

MOLES, % COMPOSITION & EMPIRICAL FORMULAS DIARY

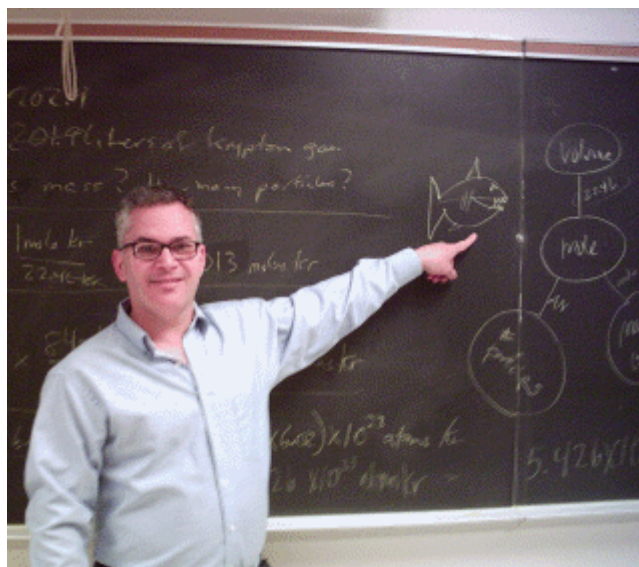
MOLES

Any time that you get to draw a shark in chemistry class is a good day. Fishing is not one of my favorite "sports", but I do enjoy fishing at Wegmans for shark (and for salmon). This is your teacher in the fall of 2006, it's the first picture of me on the site!

Near that fish is the "mole island" conceptual map of how moles are related to mass, volume, and the number of particles.

The Mole is central to your understanding much more of chemistry, you must work hard to truly grasp their significance.

Avogadro's number sets the basic ratio between the mole and how many particles it is. Like a dozen is twelve, a mole is exactly 6.02×10^{23} particles.



Half a dozen is six, and half a mole is therefore half of Avogadro's Number: or 3.01×10^{23} particles. This relationship of moles to an exact number of particles allows us to mathematically connect masses of substances to the number of particles present.

Particles can be atoms if the substance comes in atoms, say noble gases, or metals. Sometimes particles can be formula units - if the substance is an ionic compound. Particles can also be molecules - when if the substance is made from two non-metals joined into a molecular compound. Particles can even be ions, if you want to count how many ions are present in a substance. Having a mole of anything of "real" size is a problem. A mole of atoms is a huge number but relatively light mass since atoms are so very small. Having a mole of something the size of a banana would be larger than the moon.

Besides the mole to number of particles ratio, there is a special mass relationship between atoms on the periodic table and the concept of moles. If you look at your Periodic Table, and see that one atom of Helium has an atomic mass of 4.00260 amu (which we usually round to 4 amu), the mass of ONE MOLE OF HELIUM is in fact 4.00260 grams, which is also usually rounded to just 4 grams.

The units change between single atoms and moles, but the periodic table provides the numeric. We can use these numbers to determine how many grams one mole of any element is, and to determine the MOLAR MASS of any compound (by just adding up individual atomic mole masses by the ratios of atoms in the compound - see below).

Examples include

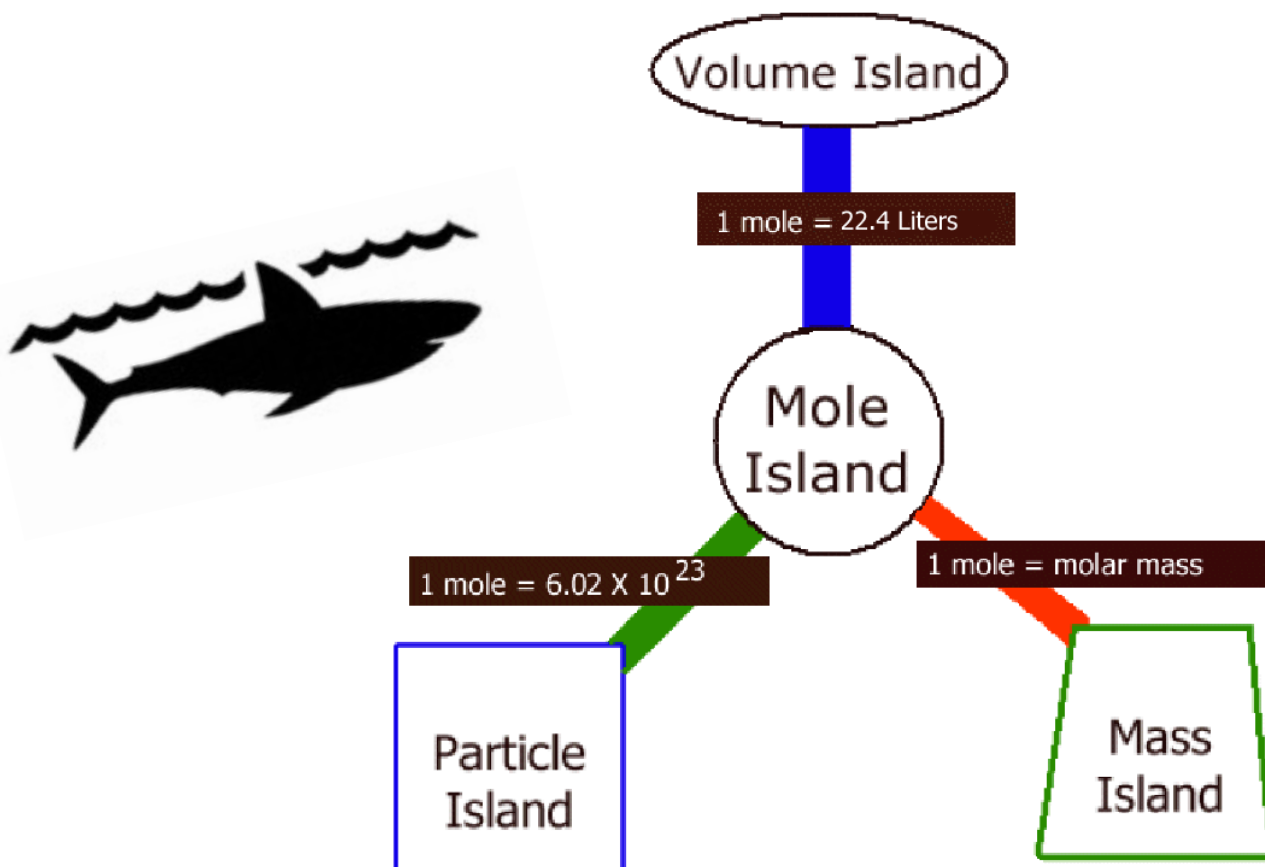
atom	atomic mass	molar mass
niobium	93 amu	93 grams/mole
zinc	65 amu	65 grams/mole
sulfur	32 amu	32 grams/mole
silicon	28 amu	28 grams/mole
NaCl	$23 + 35 = 58$ amu	58 grams/mole
NaOH	$23 + 16 + 1 = 40$ amu	40 grams/mole
$C_6H_{12}O_6$	$72 + 12 + 96 = 180$ amu	180 grams/mole

With gases, the mole to volume relationship is the "simplest" to connect. At standard temperature and pressure (zero centigrade and one atmosphere pressure), one mole of any gas is equal to 22.4 Liters of volume. In our class the gas parameters of pressure and temperature will be at STP until we study gases later in the year. So, the number to remember is 22.4 Liters.

Examples include

gas	formula	volume of one mole at STP
one mole of helium	He	22.4 liters
one mole of carbon dioxide	CO ₂	22.4 liters
two moles of krypton	Kr	44.8 liters
one half mole neon	Ne	11.2 liters
1.0 mole nitrogen dioxide	NO ₂	22.4 liters
3.0 moles iodine gas	I ₂	67.2 liters

Mole Islands... The drawing below describes the connection between MOLES in the middle, with the "islands" that surround it. The only way to make your way from any island to another is to take the ONLY BRIDGE available, and PAY THE TOLL as indicated. Use the **BLACK** tolls to make your conversions factors. If you "cheat" and try to skip the mole conversion, the sharks will eat you. Stay on the "BRIDGES"!



Going from one part of the diagram to another (going from one island to another) you may only use the bridges as shown. Each bridge has its own toll to pay, indicated in **BLACK**. Use these "tolls" to set up one step dimensional analysis conversions to change from one unit to another (for example, from LITERS TO MOLES, or MOLES TO GRAMS, or MOLES TO NUMBER OF PARTICLES).

Example: if you know that you have 50.0 grams of a substance, you need to get its MOLAR MASS to convert that to MOLES. Once you determine the number of moles it is, then you can find the volume (if it's a gas) or the number of particles.

Problems that start at a different unit, say volume or number of particles, **ALWAYS convert to moles FIRST.**

Once you convert to moles, the other conversions are just one step away. Moles are central to chemistry and this diagram will help you keep it all straight. It will require practice, and if you don't practice, it will be very apparent.

MOLAR MASS

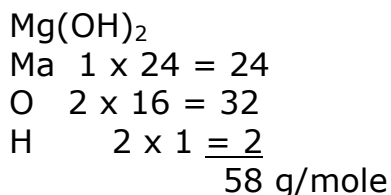
By definition, how many grams exactly one mole of a substance weighs. If it is just an element, read the atomic mass on the periodic table, and change the "AMU" units to "GRAMS" instead. If it is a compound, write the PROPER FORMULA of the compound, and multiply the number of atoms by the proper atomic masses, and then add them all up. Units will be GRAMS PER MOLE.

For example, determine the MOLAR MASS of sodium hydroxide. NaOH, has 3 atoms, one each of sodium, oxygen, and hydrogen.	Molar Mass of <u>NaOH</u> Na - sodium $1 \times 23 = 23$ O - oxygen $1 \times 16 = 16$ H - hydrogen $1 \times 1 = \underline{1}$ total = 40 grams/mole
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% Composition by Mass

Is a measurement of the proportion of mass one part of a compound makes up in the compound. For example, what is the proportion of magnesium in magnesium hydroxide by mass?

To figure this out, you need first to set up the molar mass of the compound, then use the formula for % composition by mass and calculate.



$$\begin{array}{l} \text{\% comp by mass of Mg in} \\ \text{magnesium hydroxide} \\ \\ \text{Mg} \quad \frac{24 \text{ g}}{58 \text{ g}} \times 100\% = 41.4\% \end{array}$$

% COMPOSITION BY MASS

All this means is if you have a particular compound, you can use the periodic table to determine its molar mass. Once you get that, you can determine what percentage of the compound is of any of the parts, as measured in grams.

A real world example is this: if you take a piece of chewing gum you can know its total mass just by weighing it on a balance. The gum is made up of an indigestible chewy part, and sugar (little else really). If you chew the gum long enough all the

$$\% \text{ Comp of Sugar in Gum} = \frac{\text{Mass of missing sugar}}{\text{Mass of gum originally}} \times 100\%$$

sugar dissolves onto your tongue. When the gum is lousy tasting, the sugar is gone. If you weigh this tasteless piece of gum, you find it weighs much less than before.

The missing mass is the missing sugar (it's not really missing, you just ate it!). You could determine the % Composition by Mass of the sugar in the gum by this formula:

You could use this formula to determine what percent by mass of your sodium chloride is a highly reactive metal and what percent by mass is a toxic gas! This too takes practice.

SEE CHART ON THE NEXT PAGE

An example...

Find the % composition of chlorine in hydrogen chloride (HCl).

HCl

% Composition by Mass

H - Hydrogen $1 \times 1 = 1$

Cl - Chlorine $1 \times 35 = 35$

MOLAR MASS = 36 g/mole

$\frac{\text{mass of the part}}{\text{mass of the whole}} \times 100\%$

$\frac{35\text{g}}{36\text{g}} \times 100\% = 97.2\% \text{ Cl by mass}$

If you have 50.0 grams of HCl, how many grams would be chlorine?

50.0 grams HCl $\times 0.972 = 48.6$ grams [97.2% = 0.972 AS A DECIMAL]

If you have 312 grams of HCl, how many grams would be chlorine?

312 grams HCl $\times 0.972 = 303$ grams

The total mass that is chlorine is always 97.2% FOR HCl, based upon the % Composition by mass we figured above.

EMPIRICAL FORMULAS

An empirical formula is a math concept more than a chemistry one. It really is the lowest ratio of atoms or ions that make up a formula. You are familiar with glucose, $C_6H_{12}O_6$, and the ratio of atoms in that is of course 6:12:6, which can be reduced to 1:2:1. The EMPIRICAL FORMULA for glucose is just CH_2O .

The ratio has NOTHING to do with the actual chemistry, density, molar mass, etc. It is a way to categorize groups of compounds, and to make you think.

The EMPIRICAL FORMULAS for these compounds are:

$C_5H_{10}O_5$	CH_2O
C_2H_2	CH
C_4H_{10}	C_2H_5
C_8H_{18}	C_4H_9
$MgSO_4$	$MgSO_4$ (this formula cannot be reduced to a lower ratio)
H_2O	H_2O (this formula cannot be reduced to a lower ratio)
CH_4	CH_4 (this formula cannot be reduced to a lower ratio)
$C_{44}H_{88}O_{44}$	CH_2O

Empirical formulas are about the LOWEST RATIO. Very often the "lowest ratio formula", such as CH_2O , is not even a real compound, it cannot even exist chemically. But the ratio can exist in your mind, or on paper.

An EMPIRICAL FORMULA is more an IDEA than a real thing. Sometimes Empirical Formulas are the same as the formula of the real compounds, like with magnesium sulfate, water, or methane gas. The last example shows that no matter how big the numbers, the lowest ratio makes the empirical formula.

Types of mole problems...

There are only a limited number of kinds of mole problems. Using your mole island construct, you can easily do all of them.

Problems for practice. Answers on next page.

1. How many grams are in 1.0 moles of NaHCO_3 , which is baking soda?
2. How many moles are in 25.0 grams of baking soda?
3. How many moles is 145.6 liters of helium gas at STP?
4. If you have 2.75 moles of CO_2 gas, how many liters does it take up at STP?
5. If you have 2.75 moles of CO_2 gas, how many particles is that?
6. If you have 3.50×10^{27} atoms of neon gas, how many moles is that?
7. If you have 75.0 gms $\text{Cl}_{2(g)}$, how many formula units & how many liters does it take up at STP?

1	The molar mass of sodium hydrogen carbonate is 84 g/mol.			
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2	$\frac{25.0 \text{ g baking soda}}{1}$	X	$\frac{1 \text{ mole baking soda}}{84 \text{ g baking soda}}$	= 0.298 moles baking soda
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3	$\frac{145.6 \text{ liters He}}{1}$	X	$\frac{1 \text{ mole He}}{22.4 \text{ liters He}}$	= 6.50 moles He
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4	$\frac{2.75 \text{ moles CO}_2}{1}$	X	$\frac{22.4 \text{ liters CO}_2}{1 \text{ mole CO}_2}$	= 61.6 liters CO_2
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5	$\frac{2.75 \text{ moles CO}_2}{1}$	X	$\frac{6.02 \times 10^{23} \text{ molecules CO}_2}{1 \text{ mole CO}_2}$	= 16.555 x 10 ²³ changes to 1.66 x 10 ²⁴ molecules CO_2
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6	$\frac{3.50 \times 10^{27} \text{ atoms Ne}}{1}$	X	$\frac{1 \text{ mole Ne}}{6.02 \times 10^{23} \text{ atoms Ne}}$	= $\frac{3.50}{6.02} \times \frac{10^{27}}{10^{23}}$ = 0.581 x 10 ⁴ changes to 5.81 x 10 ³ moles Neon
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7	$\frac{75.0 \text{ grams Cl}_2}{1}$	X	$\frac{1 \text{ mole Cl}_2}{70 \text{ grams Cl}_2}$	= 1.07 moles Cl_2	go to the next line
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$\frac{1.07 \text{ moles Cl}_2}{1}$	X	$\frac{22.4 \text{ liters Cl}_2}{1 \text{ mole Cl}_2}$	= 24.0 liters Cl_2
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