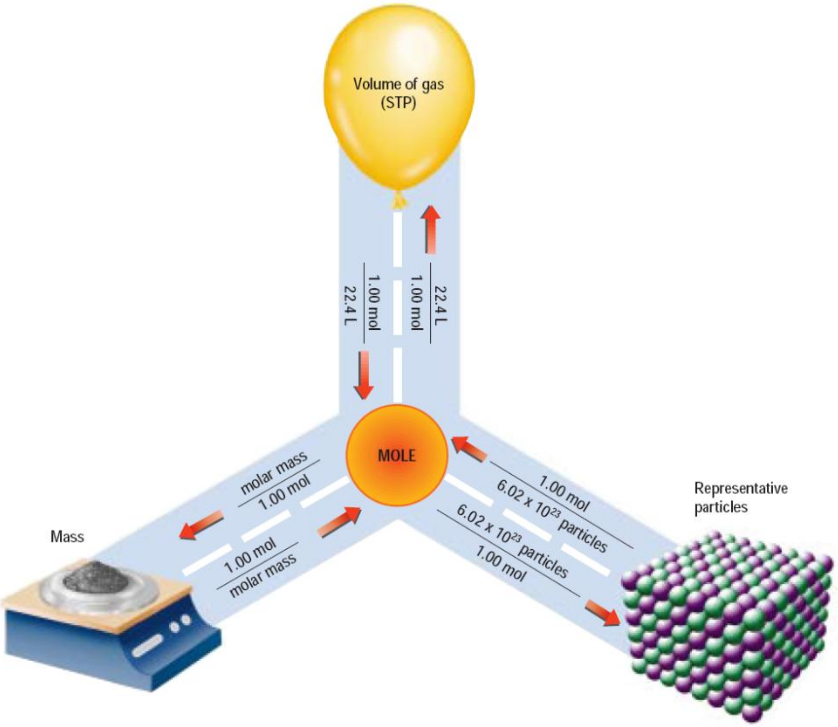
<u>NOTES: 10.2 -</u> Molar Volume, Density; and the Mole "Road Map";

<u>NOTES: 10.3</u> Percent Composition



What is a MOLE?

- VIDEO CLIP!

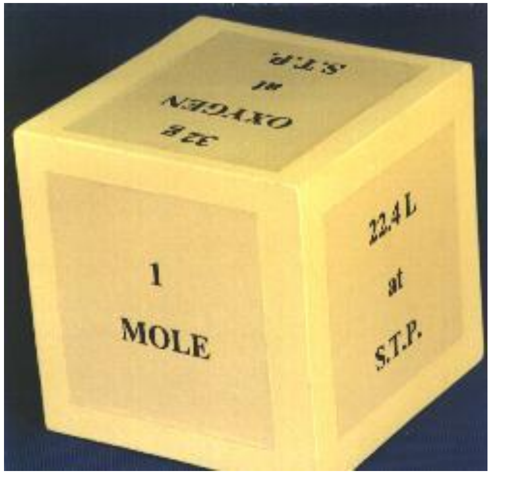
<u>Volume ↔ Mole Conversion:</u>

- **Conversion Factor:**

  -1 mole of any gas at STP = **<u>22.4 L</u>**

  -this is known as **<u>MOLAR VOLUME</u>**

  -"STP" = standard temperature ($0^{o}C$) and standard pressure (1 atm)

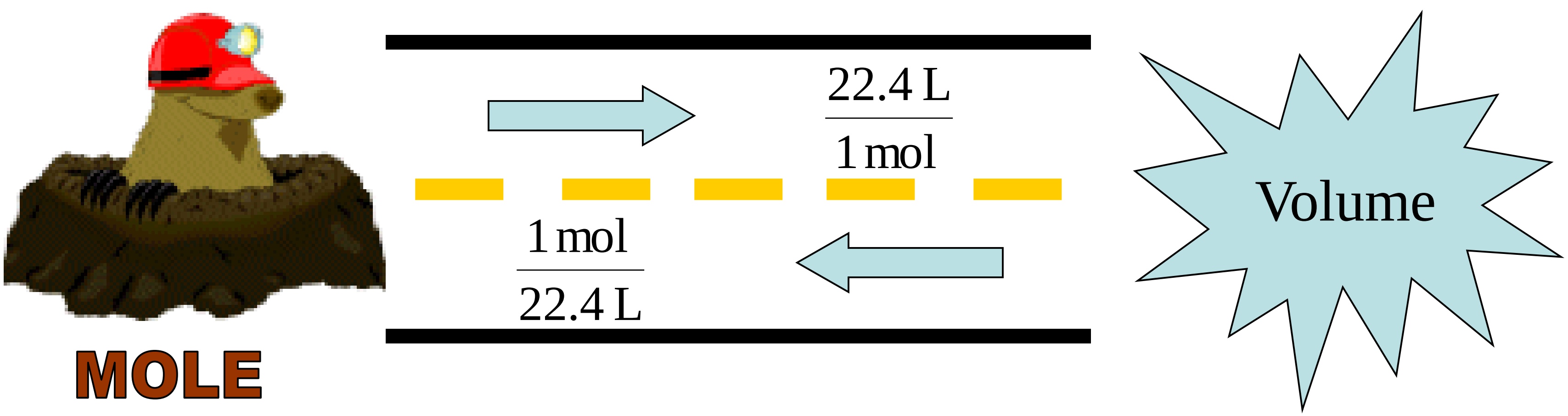


Mole ⟺ Volume Calculations

1 mole = 22.4 L of gas at STP

STP = standard temperature and pressure

(0 °C & 1 atm)



Volume Example #1:

Determine the volume, in liters, of 0.600 mol of SO_2 gas at STP.

$$\frac{0.600\,\text{mol}\ SO_2}{} \times \frac{22.4\,\text{L}}{1\,\text{mol}}$$

$$= 13.4\,\text{L}\ SO_2$$

Volume Example #2:

Determine the number of moles in 33.6 L of He gas at STP.

$$\frac{33.6\,\mathrm{L\ He}}{} \times \frac{1\,\mathrm{mol}}{22.4\,\mathrm{L}}$$

$$= 1.50\,\mathrm{mol\ He}$$

<u>DENSITY:</u>

Density = <u>Mass / Volume</u>

When given the density of an unknown gas, one can multiply by the molar volume to find the molar mass. The molar mass can then allow for identification of the gas from a list of possibilities.

Density Example (part A):

The density of an unknown gas at STP is 2.054 g/L. (a) What is the molar mass?

$$\text{density} \times \text{molar volume} = \text{molar mass}$$

$$\frac{2.054\text{g}}{\text{L}} \times \frac{22.4\text{L}}{1\text{mol}} = 46.01\,\text{g/mol}$$

Density Example (part B):

The density of an unknown gas is 2.054 g/L.
(b) Identify the gas as either nitrogen, fluorine, nitrogen dioxide, carbon dioxide, or ammonia.

molar mass = 46.0 g/mol (from part a)

Nitrogen = N_2 = 2(14.0) = 28.0 g/mol

Fluorine = F_2 = 2(19.0) = 38.0 g/mol

$\boxed{\text{Nitrogen dioxide} = NO_2 = 14.0 + 2(16.0) = 46.0 \text{ g/mol}}$

Carbon dioxide = CO_2 = 12.0 + 2(16.0) = 44.0 g/mol

Ammonia = NH_3 = 14.0 + 3(1.0) = 17.0 g/mol

Density Example #2:

The density of a gaseous compound containing carbon and oxygen is 1.964 g/L at STP. How many moles are in a 47.0 g sample of this gaseous mixture?

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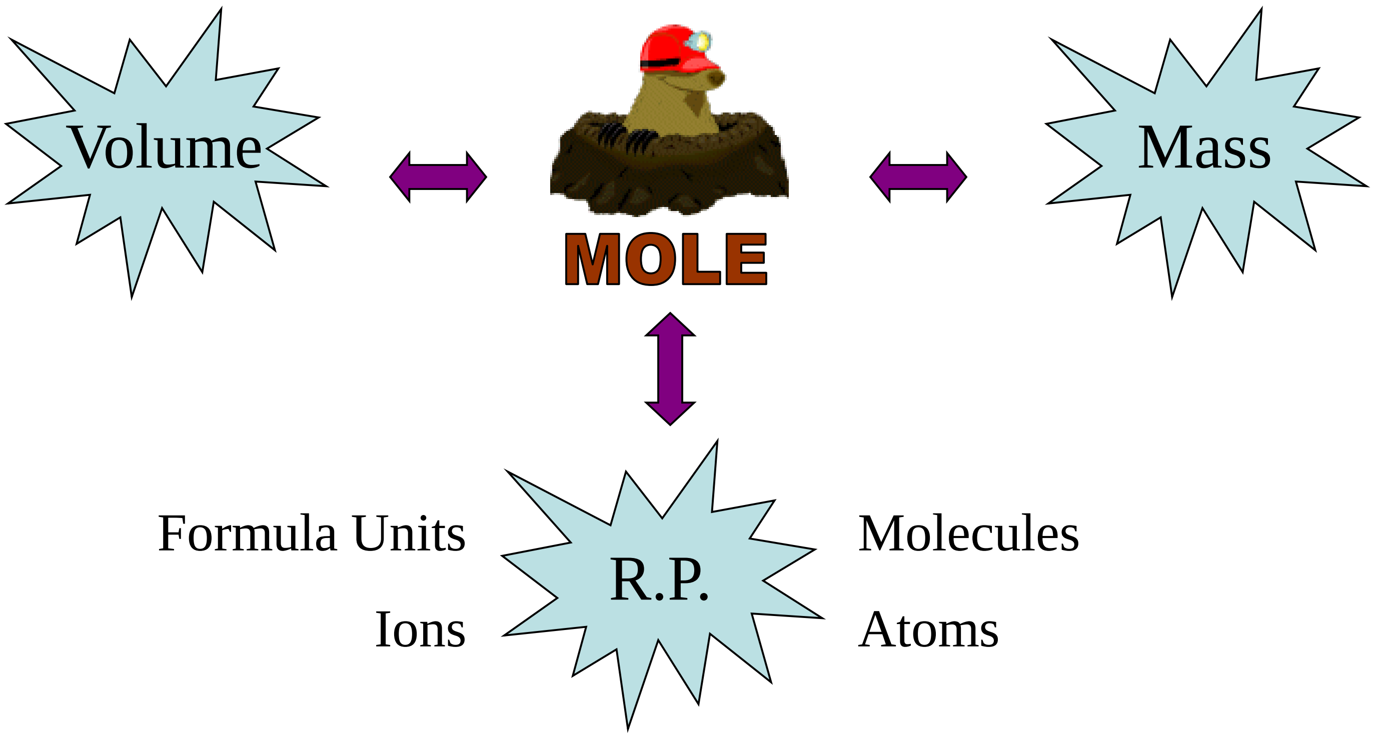
$$\frac{47.0\,\mathrm{g}}{} \times \frac{1\mathrm{L}}{1.964\,\mathrm{g}} = 23.93L$$

Density Example #2:

The density of a gaseous compound containing carbon and oxygen is 1.964 g/L at STP. How many moles are in a 47.0 g sample of this gaseous mixture?

$$\frac{23.93\,\mathrm{L}}{} \times \frac{1\,\mathrm{mol}}{22.4\,\mathrm{L}} = 1.07 mol$$

CARRY YOUR UNITS...



...AND YOUR UNITS WILL CARRY YOU!

The MOLE "Road Map"…

All Roads Lead to the Mole!!

Mixed Mole Conversions:

1 mole = <u>6.02×10^{23} RP's</u>

= <u>molar mass</u>

= <u>22.4 L of gas @STP</u>

All Roads Lead to the Mole.

Always convert to units of moles first when converting between grams, liters, and representative particles.

Mixed Mole Example #1:

How many carbon atoms are in a 50.0-carat diamond that is pure carbon? Fifty carats is the same as 10.0 g.

$$\frac{10.0\text{g}}{} \times \frac{1\,\text{mol}}{12.0\text{g C}} \times \frac{6.02\times 10^{23}\,\text{atoms}}{1\,\text{mol C}}$$

$$= 5.02\times 10^{23}\ \text{atoms C}$$

Mixed Mole Example #2:

How many **individual atoms** are in 22.0 g of water?

$$\frac{22.0\text{g } H_2O}{} \times \frac{1\text{ mol}}{18.0\text{g}} \times \frac{6.02\times 10^{23}\text{ molecules}}{1\text{ mol}} \times \frac{3\text{ atoms}}{1\text{ molecule}}$$

$$= 2.21\times 10^{24}\text{ atoms}$$

Mixed Mole Example #3:

What is the volume of a sample of nitrogen dioxide gas measured at STP that has a mass of 29.3 g?

*****NO_2 = 46.0 g/mol***

$$\frac{29.3\text{g } NO_2}{} \times \frac{1\,\text{mol}}{46.0\text{g}} \times \frac{22.4L}{1\,\text{mol}} = 14.3L$$

Mixed Mole Example #4:

What is the mass in grams of in 3.41 x 10^{22} molecules of CBr_4?

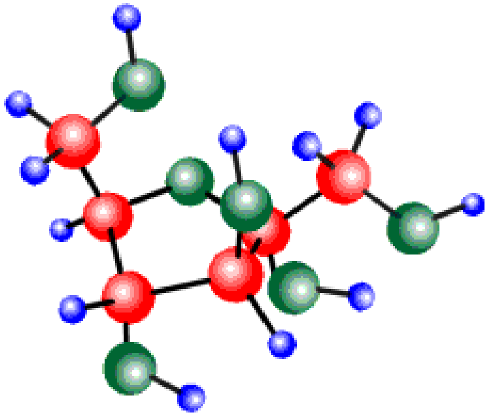
*****CBr_4 = 331.6 g/mol***

$$\frac{3.41\times10^{22}\,\mathrm{molecules}}{}\times\frac{1\,\mathrm{mol}}{6.02\times10^{23}\,\mathrm{molecules}}\times\frac{331.6g}{1\,\mathrm{mol}}$$

$$=18.8\mathrm{g}$$

Percent Composition

The chemical composition can be expressed as the mass percent of each element in the compound.



Example #1: Determine the percent composition of C_3H_8.

Assume you have one mole of the substance.

$$\%C = \frac{36.0g}{44.0g} \times 100 = \underline{81.8\%}$$

3 C = 3(12.0) = 36.0 g

$$\%H = \frac{8.0g}{44.0g} \times 100 = \underline{18.2\%}$$

8 H = 8(1.0) = 8.0 g

molar mass = 44.0 g

Example #2: Determine the percent composition of iron(III) sulfate.

Iron(III) sulfate = $Fe_2(SO_4)_3$

2 Fe = 2(55.8) = 111.6 g

3 S = 3(32.1) = 96.3 g

12 O = 12(16.0) = 192.0 g

molar mass = 399.9 g

$$\%Fe = \frac{111.6g}{399.9g} \times 100 = \underline{27.9\%}$$

$$\%S = \frac{96.3g}{399.9g} \times 100 = \underline{24.1\%}$$

$$\%O = \frac{192.0g}{399.9g} \times 100 = \underline{48.0\%}$$

Hydrated Compounds:

Some compounds exist in a "hydrated" state.

Some specific number of water molecules are present for each molecule of the compound.

Example: **oxalic acid $(COOH)_2$ can be obtained in the laboratory as $(COOH)_2 \cdot 2H_2O$.**

Note: *the dot in $(COOH)_2 \cdot 2H_2O$ shows that the crystals of oxalic acid contain 2 water molecules per $(COOH)_2$ molecule.*

Example: **oxalic acid $(COOH)_2$ can be obtained in the laboratory as $(COOH)_2 \cdot 2H_2O$.**

● The molar mass of $(COOH)_2$ = **90.0 g/mol**

$$\% \text{ anhydrous molecule} = \frac{90.0}{126.0} \times 100 = \underline{71.4\%}$$

● The molar mass of $(COOH)_2 \cdot 2H_2O$ = **126.0 g/mol**

$$\% \text{ water} = \frac{36.0}{126.0} \times 100 = \underline{28.6\%}$$

● Water can be driven out of a hydrated compound by heating it to leave an “anhydrous” (without water) compound.

Example: A 7.0 g sample of calcium nitrate, $Ca(NO_3)_2 \bullet 4H_2O$, is heated to constant mass. How much anhydrous salt remains?

% mass anhydrous salt =

$$\frac{164.1\text{g}}{236.1\text{g}} \times 100 = 69.5\%$$

Mass of anhydrous salt remaining =

$$.695 \times 7.0\text{g} = 4.9\ \text{g}$$