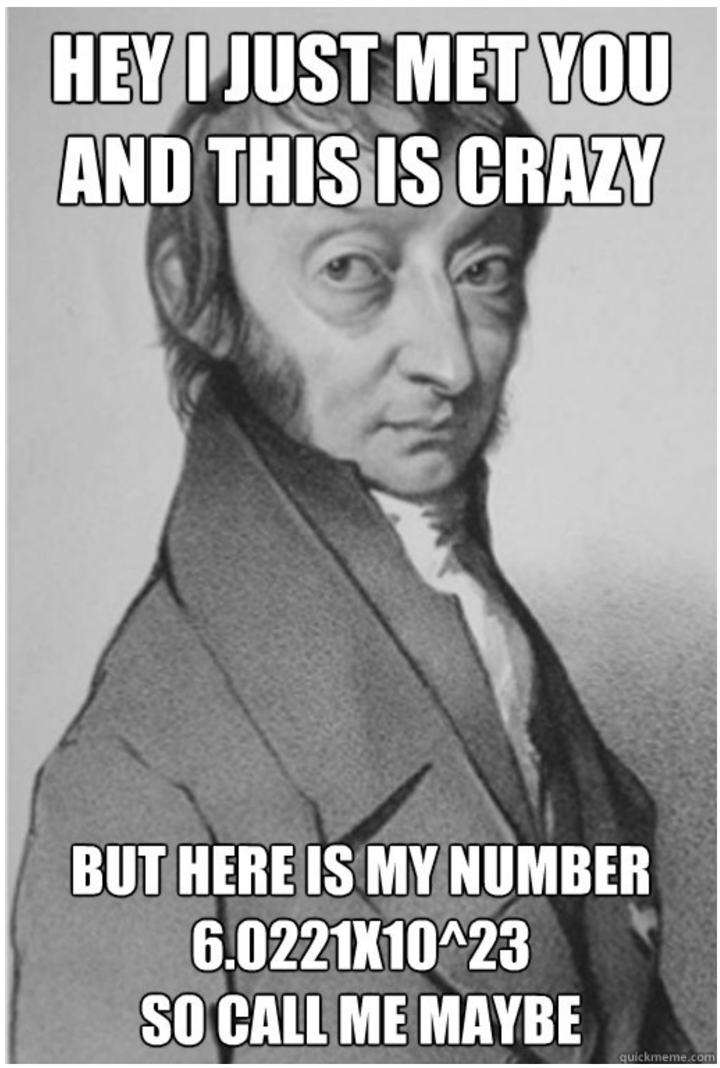
CHEMISTRY 11

UNIT 4: THE MOLE



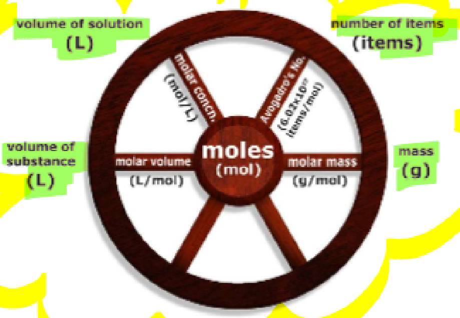
BOOK 1: INTRODUCTION TO THE MOLE

Name:___KEY_________ Block:______

The Mole — The Central Unit of Chemistry units

What mass of oxygen has the same number of atoms as 1 g of hydrogen? An oxygen atom (16 u) weighs 16 times as much as a hydrogen atom (1 u). Therefore, it would require 16 g of oxygen to have the same number of atoms as 1 g of hydrogen. Chemists extended this reasoning to all the elements. For example, 55.8 g Fe, 35.5 g Cl, 23.0 g Na, and 12.0 g C contain the same number of atoms. since these masses are in the same ratios as their individual atomic masses. **How many atoms are there in the atomic mass of any element expressed in grams?** *Originally chemists didn't know and even now they only have a very rough estimate but they nevertheless gave a name to that number.* **They called this number a "MOLE."**

mole = mol ← (unit symbol)



The mole is at the centre of the chemical measurement.

The Mole Concept

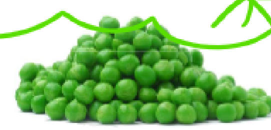
A **mole** is ... a quantity equal to the number of atoms in the atomic mass of ANY element, expressed in grams (g)

(e.g., the number of atoms in 1.0 g H, 16.0 g O, 63.5 g Cu).

NOTE: 1 mole ≠ 1 gram

1 mole = the number of particles in the atomic mass of an element

"How much is a mole?", you ask.... THE GREEN PEA ANALOGY



If you were to select one hundred (10^2) average-size peas, you would find that they occupy roughly a volume of 20 cm³. One million peas (10^6) are just enough to fill an ordinary household refrigerator and a billion (10^9) peas will fill a three bedroom house from basement to attic. A trillion (10^{12}) peas will fill a thousand houses, the number you might find in a small town. A quadrillion (10^{15}) peas will fill all of the buildings in a city the size of Victoria.

Obviously you will run out of buildings very soon. Let us try a larger measure. Say there is a blizzard over all the western provinces, except that instead of snowing snow, it snows peas. All of British Columbia, Alberta, and Saskatchewan lie covered to a depth of 1 metre. The blanket of peas drifts across roads, banks up against the sides of houses, and covers all the fields and forests. Think of flying across the province with this blanket of peas extending as far as you can see. This gives you an idea of our next number. In the entire blanket there are about a quintillion (10^{18}) peas.

Imagine that this blizzard falls over the entire land surface of the planet! North America, South America, Africa, Europe, Asia, Australia and Antarctica are all buried one metre deep. This global blanket contains about one sextillion (10^{21}) peas. Then imagine that the oceans are frozen over and the blanket now covers the entire land and sea area of the Earth to a one metre depth. Go out among the neighbouring stars and collect 250 planets the size of Earth and cover each of them with a blanket of peas one metre deep. Then you will have a mole of peas.

Furthermore, go out into the farthest reaches of the Milky Way and collect 250 000 planets, each the size of the Earth, and cover them with a blanket of peas one metre deep. You now have about one octillion (10^{27}) peas – which is roughly the number of atoms which make up your body.

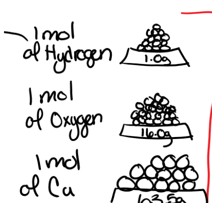
– adapted from the original Green Pea Analogy (Author unknown).

The number of **things** (particles) in a mole IS Avagadro's number 6.02×10^{23}

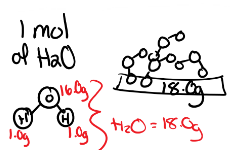
For Example:

- *A **mole** is the number of carbon atoms in exactly 12g of Carbon (atomic mass of carbon)* ← molar mass of carbon
- *1 **mole** of sulphur atom is exactly 32.1 g*
- *1 molecule of H_2O contains 2 moles of hydrogen and 1 mole of oxygen....but it is also considered to be 1 mole of water.*

When we say a mole refers to *"the number of things"*, those "things" can be atoms, molecules, compounds, etc...



= 1 mol = 6.02×10^{23} atoms



1 mol = 6.02×10^{23} molecu

1 mol of anything = 6.02×10^{23} of anything (atoms, molecules ...)

1 water particle / molecule / compound = 1 "thing" = 1 mol of H_2O

$2(1.0g) + 1(16.0g) = 18.0g$

Introducing Molar Mass

Experimental work by the English chemist John Dalton (1766–1844) was concerned with how much of one element could combine with a given amount of another element. He put forth the following hypotheses.

- Molecules are made up of " atoms " of various elements.
- If compound B contains twice the mass of element X as does compound A, then compound B must contain twice as many atoms of X.

Dalton did not attempt to figure out the mass of an individual atom of any element. Instead he simply assigned an arbitrary mass to each element, assuming that hydrogen was the lightest element and therefore could be assigned a mass of "1".

Figure 3.1.2 *The mass of an oxygen atom is equal to the mass of 16 hydrogen atoms.*

Dalton's experiments found that Carbon was 6 times heavier than Hydrogen, so C was given the mass 6 (we know C = 12)

Similarly, Oxygen was 16 times heavier than Hydrogen, and was assigned a mass of 16.

In this way Dalton was able to calculate relative masses for several different elements.

(relative to hydrogen)

(* atomic mass = "u")

Molar Mass (g)

The mass of one mole of an element's atoms is called that element's 'molar mass'. It follows from simply restating the definition of a mole that the molar mass of an element is its atomic mass expressed in grams.

For example, "one mole is the number of atoms in 16 g of oxygen" can be restated as "one mole of oxygen atoms weighs 16 g."

grams

The atomic mass of the elements can be found in the Periodic Table.

The atomic mass of oxygen is 16 u and thus the molar mass of oxygen is 16.0 g

This is better expressed as a conversion factor for calculation purposes:

16 g per mole of oxygen or $\frac{16g}{1 mol}$ or $\frac{1 mol}{16g}$

O_2 (16u)–(16u) molecular ~~molar~~ mass $2 \times (16) = 32.0g$ molecule

if I have 1 oxygen: O (16u) molar mass $1 \times (16u) = 16.0g$ #1

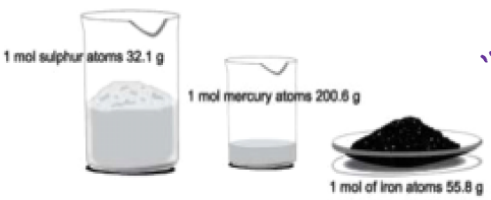


Figure 3.2.3 *The mass of 1 mol of a chemical depends on the atoms that make it up.*

"The molecular mass or "formula mass" of a compound is the sum (+) of its constituent atomic masses

For Example: H_2O: 2 × H and 1 × O

One mole of water molecules consists of 1mol of oxygen atoms (16 g) and 2mol of hydrogen atoms (2 g) and therefore weighs 18 g.

~~1 mole of molecules consists of ___ of oxygen atoms (___ g) and ___ of hydrogen atoms (2 g) and therefore has a mass of 18 g.~~

1 mol of NaCl formula units consists of 1 mol of sodium atoms (23.0 g) and 1 mol of chlorine atoms (35.5 g) for a total mass of 58.5 g

molecular mass

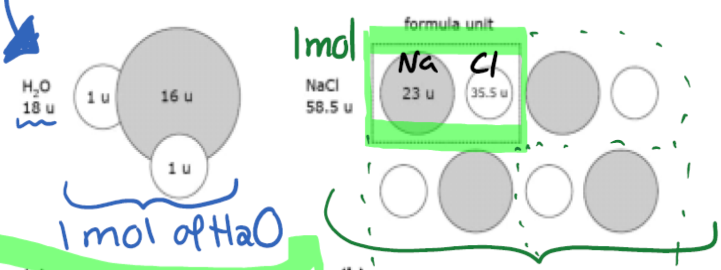


Figure 3.2.4 *(a) The molecular mass of water is the sum of the masses of the oxygen and hydrogen atoms. (b) The formula mass of NaCl is the sum of the masses of sodium and chlorine atoms.*

Just as the molar mass of an element is simply its atomic mass expressed in grams, the **molar mass** of a compound is simply its molecular or formula mass expressed in grams.

1 mol of H = 1.0g = 6.0221×10^{23} H atoms

1 mol of H_2O = 18.0g = 6.0221×10^{23} H_2O molecules

1 mol = 6.0221×10^{23} of something

6.02 x 10^{23}

The **MOLAR MASS** is the mass of ONE MOLE of particles.

This definition leads to the following statement.

$\boxed{\text{the mass shown in your Periodic Table, but represented in grams (g.)}}$

EXAMPLES:

Element	Atomic mass shown on periodic table (amu or u)	Molar mass of element (g)
C	12.0 u	12.0 g
Fe	55.8 u	55.8 g
S	32.1 u	32.1 g

* **IMPORTANT:** Unless specifically asked to use more precise values, always use masses rounded off to one decimal place. 2-3sf. The masses of the elements are given in the ***Periodic Table of the Elements*** and the table ***Atomic Masses of the Elements*** at the back of this book.

Sample Problem — Determining a Compound's Molar Mass

What are the **atomic mass** (amu) and **molecular mass** of $Al_2(SO_4)_3$?

Al : 27.0u
S = 32.1 u
O = 16.0u

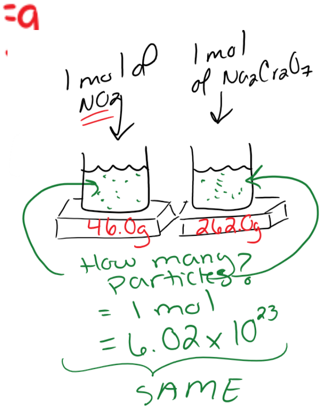
What to Think About	How to Do It
1. 1 $Al_2(SO_4)_3$ consists of 2 Al's, 3 S's, and 12 O's.	Molecular Mass = 2×(Al) + 3×(S) + 12(O) 2(27.0u) + 3(32.1u) + 12(16.0u)
2. 1 mol $Al_2(SO_4)_3$ consists of 2 mol Al, 3 mol S and 12 mol O.	Molecular Mass (g) of $Al_2(SO_4)_3$ = 342.3g

NOTE: **Molecular mass** is the *mass of a molecule:* the sum of the atomic weights of each element multiplied by the number of atoms of that element in the molecular formula.

atomic mass = amu or u   molar or molecular mass = g

PRACTICE — Determining a Compound's Molar Mass

1. What is the molecular mass (g) of nitrogen dioxide? NO_2
   (1×N) + (2×O)
   (1×14.0u) + (2×16.0u) = 46.0g
2. What is the molecular mass (g) of $Na_2Cr_2O_7$? (2×Na) + (2×Cr) + (7×O)
   (2×23.0u) + (2×52.0u) + (7×16.0u) = 262.0g
3. What is the molecular mass (g) of iron(III) sulphide? (2×Fe) + (3×S)
   Fe^{3+} S^{2-} 2:3 Fe_2S_3
   (2×55.8u) + (3×32.1u) = 207.9g



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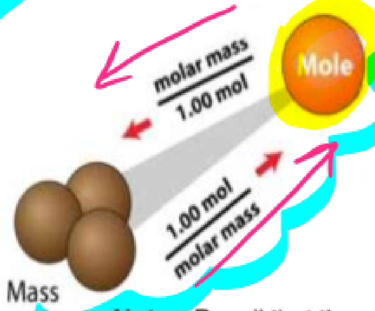
Assignment #1- Hebden Questions page 80
6 (a,e,i,m,c,g,k,o) #7 a + c
Complete ALL assignments on a seperate piece of paper and attach to your booklet when handing in at the end of the unit.

atomic mass, expressed in grams (g)

Conversions Using Molar Mass



Name	Equivalence Statement	Conversion Factors	
Molar mass	1 mol = ? g	$\frac{?\ g}{1\ mol}$	$\frac{1\ mol}{?\ g}$
Example: H_2O	1 mol = 18.0 g	$\frac{18.0g}{1mol}$	$\frac{1mol}{18.0g}$

The use of the molar mass allows you to calculate the mass of a substance when given number of moles, and to find the number of moles of a substance, when given the mass.

mass (g) ⟷ mole (mo

Note: Recall that the unit symbol for "mole" is mol

Example: three moles = 3 mol

~~mols~~ ~~cmls~~

Under **no** circumstances must you use any other unit symbol for the unit "mole", because all other possibly–suitable abbreviations (m, mo, M, Mo) already have special meanings in chemistry. Using any unit symbol besides "mol" will cause you great confusion later on in this course.

oxygen atoms = O

oxygen gas = O_2

diatomic molecules

HOFBrINCl

$H_2O_2F_2$

Sample Problem — Converting Moles to Mass

What is the mass of 3.2 mol of oxygen atoms?

(1 mol) molar mass of O = 16.0g } periodic table.

What to Think about	How to Do It
1. Convert: mol O → g O 2. Setup: 3.2 mol O × $\frac{?\ g\ O}{1\ mol\ O}$ 3. Conversion factor: 16.0 g O per 1 mol O	convert mol → mass $\frac{1mol}{16.0g}$ or $\frac{16.0g}{1mol}$ $3.2mol \times \frac{16.0g}{1\ mol}$ = 51.2 g = $\boxed{51g\ of\ O}$ 2sf 3sf. 2sf. conversion factor

1 mol of X has a mass of (molar mass X) g produces the conversion factors

$\frac{1\ mol}{(molar\ mass\ of\ X)\ g}$ or $\frac{(molar\ mass\ of\ X)\ g}{1\ mol}$

EXAMPLE: **What is the mass of 1.36 x 10^{-3} mol of SO_3?** → molar mass S = 32.0g, 3×O = 3(16.0g), 80.1g.

$1.36\times10^{-3}mol \times \frac{80.1g}{1mol} = 1.09\times10^{-1}g\ SO_3$ (3sf.)

EXAMPLE: **How many moles of N_2 are there in 50.0 g of N_2?** → molar mass of N_2(g) 2(14.0g) $\boxed{1mol = 28.0g}$

mass → mol

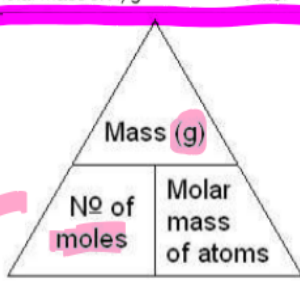
$50.0g\ \cancel{N_2} \times \frac{1\ mol}{28.0g\ \cancel{N_2}} = 1.79\ mol\ N_2$

EXAMPLE: **How many moles of CH_3OH are there in 0.250 g of CH_3OH?**

C·1(12.0g)
H·4(1.0g)
O·1(16.0g)
32.0g

$0.250g \times \frac{1mol}{32.0g}$ = 0.0078125..... = 7.81×10^{-3} mol

3s.f.   3s.f.



The units of Molar Mass are actually: $\frac{g}{mol}$. This means that the molar mass can be calculated by dividing the mass of a substance by the number of moles contained in the substance.

1 mol = 43.64g

EXAMPLE: **If 0.140 mol of acetylene gas has a mass of 3.64 g, what is the molar mass of acetylene?** → $\frac{g}{mol}$

$\boxed{Molar\ Mass = \frac{mass(g)}{no.\ of\ mol}}$

$no.\ of\ mol = \frac{mass(g)}{molar\ mass}$

$mass(g) = (no.\ of\ mol)(molar\ mass)$

$Molar\ mass = \frac{3.64g}{0.140mol} = 26.0\frac{g}{mol}$

acetylene = C_2H_2 → check 2(12.0) 2(1.0) = 26.0g ✓

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Assignment #2- Mole problems #0 Worksheet
Complete ALL assignments on a separate piece of paper and attach to your booklet when handing in at the end of the unit.

Mole Problems #0 - Answer Key

1\. a) Fe_2O_3

$2 \times Fe = 2\,mol \times 55.8\ g/mol = 111.6\ g$
$3 \times O = 3\,mol \times 16.0\ g/mol = 48.0\ g$
ADD = $\boxed{159.6\ g}$

b) H_3PO_4

$3 \times H = 3\,mol \times 1.0\ g/mol = 3.0\ g$
$1 \times P = 1\,mol \times 31.0\ g/mol = 31.0\ g$
$4 \times O = 4\,mol \times 16.0\ g/mol = 64.0\ g$
ADD = $\boxed{98.0\ g}$

c) Be_5As_2

$5 \times Be = 5\,mol \times 9.0\ g/mol = 45.0\ g$
$2 \times As = 2\,mol \times 74.9\ g/mol = 149.8\ g$
ADD = $\boxed{194.8\ g}$

d) Rb_2SO_3 = Rubidium sulfite

$2 \times Rb = 2\,mol \times 85.5\ g/mol = 171.0\ g$
$1 \times S = 1\,mol \times 32.1\ g/mol = 32.1\ g$
$3 \times O = 3\,mol \times 16.0\ g/mol = 48.0\ g$
ADD = $\boxed{251.1\ g}$

e) $Al_2(SO_4)_3$ = Aluminum sulfate

$2 \times Al = 2\,mol \times 27.0\ g/mol = 54.0\ g$
$3 \times S = 3\,mol \times 32.1\ g/mol = 96.3\ g$
$12 \times O = 12\,mol \times 16.0\ g/mol = 192.0\ g$
ADD = $\boxed{342.3\ g}$

f) $Mg(OH)_2$: Magnesium hydroxide

$1 \times Mg = 1\,mol \times 24.3\ g/mol = 24.3\ g$
$2 \times O = 2\,mol \times 16.0\ g/mol = 32.0\ g$
$2 \times H = 2\,mol \times 1.0\ g/mol = 2.0\ g$
ADD = $\boxed{58.3\ g}$

2\. a) $\dfrac{2.50\ mol\ K_2CrO_4}{} \left| \dfrac{194.2\ g\ K_2CrO_4}{1\ mol\ K_2CrO_4} \right. = 485.5\ g\ K_2CrO_4$

$\boxed{= 486\ g\ K_2CrO_4\ \text{(SIG FIGS)}}$

b) $\dfrac{0.25\ mol\ Ba(NO_3)_2}{} \left| \dfrac{261.3\ g\ Ba(NO_3)_2}{1\ mol\ Ba(NO_3)_2} \right. = 65.3\ g\ Ba(NO_3)_2$

$\boxed{= 65\ g\ Ba(NO_3)_2}$

c) $\frac{0.375 \text{ mol } Na_2Cr_2O_7}{} \left| \frac{262 \text{ g } Na_2Cr_2O_7}{1 \text{ mol } Na_2Cr_2O_7} \right. = \boxed{98.3 \text{ g } Na_2Cr_2O_7}$

d) $\frac{0.25 \text{ mol } NaCH_3COO}{} \left| \frac{82 \text{ g } NaCH_3COO}{1 \text{ mol } NaCH_3COO} \right. = \boxed{21 \text{ g } NaCH_3COO}$

e) $\frac{0.418 \text{ mol } Fe(NO_3)_3}{} \left| \frac{241.8 \text{ g } Fe(NO_3)_3}{1 \text{ mol } Fe(NO_3)_3} \right. = \boxed{101 \text{ g } Fe(NO_3)_3}$

f) $\frac{1.872 \text{ mol } Cu(CH_3COO)_2}{} \left| \frac{181.5 \text{ g } Cu(CH_3COO)_2}{1 \text{ mol } Cu(CH_3COO)_2} \right. = \boxed{339.8 \text{ g } Cu(CH_3COO)_2}$

3\. a) $\frac{50.0 \text{ g } C_6H_{12}O_6}{} \left| \frac{1 \text{ mol } C_6H_{12}O_6}{180 \text{ g } C_6H_{12}O_6} \right. = \boxed{0.278 \text{ mol } C_6H_{12}O_6}$

b) $\frac{25.00 \text{ g } K_3PO_4}{} \left| \frac{1 \text{ mol } K_3PO_4}{212.3 \text{ g } K_3PO_4} \right. = \boxed{0.1178 \text{ mol } K_3PO_4}$

c) $\frac{15.57 \text{ g } Bi(OH)_3}{} \left| \frac{1 \text{ mol } Bi(OH)_3}{260.0 \text{ g } Bi(OH)_3} \right. = \boxed{0.05988 \text{ mol } Bi(OH)_3}$

d) $\frac{3.50 \text{ g } AsCl_3}{} \left| \frac{1 \text{ mol } AsCl_3}{181.4 \text{ g } AsCl_3} \right. = \boxed{0.0193 \text{ mol } AsCl_3}$

e) $\frac{27.85 \text{ g } Fe_3(PO_4)_2}{} \left| \frac{1 \text{ mol } Fe_3(PO_4)_2}{357.4 \text{ g } Fe_3(PO_4)_2} \right. = \boxed{0.07792 \text{ mol } Fe_3(PO_4)_2}$

f) $\frac{4.90 \text{ g } Al_2(CO_3)_3}{} \left| \frac{1 \text{ mol } Al_2(CO_3)_3}{234 \text{ g } Al_2(CO_3)_3} \right. = \boxed{0.0209 \text{ mol } Al_2(CO_3)_3}$

Part B: Molar Volume

of mol ⟷ volume of a gas @ STP

Figure 3.4.1 *Gay-Lussac was an avid hot-air balloonist and conducted some of his experiments aloft.*

Just as the mass of a mole of a substance is called its molar mass, the ***volume of a mole*** of a substance is called its molar volume.

The **molar volume** of a substance is the space occupied by 1 mole of it's particles. A solid's or a liquid's molar volume is determined by the size and spacing of its particles. The size of the particles has little effect on a gas's molar volume because the average distance between the particles is so much greater than their size.

↑ K.E.

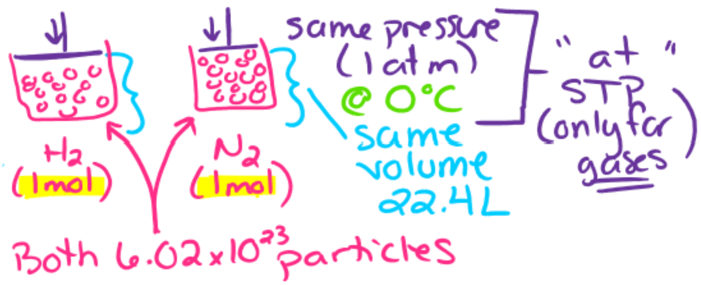
At a higher temperature, a substance's particles are moving faster and are thereby hitting each other harder and bouncing farther apart. Since its particles have spread farther apart, a substance's molar volume is greater at higher temperatures.

Liquids and gases are more frequently measured by volume than by mass. A substance's molar volume allows you to convert the volume of the substance into its number of moles.

Avogadro's Hypothesis: Equal volumes of different gases, at the same temperature and pressure, contain the same number of particles. $= 6.02 \times 10^{23}$

STANDARD **T**EMPERATURE AND **P**RESSURE (**STP**) = 0°C and 101.3 kPa. → 1 mol = 22.4 L

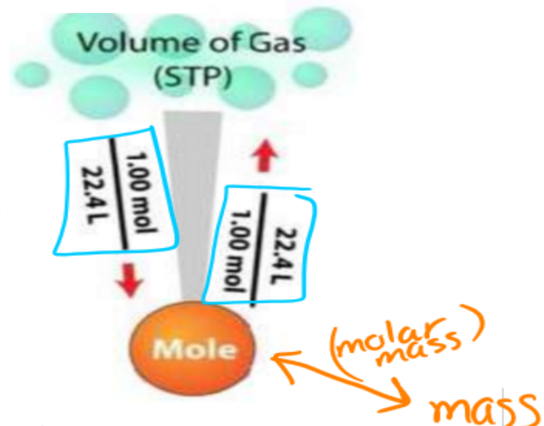
Meaning: *All gas samples with the same pressure, temperature and number of particles occupy identical volumes. This implies* equal number of moles *of every gas at STP occupy identical volumes.*



Experimentally–determined fact:

$\boxed{\text{1 mol of ANY GAS at STP has a volume of 22.4 L.}}$

In other words **the MOLAR VOLUME of *any* gas (g) at STP is 22.4 L.**



$\boxed{\text{Conversion factor: } \frac{1\text{ mol}}{22.4\text{ L}} \text{ or } \frac{22.4\text{ L}}{1\text{ mol}}}$

***NOTE:** These conversion factors **ONLY** apply to **gases** and **only at STP.**

EXAMPLE: **What is the volume occupied by 0.350 mol of SO_2(g) at STP?**

given mol → volume   1 mol = 22.4 L

$$\frac{0.350\,\cancel{\text{mol}}}{} \times \frac{22.4\text{ L}}{1\,\cancel{\text{mol}}} = \underline{7.84}\text{ L of } \boxed{SO_2(g)}$$

EXAMPLE: **How many moles of gas are contained in a balloon with a volume of 10.0 L at STP?**

given vol. → mol   1 mol = 22.4 L

$$\frac{10.0\,\cancel{\text{L}}}{} \times \frac{1\text{ mol}}{22.4\,\cancel{\text{L}}} = \underline{0.446}\text{ mol}$$

3 s.f.   3 s.f. $(4.46 \times 10^{-1}\text{ mol})$

chemistry homework

Assignment #3- Practice Problems #1-3 & Hebden Questions #11-12 page 83
Complete ALL assignments on a separate piece of paper and attach to your booklet when handing in at the end of the unit.

PRACTICE

Practice Problems — **Converting Moles to Volume and Volume to Moles**

1. What volume of oxygen gas at STP contains 1.33 mol of O_2?

$1.33\ \cancel{mol} \times \frac{22.4L}{1\ \cancel{mol}} = 29.8L$

2. In British Columbia, the burnt-match odor of sulphur dioxide is often associated with pulp and paper mills. How many moles of SO_2 are in 9.5 L of SO_2 at STP?

$9.5\cancel{L} \times \frac{1\ mol}{22.4\cancel{L}} = 0.42\ mol$

3. Silicon dioxide, better known as quartz, has a molar volume of 22.8 cm^3/mol. What is the volume of 0.39 mol of SiO_2?

NOT @STP

$\frac{22.8cm^3}{1\ mol} = \frac{22.8mL}{1\ mol} = \boxed{\frac{0.0228L}{1\ mol}}$

conversion factor

In this question we need the conversion factor b/c NOT @STP.

$0.39\ \cancel{mol} \times \frac{0.0228L}{1\ \cancel{mol}} = 0.008892L$

$= 8.9 \times 10^{-3}L$

* 2 s.f.

Part C: Avagadro's Number

The number of things in a mole is also referred to as Avagadro's Number in honour of the Italian scientist whose insight regarding gases led to a technique for determining the relative atomic masses of non-metals.

Just as a dozen is 12 of anything, a mole is approx. 6.02×10^{23} of anything. While a dozen is a fairly small number, a mole is an absurdly large number. As you know, a mole of peas would cover the entire Earth's surface with a layer over 200 m deep.

1 mol of carbon 12g

Figure 3.2.1 $6.02214179 \times 10^{23}$ carbon atoms

Just as a dozen is a convenient unit of quantity for a baker to group buns and doughnuts, a mole is a convenient "unit of quantity" for a chemist to group atoms and molecules.

Chemists currently estimate that a mole is $6.02214179 \times 10^{23}$ give or take a few million billion.

The actual number isn't important unless you're working at the atomic level because whatever the number is, it's the same for a mole of anything (1 mol of H or 1 mol of C $= 6.02 \times 10^{23}$ atoms)

Most chemical conversions involve the mole.

The key to conversion is the conversion factor.

Chemists know or know where to find the conversion factors they need.

Name	Equivalence Statement	Conversion Factors	
Avogadro's number	1 mol = 6.02×10^{23} particles "items"	$\frac{6.02 \times 10^{23}\ items}{1\ mol}$	$\frac{1\ mol}{6.02 \times 10^{23}\ items}$

atoms
particles
molecules
grandma's....

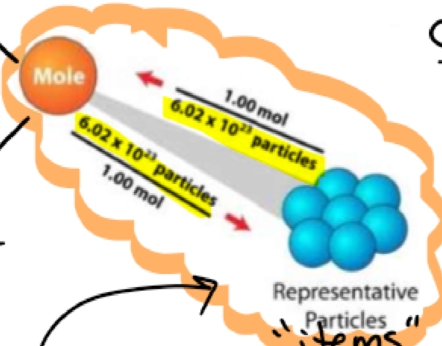
Chemists often refer to ***"a mole of a substance"***, rather than a mole of a substances particles.... this is because the mole is a way to express the amount of material

6.02×10^{23} flexible units

can be atoms/particles/molecules } whatever the Question refers to.

mass

Conversions Using Avogadro's Number



Volume

"items"

1 mol = 6.02×10^{23} molecules

EXAMPLE: How many molecules are there in 0.125 mol of molecules?

$$\frac{0.125\ \cancel{mol}}{} \Big| \frac{6.02 \times 10^{23}\ \text{molecules}}{1\ \cancel{mol}} = 7.525 \times 10^{22} = 7.53 \times 10^{22}\ \text{molecules}$$

3sf

EXAMPLE: How many moles of N atoms are there in 5.00 x 10^{17} N atoms?

#atoms → mol

$$\frac{5.00 \times 10^{17}\ N\ \cancel{atoms}}{} \Big| \frac{1\ mol}{6.02 \times 10^{23}\ N\ \cancel{atoms}} = \boxed{8.31 \times 10^{-7}\ mol}$$

EXAMPLE: A light source emits 8.50 x 10^{17} photons per second. (A photon is a "particle of light".) How many moles of photons are emitted by the light source in one minute?

"items" → $\frac{photons}{s}$ → $\frac{mol}{min}$

1 mol = 6.02×10^{23} photons

$$\frac{8.50 \times 10^{17}\ \cancel{photons}}{1\ \cancel{s}} \Big| \frac{1\ mol}{6.02 \times 10^{23}\ \cancel{photons}} \Big| \frac{60\ \cancel{s}}{1\ min} = \boxed{8.47 \times 10^{-5}\ \frac{mol}{min}}$$

3sf 3sf

Note: The above three examples don't depend on whether the particles are atoms, molecules, or whatever. The calculations just deal with the numbers of particles.

Avogadro's Number is also used to find the molar mass of particles if the mass of one particle is known.

"molar mass"

EXAMPLE: A particular variety of carbon atom has a mass of 2.16 x 10^{-23} g/atom. What is the mass of a mole of this variety of carbon atom?

$\frac{g}{mol}$

$\frac{g}{atom} \rightarrow \frac{g}{mol}$ Keep mass!

$$\frac{2.16 \times 10^{-23}\ g}{1\ \cancel{atom}} \Big| \frac{6.02 \times 10^{23}\ \cancel{atoms}}{1\ mol} = 13.0\ \frac{g}{mol}$$

$\left(1.30 \times 10^{1}\ \frac{g}{mol}\right)$

chemistry homework

Assignment #4- Practice Problems #1-6 & Hebden Questions #16-17 page 84

Complete ALL assignments on a separate piece of paper and attach to your booklet when handing in at the end of the unit.

PRACTICE —Converting Moles to Number of Items

1. Chromium ions are responsible for the beautiful colours of rubies and emeralds. How many chromium ions (Cr^{3+}) are in 3.5 mol of chromium ions?

2. 30.0 mol H_2O = ____________ molecules H_2O

3. How many atoms of sodium are in 0.023 mol Na?

1. 2.1×10^{24} ions Cr^{3+}
2. 1.81×10^{25} molecules H_2O
3. 1.4×10^{22} atoms Na

PRACTICE — Converting Number of Items to Moles

4. Incandescent lights are filled with argon to prevent the glowing filament from burning up. How many moles of argon do 1.81×10^{22} atoms of argon represent?

~~1~~4. 0.0301 mol Ar
~~2~~5. 3.74 mol CO_2
~~3~~6. 0.154 mol NaCl

5. 2.25×10^{24} molecules CO_2 = ___________ mol CO_2?

6. A 1-L intravenous bag of saline solution contains 9.27×10^{22} formula units of NaCl. How many moles of NaCl is this?

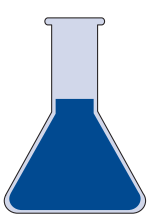
COMBINED EXERCISES:

15. Calculate the number of moles contained in the following.
    (a) 10.6 L of $SO_2(g)$ at STP
    (b) 7.50×10^{21} molecules of HNO_3
    (c) 425 mg of $Ca(OH)_2$
    (d) 4.25×10^{12} molecules of Fe_2O_3
    (e) 0.950 kg of NaOH
    (f) 25.0 mL of $N_2(g)$ at STP
    (g) 5.50×10^{25} molecules of CCl_4
    (h) 0.120 L of $NO_2(g)$ at STP

16. Calculate the volume of the following gases at STP.
    (a) 0.235 mol of $B_2H_6(g)$
    (b) 9.36 mol of $SiH_4(g)$
    (c) 2.55×10^{3} mol of $C_2H_6(g)$

17. Calculate the mass of each of the following.
    (a) 0.125 mol of $CO_2(g)$ at STP
    (b) 5.48 mol of $FeCl_3(s)$
    (c) 6.54×10^{-4} mol of HCN(g) at STP
    (d) 15.4 mol of $Ni(OH)_2(s)$

18. Calculate the mass of 1 mol of each of the following.
    (a) $Na_2B_4O_7 \cdot 10H_2O$
    (b) Grandma Smith, an average grandmother, having a mass of 52 kg. (Express your answer in kilograms.)
    (c) a bismuth atom with a mass of 3.52×10^{-22} g
    (d) an electron having a mass of 9.1×10^{-28} g.
    (e) $Cu_3(OH)_2(CO_3)_2$
    (f) a book having a mass of 1.34 kg



Complete Lab Activity 4B: Copper & Iron Nail

chemistry homework

Assignment #5- Hebden Questions #15 & 18 page 84
Complete ALL assignments on a separate piece of paper and attach to your booklet when handing in at the end of the unit.

(you can work on this while your lab is processing)

Vol

↑ molar vol 22.4L

mol

molar mass ↙ mass

↘ Avagadro's# particles.

MULTIPLE CONVERSIONS BETWEEN MOLES, MASS, VOLUME AND NUMBER OF PARTICLES

Before jumping into the middle of some complex conversion factor calculations, a simple process must be understood: how to find the number of atoms in a given number of molecules

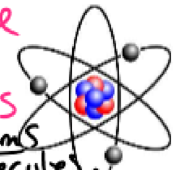
This calculation simply involves counting the number of atoms in one molecule and then multiplying by the number of molecules involved.

EXAMPLE: **How many atoms** are there in 5 molecules of $\boxed{CuSO_4{\cdot}5H_2O}$?

1 molecule → equal

$1+1+4+5(2+1) = 21$ atoms

$\dfrac{1\text{ molecule}}{21\text{ atoms}}$ or $\dfrac{21\text{ atoms}}{1\text{ molecules}}$.



$$\frac{5\ \cancel{\text{molecules}}}{} \left|\ \frac{21\text{ atoms}}{1\ \cancel{\text{molecules}}}\right. = 105\text{ atoms}$$

EXAMPLE: **How many HYDROGEN atoms** are there in 30 molecules of $\boxed{H_3PO_4}$?

1 molecule = 3 H atoms

$$\frac{30\ \cancel{\text{molecules}}}{} \left|\ \frac{3\text{ H atoms}}{1\ \cancel{\text{molecule}}}\right. = 90\text{ H atoms}$$

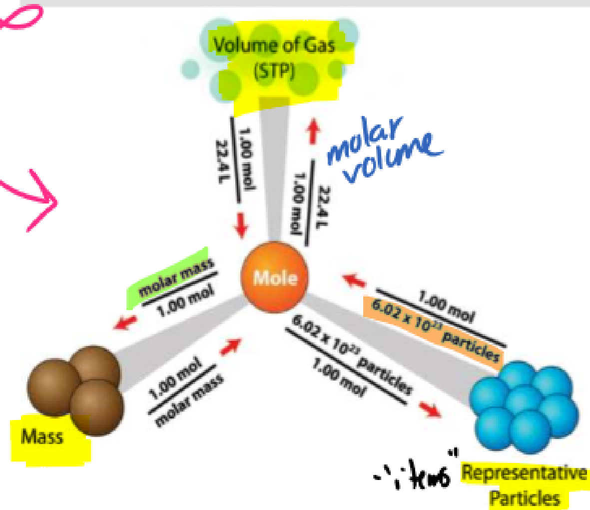
* "oxygen atom" = O ← 1 atom

"oxygen gas" = O_2 ← 2 atoms

Part D: Two-Step Mole Conversions

The mole serves as a link between the invisible world of atoms and observable quantities of chemicals. The mole is the central unit of chemistry and allows us to keep track of atoms and molecules.

A quantity can only only be related to another substance through THE MOLE!



For now, our conversions are limited to those between moles and particle, between moles and grams and moles and volume.

The beauty of the "***Mole Wheel***" that as you learn more chemical quantities they can simply be added to the outside of the spoke in relation to the mole.

Think of the mole as the hub of a wheel with the spokes leading out to all the other units. In our wheel model, the spokes represent the conversion factors

In order to relate or "connect" a new chemical quantity to all of the others you only need to connect it to the mole.

The mole is "central" to all conversions between mass, particles and volume: each calculation goes form **STARTING UNIT --> MOLES --> FINAL UNIT**

(given value)

EXAMPLE: What is the volume occupied by 50.0 g of $NH_3(g)$ at STP?

→ 1 mol = 22.4L

given mass → mole → volume (mass → volume)

$$\frac{50.0\ \cancel{g}}{} \left|\ \frac{1\ \cancel{mol}}{17.0\ \cancel{g}}\right| \frac{22.4\ L}{1\ \cancel{mol}} = 65.9\ L$$

molar mass of $NH_3(g)$ — molar volume

$\leftarrow$ molar mass of Cl = $35.5\,\frac{g}{mol}$

EXAMPLE: What is the mass of 1.00×10^{12} atoms of Cl?

given atoms $\rightarrow$ mol $\rightarrow$ mass

$$\frac{1.00\times10^{12}\ \cancel{\text{atoms Cl}}}{} \times \frac{1\ \cancel{\text{mol}}}{6.02\times10^{23}\ \cancel{\text{atoms}}} \times \frac{35.5\text{ g}}{1\ \cancel{\text{mol}}} = \underline{5.90\times10^{-11}}\text{ g Cl}$$

EXAMPLE: How many oxygen atoms are contained in 75.0 L of $SO_3(g)$ at STP?

$\boxed{\text{3 O atoms per molecule}}$.

↳ 1 mol = 22.4 L

given volume $\rightarrow$ MOL $\rightarrow$ molecules of SO_3 $\rightarrow$ atoms of O

$$\frac{75.0\ \cancel{\text{L}}}{} \times \frac{1\ \cancel{\text{mol}}}{22.4\ \cancel{\text{L}}} \times \frac{6.02\times10^{23}\ \cancel{\text{molec. } SO_3}}{1\ \cancel{\text{mol}}} \times \frac{3\text{ O atoms}}{1\ \cancel{\text{molecule of } SO_3}} = \underline{6.05\times10^{24}}\text{ O atoms}$$

chemistry homework

Assignment #6- Practice Problems #1-3 & Hebden Questions #21-24 *(odd letters)* page 85-87

Complete ALL assignments on a separate piece of paper and attach to your booklet when handing in at the end of the unit.

PRACTICE — Two-Step Conversions

1. Fill in the missing entries to determine the mass in grams of a billion billion (1×10^{18}) sulphur dioxide molecules.

$$(1 \times 10^{18}\text{ molecules } SO_2) \times \frac{1\text{ mol } SO_2}{\_\_\_\_\_\_\_\_\text{ molecules } SO_2} \times \frac{\_\_\_\_\_\_\text{ g } SO_2}{1\text{ mol } SO_2} = \_\_\_\_\_\_\_\_\text{ g } SO_2$$

2. How many atoms are in 2.1 g Br?

ANSWERS:

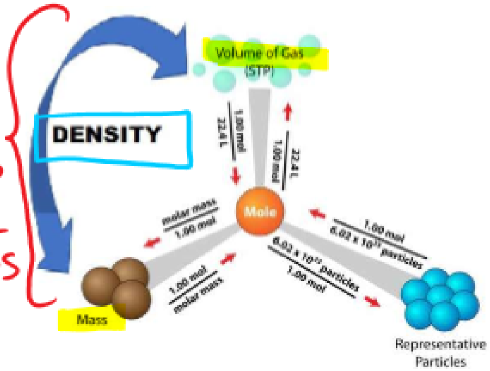
1. 1×10^{-4} g SO_2
2. 1.6×10^{22} atoms Br
3. 1.79×10^{-22} g Ag

3. What is the mass in grams of one atom of Ag?

Part E: Molar Volume and Density

Density is the amount of matter in a given volume of an object or material. In other words, it is the mass per unit volume. Density is a conversion factor that relates a substance's mass directly to its volume without any reference to the mole. In terms of our wheel model, density is the section of the rim that connects MASS and VOLUME.

For NOT @STP or NOT gases



So far, the volumes used all refer to a gaseous substance at STP. If Density is mentioned at any point in a question, you should immediately recal that:

$$\text{Density} = \frac{\text{mass}}{\text{volume}} = \frac{(g)}{(L)}$$

1000 mL = 1L

NOT @STP

Sample Problem — Converting Volume Directly to Mass The density of methanol, CH_3OH, at 20°C is 0.813 g/mL. What is the mass of 0.500 L of the alcohol at 20°C? Density needs to be in mL	
What to Think about	**How to Do It** volume $\xrightarrow{\text{(density)}}$ mass
1. Convert: L $CH_3OH \rightarrow$ g CH_3OH 2. Setup: 0.500 L $CH_3OH \times \frac{?\ g\ CH_3OH}{1\ L\ CH_3OH}$ 3. Conversion factor: 813 g/L	$\frac{0.500\,L \mid 1000\,mL}{\mid 1\,L} = 500\,mL$ (starting volume) $\frac{500\,\cancel{mL} \mid 0.813\,g}{\mid 1\,\cancel{mL}} = \boxed{407\ g\ of\ CH_3OH}$ 406.5 3sf use density as conversion factor

Molar volume and density are related through molar mass $\quad \frac{\cancel{g}}{mol} \times \frac{L}{\cancel{g}} = \frac{L}{mol}$

mass / D | Vol.

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

we can calculate for any substance

Sample Problem — Calculating Molar Volume from Density In an episode of the television show "MythBusters," the team floated an aluminum foil boat on the invisible gas, sulphur hexafluoride, SF_6. SF_6 has a density of 6.00 g/L at room temperature and pressure, about six times that of air. not @STP What is the molar volume of SF_6 under these conditions?	
What to Think about	**How to Do It**
1. molar volume = $\frac{\text{molar mass}}{\text{density}}$ 2. ~~Setup: $\frac{?\ g\ SF_6}{1\ mol\ SF_6} \times \frac{1\ L\ SF_6}{6.00\ g\ SF_6}$~~ 3. Conversion factors: 146.1 g SF_6 per 1 mol SF_6	calculate molar mass $SF_6 = S(32.1g) + 6 \cdot F(19.0g) = 146.1g$ Molar Volume $= \frac{m.m}{D} = \frac{146.1\,\cancel{g} \mid 1\,L}{1\,mol \mid 6.00\,\cancel{g}} = 24.35\ \frac{L}{mol}$ 3 s.f. $= \boxed{24.4\ \frac{L}{mol}}$ m.m / V | D

Tips for approaching Density problems involving The Mole:

TYPE A
- **If the volume of a solid or liquid is the unknown,** calculate the volume from $V = m/d$. If the mass is not known, find the mass from the moles of the substance present. If the **molar volume** is the unknown, the molar mass is used in the calculation. **(Note that you cannot use the molar volume of a gas, 22.4 L, when calculating the volume of a liquid or solid.)**

TYPE B
- **If the density is unknown**, you will need both mass and volume to calculate: $d = m/V$. The mass can be found if the number of moles is known. If neither the mass nor volume is given, the density of a gas at STP can be found by using the mass of 1 mol and the volume of 1 mol at STP.

TYPE C
- **If the number of moles is unknown,** use the density and volume to calculate $m = d \cdot V$ and then convert the mass to moles.

TYPE D
- **If the molar mass of a gas at STP is unknown,** the data given is usually the mass and volume of a small amount of gas. In this case, find the density of the gas using the given mass and volume and then combine the density with the volume of 1 mol (22.4 L) to find the mass of 1 mol.

can solve for volume in mL (density given in g/mL)

liquid (means can't use 1 mol = 22.4 L)

TYPE A

EXAMPLE: **What is the volume occupied by 3.00 mol of ethanol, $CH_3CH_2OH(l)$? (d = 0.790 g/mL)**

moles → mass → volume

M.V. = $\frac{M.M}{D}$

$$\frac{3.00\ \text{mol}}{} \left| \frac{46.0\ g}{1\ \text{mol}} \right| \frac{1\ mL}{0.790\ g} = \boxed{175\ mL\ (1.75\times10^2\ mL)}$$

molar mass of CH_3CH_2OH = 46.0 g

molar mass

density

TYPE B

EXAMPLE: **What is the density of $O_2(g)$ at STP?**

1 mol = 22.4 L

$O_2 = 2(16.0\ g) = 32.0\ g$

$$D = \frac{m}{V} = \frac{\text{molar mass}}{\text{molar volume}} = \frac{32.0\ g/\text{mol}}{22.4\ L/\text{mol}} = \boxed{1.43\ g/L}$$

OR

$$\frac{32.0\ g}{1\ \text{mol}} \left| \frac{1\ \text{mol}}{22.4\ L} \right.$$

TYPE C

EXAMPLE: **How many moles of Hg(l) are contained in 100 mL of Hg(l)? (d = 13.6 g/mL)**

liquid (not a gas @ STP)

volume → mass → moles

$$\frac{100\ mL}{} \left| \frac{13.6\ g}{1\ mL} \right| \frac{1\ \text{mol}}{200.6\ g} = 6.78\ \text{mol} = \boxed{7\ \text{mol}}$$

1 s.f.   1 sf.

TYPE D

EXAMPLE: **A 2.50 L bulb contains 4.91 g of a gas at STP. What is the molar mass of the gas?**

1 mol = 22.4 L
molar vol.

$$\text{density} = \frac{m}{V} = \frac{4.91\ g}{2.50\ L}$$

$$D = \frac{m.m}{m.v} \quad \therefore mm = (D)(m.v)$$

calc.

$$D = 1.964\ \frac{g}{L}$$

3 sf.

$$\text{molar mass} = \left(1.964\ \frac{g}{L}\right)\left(22.4\ \frac{L}{\text{mol}}\right) = 43.9936\ \frac{g}{\text{mol}} = 44.0\ \frac{g}{\text{mol}}$$

TYPE C+

EXAMPLE: **$Al_2O_3(s)$ has a density of 3.97 g/mL. How many atoms of Al are in 100 mL of Al_2O_3?**

(no conversion needed)

calc. molar mass.

$2 \cdot (Al) = 2(27.0\ g)$
$3 \cdot (O) = 3(16.0\ g)$
102.0 g

1 mol = 6.02×10^{23} molecules.
2 Al atoms = 1 molecule Al_2O_3 *

volume $\xrightarrow{\text{density}}$ mass → mol → molecules → atoms

$$\frac{100\ mL}{} \left| \frac{3.97\ g}{1\ mL} \right| \frac{1\ \text{mol}}{102.0\ g} \left| \frac{6.02\times10^{23}\ \text{molec}\ Al_2O_3}{1\ \text{mol}} \right| \frac{2\ Al\ \text{atoms}}{1\ \text{molec.}\ Al_2O_3}$$

density   molar mass   Avagadro's #

$= 4.686157\ldots\times10^{24}$

1 sf.

$\boxed{5\times10^{24}\ Al\ \text{atoms}}$

chemistry homework

Assignment #7- Practice Problems #1-3 & Hebden Questions #25-34 *(all)* page 88 + The Mole Review
Complete ALL assignments on a separate piece of paper and attach to your booklet when handing in at the end of the unit.

$$\boxed{\text{Density} = \frac{\text{molar mass}}{\text{molar volume}}}$$

Practice Problems — Calculating Molar Volume and Density

1. Gold has a density of 19.42 g/cm³. The standard gold bar held as gold reserves by central banks weighs 12.4 kg. What is the volume of the standard gold bar?

19.42 g/mL

$1\,cm^3 = 1\,mL$

$$\frac{12.4\,kg}{} \Big| \frac{10^3\,g}{1\,kg} = 12.4 \times 10^3\,g$$

$$D = \frac{m.m}{m.V} \quad \therefore\ mV = \frac{mm}{D} = \frac{12.4 \times 10^3\,\cancel{g}}{19.42\,\frac{\cancel{g}}{mL}}$$

$= 639\,mL$

2. Mercury has a density of 13.534 g/mL at room temperature. What is the mass of 12.7 mL of mercury?

same… no conversion

$$D = \frac{m.m}{mV} \quad \therefore\ mm = \text{Density} \cdot \text{molar volume}$$

$\frac{L}{mol}$ or $\frac{mol}{L}$

$$\frac{13.534\,g}{1\,mL} \Big| \frac{12.7\,mL}{1\,mol} = 171.88$$

$\boxed{= 172\,g}$

3s.f.

3. Although ethanol (C_2H_5OH) is best known as the type of alcohol found in alcoholic beverages, its largest use is as a fuel or fuel additive. The density of ethanol is 0.789 g/mL. What is the molar volume of ethanol?

calc.

molar mass $= 2(12.0) + 6(1.0) + (16.0g) = 46.0g$

$$D = \frac{\text{molar mass}}{\text{molar volume}}$$

$$\therefore\ m.V = \frac{m.m}{D}$$

$$\text{molar Volume} = \frac{46.0g}{0.789\frac{g}{mL}} = \boxed{58.3\,mL}$$

Th E Mo L E Re V I E W

KEY

1. a) 6.57L H_2S b) $2.69x10^{-3}$L c) 0.019L BrF d) $3.2x10^{3}$ L B_2H_6

2. a) $3.271x10^{-22}$g Au b) $3.6x10^{-7}$g AgCl c) 0.469g C_3H_6 d) 13.0g SF_6

3. a) 0.0391mol $C_{10}H_8$ b) $2.47x10^{-3}$ mol K_3PO_4 c) 0.268mol NO_3F d) $4.46x10^{-5}$mol O_3
   e) $7.56x10^{-12}$mol Pt f) $1.000x10^{-7}$mol PCl_5

4. a) $7.53x10^{6}$g/mol b) 413g/mol c) 178g/mol d) 248.2g/mol e) 93.0g/mol f) 329.6g/mol

5. a) $1.52x10^{-3}$g/mL b) 0.01020L/mol c) 0.0207mol CS_2 d) 0.704g/mL e) 0.899mL Ag
   f) 2.28g/mL g) 129mol C_2H_5OH h) 34.0g/mol i) 0.418mL NaCl j) 62.2g/mol
   k) 0.013L/mol

6. a) 18 atoms b) $5.39x10^{7}$L COF_2 c) $4.38x10^{23}$ molecules d) $1.12x10^{-3}$mol HCN
   e) 10.5L ClF_3 f) 0.457mol Fe g) $3.36x10^{21}$molecules NOCl h) 9.755g Pt i) 136.5g/mol
   j) $2.32x10^{-3}$g/mL k) 0.0935g Kr l) $8.573x10^{-3}$L/mol m) 63.9g/mol n) 1.05g/mL
   o) $6.99x10^{-4}$mol CuSCN p) 3.73mL q) $5.49x10^{-4}$g/mol r) 51.9g/mol s) 1.74mol HgS