume, which turned out to be 60.0 L, how many molecules of propane ($C_3H_{8(g)}$) were burned to create that quantity of gas?

Given: $1 C_3H_{8(g)} + 5\ O_{2(g)} \rightarrow 3\ CO_{2(g)} + 4\ H_2O_{(g)}$

Convert 60.0 L to moles by dividing by 22.4 L at STP = 2.26 moles CO_2

Apply molar ratio from balanced equation:

2.26 moles $CO_2(1 C_3H_{8(g)} / 3\ CO_{2(g)})$ = 0.893 moles $C_3H_{8(g)}$

Convert to molecules using 6.02×10^{23} molecules/ mole:

0.893 moles $C_3H_{8(g)}$ * 6.02×10^{23} molecules/ mole = 5.38×10^{23} molecules