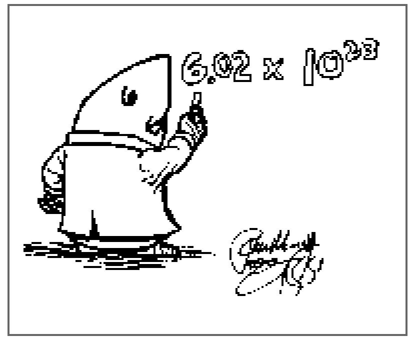
Chapter 7: MOLES, % COMPOSITION & EMPIRICAL FORMULAS DIARY

Moles



Molecules and atoms are extremely small objects - both in size and mass. Consequently, working with them in the laboratory requires a large collection of them. How large does this collection need to be? A standard needs to be introduced. This standard is the "mole". The mole is based upon the carbon-12 isotope. We ask the following question: How many carbon-12 atoms are needed to have a mass of exactly 12 g. That number is NA - Avogadro's number. Thus, NA is defined by
NA x (mass of carbon-12 atom) = 12 g

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. Sometimes particles can be formula units, like if the substance is an ionic compound. Particles can also be molecules as in when 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. Looking at the Periodic Table, we can 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 (aka gram atomic 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

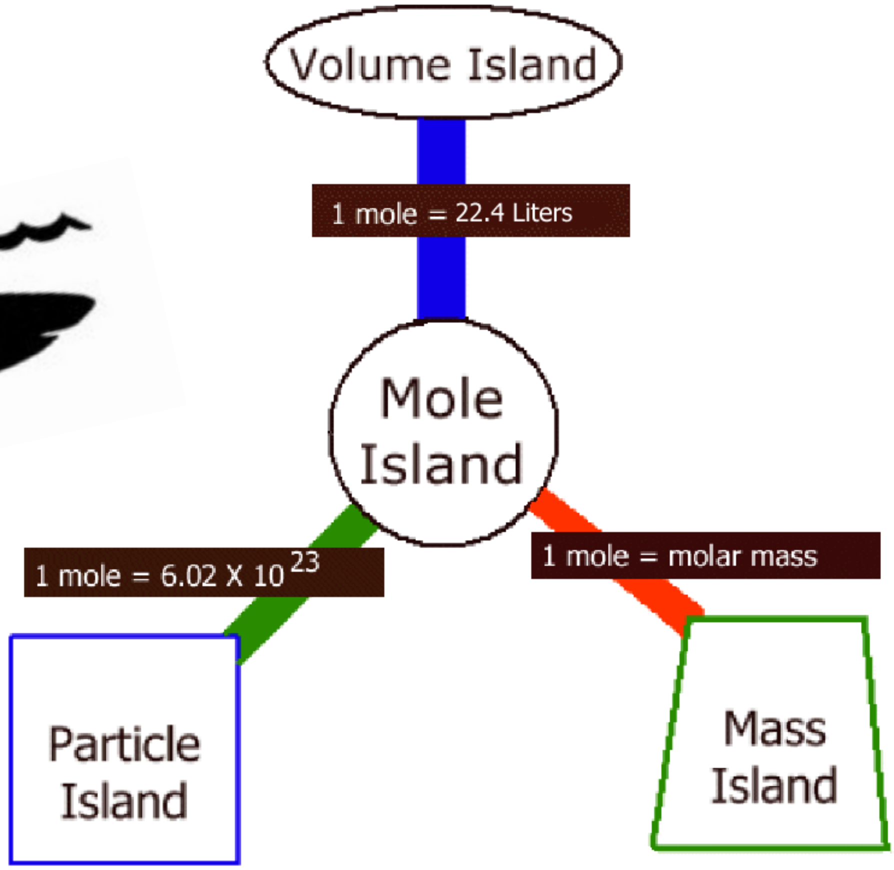
With gases, the mole to volume relationship is the "simplest" to connect. At standard temperature and pressure (zero centigrade and one atmosphere pressure), one mole of any gas is equal to 22.4 Liters of volume. In our class the gas parameters of pressure and temperature will be at STP until we study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

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 $Cl_{2(G)}$, 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.

2	$\frac{\text{25.0 g baking soda}}{1}$	X	$\frac{\text{1 mole baking soda}}{\text{84 g baking soda}}$	= 0.298 moles baking soda

3	$\frac{\text{145.6 liters He}}{1}$	X	$\frac{\text{1 mole He}}{\text{22.4 liters He}}$	= 6.50 moles He

4	$\frac{\text{2.75 moles } CO_2}{1}$	X	$\frac{\text{22.4 liters } CO_2}{\text{1 mole } CO_2}$	= 61.6 liters CO_2

5	$\frac{\text{2.75 moles } CO_2}{1}$	X	$\frac{\text{6.02 x 1023 molecules } CO_2}{\text{1 mole } CO_2}$	= 16.555 x 1023 changes to 1.66 x 10 24 molecules CO_2

6	$\frac{3.50\text{x}10^{27}\text{ atoms Ne}}{1}$	X	$\frac{\text{1 mole Ne}}{6.02 \text{ x } 10^{23}\text{ atoms Ne}}$	$= \frac{3.50}{6.02} \text{ x } \frac{10^{27}}{10^{23}}$	= 0.581 x 10^4 changes to 5.81 x 10^3 moles Neon

7	$\frac{\text{75.0 grams } Cl_2}{1}$	X	$\frac{\text{1 mole } Cl_2}{\text{70 grams } Cl_2}$	= 1.07 moles Cl_2	go to the next line

$\frac{\text{1.07 moles } Cl_2}{1}$	X	$\frac{\text{22.4 liters } Cl_2}{\text{1 mole } Cl_2}$	= 24.0 liters Cl_2

Molar mass

By definition, the molar mass or "gram atomic mass" is 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 NaOH** Na - sodium 1 x 23 = 23 O - oxygen 1 x 16 = 16 H - hydrogen 1 x 1 = 1 total = 40 grams/mole

% Composition by Mass

% 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.

$Mg(OH)_2$
Ma 1 x 24 = 24
O 2 x 16 = 32
H 2 x 1 = 2
58 g/mole

% comp by mass of Mg in magnesium hydroxide

Mg $\dfrac{24\text{ g}}{58\text{ g}} \times 100\% = 41.4\%$

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

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

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

HCl

H - Hydrogen 1 x 1 = 1
Cl - Chlorine 1 x 35 = 35

MOLAR MASS = 36 g/mole

% Composition by Mass

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

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

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

50.0 grams HCl X 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 X 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 EMPRICIAL 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.