Chemistry Moles Packet

INTRODUCTION TO MOLES

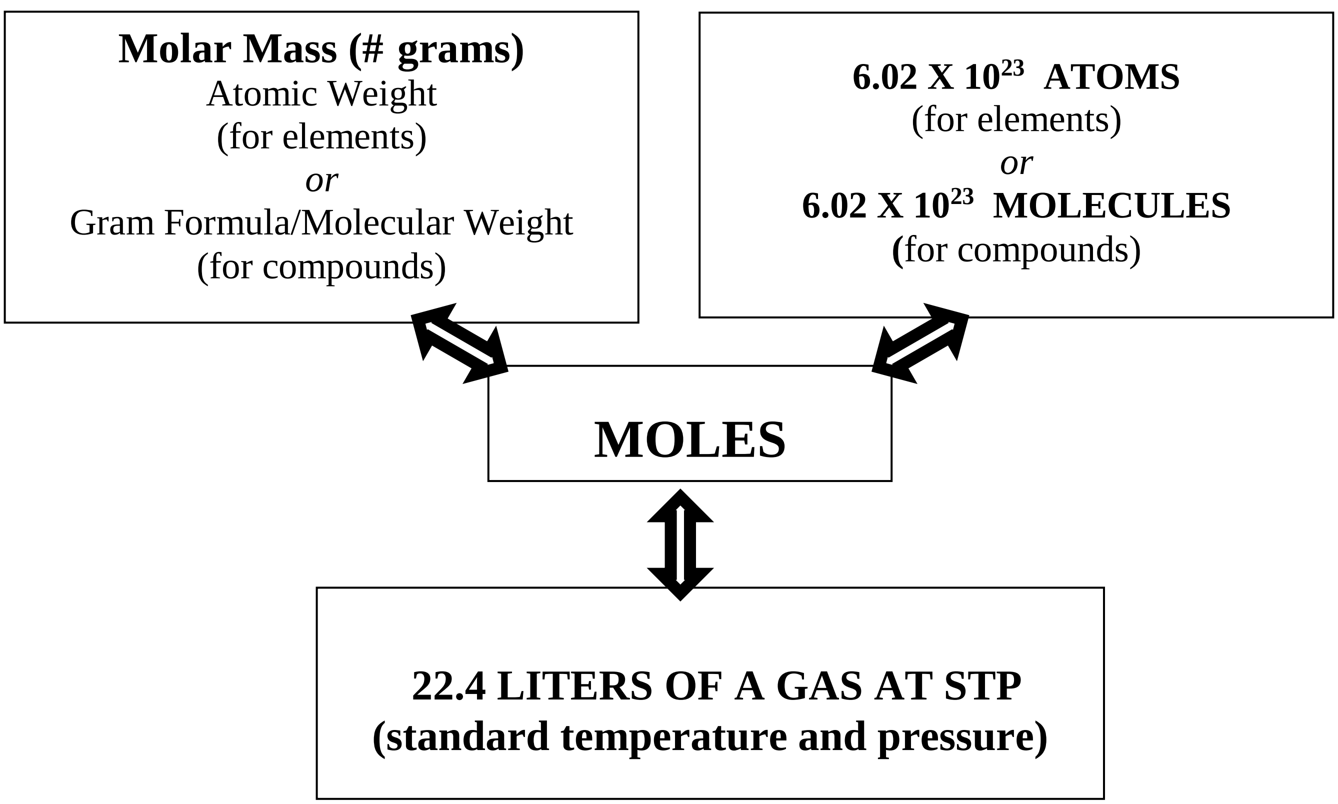
We are about to start on a unit of chemical calculations called “stoichiometry”. **Stoichiometry is how we calculate the relationships between the amounts of reactants and the amounts of products.** For example, if we know the amount of reactants we have, we can use stoichiometry to calculate how many products the chemical reaction will produce.

This packet provides an organized, step-by-step approach for these problems. If you follow this system and complete work each day as it is assigned, this will not be a difficult unit. That is good, because stoichiometry is one of the most central and important concepts in chemistry. **It is essential that you understand this unit in order to move forward in chemistry. In short: do not fall behind or you will be lost.**

These problems involve numbers but no difficult mathematics. All you will ever have to do is add, multiply or divide. **You will be expected to have a functioning calculator with you for every chemistry class.** As we solve these problems we will apply the factor-label (dimensional analysis) method you mastered early in the class, and we will frequently use scientific notation. Remember chapter 3 you followed the procedure:

what you are looking for	=	what you are given	X	a fraction or series of fractions where the numerator is equivalent to the denominator	=	answer

The only new concept we will introduce in this unit is the idea of a mole. **A mole is a quantity of matter that we use for conversion purposes.** We can convert from grams to moles, liters to moles (for gases), and atoms or molecules to moles. If you can convert any of these things to moles (and therefore moles to any of these things) we can convert grams to liters or molecules, liters to grams of molecules, and molecules to liters or grams.



CHEMISTRY WORKSHEET # 1 MOLAR MASS (GRAM MOLECULAR/FORMULA WEIGHTS)

We know that grams are actually a measure of the mass of matter and not the weight. Mass is the quantity of matter present; weight is a measure of the pull of gravity on matter and is measured in pounds or newtons. However, it is common usage in chemistry to talk about the "gram formula weight" rather than the technically correct term "gram formula mass" or "molar mass".

We have learned that the smallest particle of an element is an atom and the periodic table tells us the atomic masses or atomic "weights" for each element. We have also learned that the smallest unit of a compound are either molecules (for covalent compounds) or a collection of positive and negative ions (for ionic compounds).

1. What is the difference between mass and weight?

2. What unit is used to measure mass? Weight?

3. What is the small unit of a covalent compound? An ionic compound?

Molar mass tells us the mass ("weight") of 1 mol of an atom or compound. In each case **we simply calculate the sum of the "weights" of the atoms in the formula to determine the weight of a mole. These weights can be found on the periodic table.**

EXAMPLE: Calculate the molar mass (gram molecular weight) of a mole of iodine, I_2. Round to 2 decimal places.

2 I = 2 X (126.90) = **253.80 g I_2/mol**

EXAMPLE: Calculate the molar mass (gram formula weight) of a mole of aluminum sulfate, $Al_2(SO_4)_3$. Round to 2 decimal places.

2 Al = 2 X (26.98) = 53. 96
3 S = 3 X (32.07) = 64.14
+ 12 O = 12 X (16.00) = 192.00

$Al_2(SO_4)_3$ = **310.10 g $Al_2(SO_4)_3$ /mol**

CALCULATE THE MOLAR MASS FOR THE FOLLOWING COMPOUNDS OR DIATOMIC ELEMENTS. SET UP EACH PROBLEM AS SHOWN IN THE EXAMPLE ABOVE. INCLUDE UNITS (G/MOL)

		FORMULA	CALCULATION	MASS OF A MOLE
1.	water	________________		________________
2.	calcium chloride	________________		________________
3.	copper(II) sulfate	________________		________________
4.	silver nitrate	________________		________________
5.	sulfuric acid	________________		________________

		FORMULA	CALCULATION	MASS OF A MOLE
6.	calcium phosphate	________________		_________________
7.	sodium carbonate	________________		_________________
8.	ammonia	________________		_________________
9.	potassium chlorate	________________		_________________
10.	lead(II) nitrate	________________		_________________
11.	sodium oxalate	________________		_________________
12.	zinc chloride	________________		_________________
13.	magnesium oxide	________________		_________________
14.	antimony(III) chloride	________________		_________________
15.	nitrogen	________________		_________________
16.	oxygen	________________		_________________
17.	fluorine	________________		_________________
18.	chlorine	________________		_________________

CHEMISTRY WORKSHEET # 2: THE MOLE AS A UNIT OF MASS

Define the term molar mass (worksheet #1): ____

Now that you know how to find the mass of one mole of a substance (molar mass) you can easily find the mass of several moles or the mass of a fraction of a mole using the factor-label technique.

EXAMPLE: What is the mass of 5.00 moles of water(H_20)?

2 H = 2 x(1.01) = 2.02
O = 1 x(16.00) = 16.00

H_20 = 18.02 g

$$\# \text{ grams } H_20 = 5.00 \text{ moles } \cancel{H_20} \times \frac{18.02 \text{ g } H_20}{1 \text{ mole } \cancel{H_20}} = \mathbf{90.10 \text{ g } H_2O}$$

NOW YOU TRY ONE: What is the mass of 0.50 moles of calcium carbonate($CaCO_3$)?

Ca =
C =
3O =

$CaCO_3$ =

#g $CaCO_3$ =

USE A SEPARATE SHEET OF PAPER TO SOLVE THE FOLLOWING PROBLEMS. SHOW YOUR WORK. ROUND MOLAR MASSES TO TWO PLACES AFTER THE DECIMAL. ADD UNITS.

1. How many grams are there in 5.00 moles of lead?

2. How many grams are there in 2.00 moles of sulfuric acid?

3. How many grams are there in 0.250 moles of sodium hydroxide?

4. How many grams are there in 2.50 moles of potassium nitrate?

5. How many grams are there in 10.0 moles of lithium carbonate?

#1-5: You are given the number of moles and asked to find the grams (mass)

MOLS → GRAMS

Now that you know how to find the mass of one mole of a substance you can easily find the number of moles there are in a given mass of the substance.

EXAMPLE: How many moles of calcium chloride are there in 333 grams of calcium chloride ($CaCl_2$)?

Ca = 1 x(40.08)= 40.08
2 Cl = 2 x(35.45)= 70.90
$CaCl_2$ = 110.98 g

$$\# \text{ moles } CaCl_2 = 333 \text{ grams } \cancel{CaCl_2} \times \frac{1 \text{ mole } CaCl_2}{110.98 \text{ g } \cancel{CaCl_2}} = \mathbf{3.00 \text{ mole } CaCl_2}$$

USE THE SAME PAPER AS THE ABOVE PROBLEMS TO SOLVE THE FOLLOWING. SHOW YOUR WORK AND PUT UNITS ON EACH ANSWER!

6. How many moles of silver nitrate are there in 80.00 grams of silver nitrate?

7. How many moles of phosphoric acid are there in 658 grams of phosphoric acid?

8. How many moles of tin (II) fluoride are there in 908 grams of tin (II) fluoride?

9. How many moles of hydrogen peroxide (H_2O_2) are there in 1000.0 grams of hydrogen peroxide?

10. How many moles of magnesium chloride are there in 148 grams of magnesium chloride?

#6-10: You are given the grams (mass) and asked to find the number of mols

GRAMS → MOLS

CHEMISTRY WORKSHEET # 3 AVOGADRO'S NUMBER

One important property of a mole is that it means a definite number of particles just like a dozen means a number of particles. While a dozen is only 12 particles a **mole is a much larger number—6.02 x 10^{23} particles.** Elements generally exist as the particles we call atoms. **A mole of carbon contains 6.02 x 10^{23} atoms of carbon.** A mole of helium contains 6.02 x 10^{23} atoms of helium. A mole of sodium contains 6.02 x 10^{23} atoms of sodium. A mole of gold contains 6.02 x 10^{23} atoms of gold. However, **we have learned about seven elements that exist as diatomic molecules—H_2, N_2, O_2, F_2, Cl_2, Br_2, and I_2. For these elements one mole is 6.02 x 10^{23} molecules. That is, 6.02 x 10^{23} molecules of hydrogen is one mole of hydrogen,** 6.02 x 10^{23} molecules of nitrogen is one mole of nitrogen, 6.02 x 10^{23} molecules of oxygen is one mole of oxygen, etc.

While atoms are the smallest part of an element that still retains the properties of that element, molecules are the smallest parts of covalent compounds that still retain the properties of that compound. (For ionic compounds the smallest part is a combination of + and – ions but for now lets just consider them to be "molecules".) Therefore, one mole of a compound contains 6.02 x 10^{23} molecules of that compound. **One mole of water contains 6.02 x 10^{23} molecules of water**, one mole of carbon dioxide contains 6.02 x 10^{23} molecules of carbon dioxide, one mole of ammonia contains 6.02 x 10^{23} molecules of ammonia, one mole of sodium chloride contains 6.02 x 10^{23} "molecules" of sodium chloride, etc. (The number 6.02 x 10^{23} is a measurement, not a definition, and is only good for three significant figures.)

In all of the above examples one mole of any substance contained the same number of particles. **But remember, they all had different masses.** The mass of one mole of each material was equal to the gram formula weight. (This is the same idea as the mass of a dozen. A dozen eggs, a dozen bricks, a dozen dump trucks all contain twelve items but the mass of a dozen eggs is certainly much different than the mass of a dozen bricks which is much different from the mass of a dozen dump trucks!)

The number **6.02 x 10^{23} is known as Avogadro's number** in honor of an Italian Professor of physics, Amadeo Avogadro, who did considerable work on the development of atomic theory and the mole concept in about 1810. Given this number we can calculate the number of particles in a known number of moles or the number of moles in a given number of particles.

EXAMPLE: How many molecules of water are there in 3.00 moles of water?

$$\text{\# molecules } H_20 = 3.00 \cancel{\text{ moles}}\ H_20 \times \frac{6.02 \times 10^{23} \text{ molecules of } H_20}{1 \cancel{\text{ mole}}\ H_20} = \mathbf{1.81 \times 10^{24} \text{ molecules } H_20}$$

EXAMPLE: How many moles of neon are there in 2.408 x 10^{24} atoms of neon?

$$\text{\# moles Ne} = 2.408 \times 10^{24} \cancel{\text{ atoms}} \text{ Ne} \times \frac{1 \text{ mole Ne}}{6.02 \times 10^{23} \cancel{\text{ atoms}} \text{ of Ne}} = \mathbf{4.00 \text{ moles Ne}}$$

USE A SEPARATE SHEET OF PAPER TO SET-UP AND SOLVE THE FOLLOWING PROBLEMS.

How many molecules are there in:

1. 2.00 moles of ammonia (NH_3)?
2. 0.50 moles chlorine?
3. 0.250 moles oxygen?
4. 4.00 moles of sulfur dioxide?
5. 2.50 moles of methane?

How many moles are there in:

6. 3.612 x 10^{24} molecules of phosgene ($COCl_2$)?
7. 3.01 x 10^{23} molecules of Freon ($CHClF_2$)?
8. 1.505 x 10^{24} molecules of sucrose ($C_{12}H_{22}O_{11}$)?
9. 1.806 x 10^{24} molecules of bromine?
10. 3.01 x 10^{24} atoms of argon?

CHEMISTRY WORKSHEET # 4 MOLAR MASSES AND AVOGADRO'S NUMBER

Now that you know two definitions of a mole (a gram formula weight and an Avogadro's number of particles) you can combine these two definitions into one problem.

EXAMPLE: How many molecules are there in 90.1 grams of water?

$$\begin{array}{lll} 2\text{ H} = & 2 \times (1.01) = & 2.02 \\ \underline{\text{O} =} & \underline{1 \times (16.00) =} & \underline{16.00} \\ H_2O & & = 18.02\text{ g} \end{array}$$

$$\#\text{ molecules } H_2O = 90.1\text{ g } \cancel{H_2O} \times \frac{1\text{ mole } \cancel{H_2O}}{18.02\text{ g } \cancel{H_2O}} \times \frac{6.02 \times 10^{23}\text{ molecules } H_2O}{1\text{ mole } \cancel{H_2O}} = \mathbf{3.01 \times 10^{24}\text{ molecules } H_2O}$$

EXAMPLE: What would be the mass of 3.01×10^{23} molecules of ammonia (NH_3)?

$$\begin{array}{lll} \text{N} = & 1 \times (14.01) = & 14.01 \\ \underline{3\text{ H} =} & \underline{3 \times (1.01) =} & \underline{3.03} \\ NH_3 & & = 17.04\text{ g} \end{array}$$

$$\#\text{ grams } NH_3 = 3.01 \times 10^{23}\text{ } \cancel{\text{molecules}}\text{ } NH_3 \times \frac{1\text{ } \cancel{\text{mole}}\text{ } NH_3}{6.02 \times 10^{23}\text{ } \cancel{\text{molecules}}\text{ } NH_3} \times \frac{17.04\text{ g } NH_3}{1\text{ } \cancel{\text{mole}}\text{ } NH_3} = \mathbf{8.52\text{ g } NH_3}$$

SOLVE THE FOLLOWING PROBLEMS ON A SEPARATE SHEET OF PAPER. YOU MUST SHOW ALL OF THE STEPS AND YOU MUST DO THE PROBLEM JUST AS ILLUSTRATED. INCLUDE UNITS!

1. How many molecules are there in 345 grams of carbon dioxide?
2. What would be the mass, in grams, of 1.204×10^{24} molecules of sulfur dioxide?
3. How many molecules of sucrose, $C_{12}H_{22}O_{11}$, are there in 454 grams of sucrose?
4. What would be the mass, in grams, of 1.806×10^{24} molecules of carbon monoxide?
5. How many molecules of water are there in 8.050×10^{3} grams of water?
6. How many oxygen molecules are in a flask that contains 1.43 grams of oxygen?
7. What would be the mass, in grams, of 1.505×10^{23} molecules of carbon disulfide?
8. How many molecules of hydrogen chloride would there be in 100.00 grams of this gas?
9. What would be the mass, in grams, of 2.408×10^{24} molecules of tetraphosphorus decaoxide?

Extra Challenge:

10. How many hydrogen molecules are there in 1 ton of hydrogen? (Hint: How many grams are there in 1 ton?)

CHEMISTRY WORKSHEET # 5 MOLE PROBLEMS—MOLAR VOLUME OF A GAS

We have learned that a mole is a mass of material and number of particles. **A mole can also be a measure of volume when we are talking about gases**. You may remember from previous science classes that all gases behave basically the same as far as the physical properties of temperature, pressure and volume. **AVOGADRO'S HYPOTHESIS SAYS THAT EQUAL VOLUMES OF GASES AT THE SAME TEMPERATURE AND PRESSURE CONTAIN EQUAL NUMBERS OF MOLECULES.** Avogadro's statement makes sense and is possible because gases are mainly empty space—only about one thousandth of the space is actually filled with molecules. The molecules "fill" the remaining space by moving rapidly through it. So the difference in size between large molecules and small molecules is insignificant compared to the total volume the gas occupies. At **standard temperature and presssure (STP = O°Celcius and 1.00 atm pressure) one mole of any gas will have a volume of 22.4 liters.** In other words, **THE MOLAR VOLUME OF ANY GAS IS 22.4 LITERS AT STP.** Once we know this we can convert from moles to liters or liters to moles for any gas at STP.

EXAMPLE: What is the volume, in liters, of a 2.00 mole sample of methane (CH_4) at STP?

$$\#\ L\ CH_4 = 2.00\ \cancel{\text{moles}}\ CH_4 \times \frac{22.4\ L\ CH_4}{1\ \cancel{\text{mole}}\ CH_4} = \mathbf{44.80\ L\ CH_4}$$

EXAMPLE: How many moles of ethane (C_2H_6) are there in 5.60 liters of ethane?

$$\#\ \text{moles}\ C_2H_6 = 5.60\ \cancel{L\ C_2H_6} \times \frac{1\ \text{mole}\ C_2H_6}{22.4\ \cancel{L\ C_2H_6}} = \mathbf{0.25\ mole\ C_2H_6}$$

COMPLETE THE FOLLOWING PROBLEMS ON A SEPARATE SHEET OF PAPER USING THE SAME SET-UP AS SHOWN ABOVE. WRITE THE FORMULAS FOR THE COMPOUNDS, NOT JUST THE NAMES. INCLUDE UNITS!

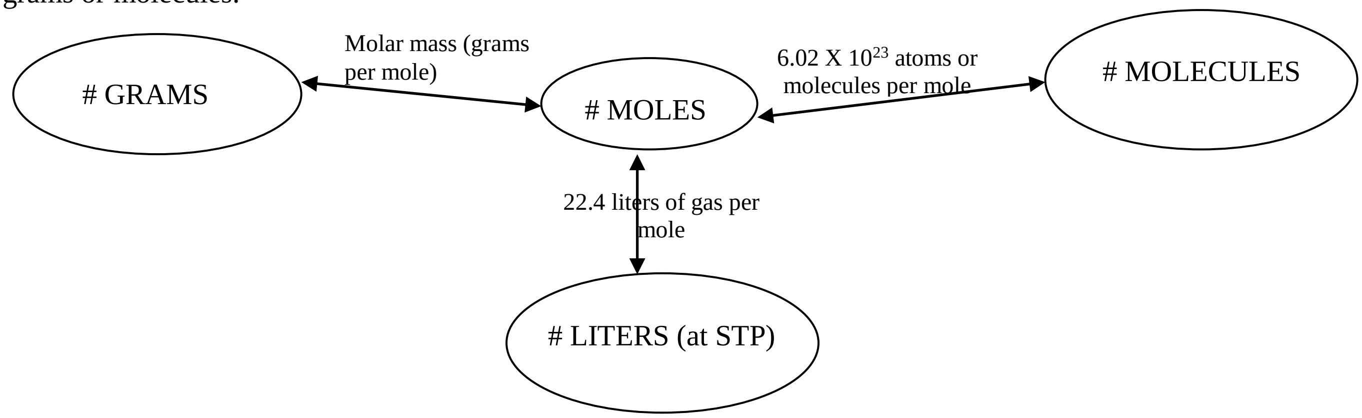
1. What is the volume, in liters, of 2.00 moles of hydrogen at STP?
2. How many liters will 5.00 moles of oxygen occupy at STP?
3. What is the volume, in liters, of 0.250 moles of carbon monoxide at STP?
4. What is the volume, in liters, of a 3.00 mole sample of carbon dioxide at STP?
5. How many moles of chlorine are there in a 67.2 liter sample of chlorine at STP?
6. A 44.8 liter sample of nitrogen at STP will contain how many moles of nitrogen?
7. How many moles of ammonia are there in 405 liters of ammonia at STP?
8. How many moles of neon would you need to fill a 33.6 liter container at STP?
9. How many moles of argon are there in 5.00×10^2 liters of argon at STP?
10. What is the volume, in liters, of 4.50 moles of fluorine at STP?

Extra Challenge

11. How many moles of nitrogen are there in a 16,500 mL sample of nitrogen at STP?

CHEMISTRY WORKSHEET # 6 MIXED MOLE PROBLEMS (GRAMS, MOLECULES, AND LITERS)

You now know three things a mole can be: a molar mass, 6.02×10^{23} molecules and, for a gas, 22.4 liters at STP. We can use this information to convert grams to molecules or liters, molecules to grams or liters, or liters to grams or molecules.



EXAMPLE 1: What would be the volume in liters of 40.36 grams of neon at STP?

$$\#\ \text{liters Ne} = 40.36\ \cancel{\text{g Ne}} \times \frac{1\ \cancel{\text{mole Ne}}}{20.18\ \cancel{\text{g Ne}}} \times \frac{22.4\ \text{L Ne}}{1\ \cancel{\text{mole Ne}}} = \mathbf{44.80\ L\ Ne}$$

EXAMPLE 2: How many molecules would there be in 56 liters of carbon dioxide at STP?

$$\#\ \text{molecules } CO_2 = 56.0\ \text{L}\ \cancel{CO_2} \times \frac{1\ \cancel{\text{mole } CO_2}}{22.4\ \text{L}\ \cancel{CO_2}} \times \frac{6.02 \times 10^{23}\ \text{molecules } CO_2}{1\ \cancel{\text{mole } CO_2}} = \mathbf{1.51 \times 10^{24}\ molecules\ CO_2}$$

SOLVE THE FOLLOWING PROBLEMS ON A SEPARATE SHEET OF PAPER.

- YOU MUST USE COMPLETE AND PROPER SET-UPS.
- SHOW THE MOLAR MASS CALCULATION WHENEVER THE PROBLEM REQUIRES YOU TO DO ONE. INCLUDE UNITS
- CIRCLE YOUR FINAL ANSWER.

1. What would be the volume, in liters, of 85.5 grams of carbon monoxide at STP?

2. How many molecules would there be in 0.500 grams of carbon disulfide?

3. What would be the mass, in grams, of 45.0 liters of nitrogen at STP?

4. How many molecules of hydrogen are in a balloon full of hydrogen with a volume of 5.34 liters at STP?

5. Your mommy buys you a helium balloon at the circus. It has a volume of 4.00 liters at STP. What mass of helium, expressed in grams, does this balloon contain?

6. How many molecules of ammonia would there be in 40.0 grams of ammonia?

7. What would be the mass, in grams, of 3.50×10^{25} molecules of chlorine?

8. What volume, expressed in liters, would 50.0 grams of fluorine occupy at STP?

9. How many grams of oxygen would there be in 1.00 liter of oxygen at STP?

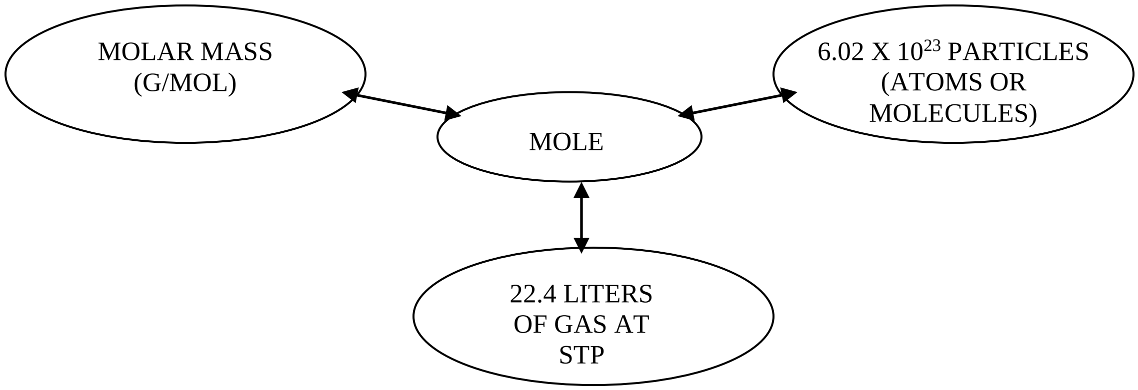
Extra Challenge:

10. How many molecules of water are there in 10 lbs of water?

CHEMISTRY WORKSHEET # 7: GENERAL REVIEW OF MOLE PROBLEMS

Now we have studied the idea of moles and learned three interpretations of a mole:

(1) A gram formula weight.
(2) An Avogadro's number of particles.
(3) 22.4 liters of gas at STP.



Solve the following problems involving the mole concept. (If you are having difficulty go back and review mole worksheets 1-6.)

Problems 1-2: moles to grams AND grams to moles

1. How many grams are there in 11.8 moles of sodium hydroxide?

   Ans. 472 grams sodium hydroxide

2. How many moles are there in 215 grams of water?

   Ans. 11.9 moles water

Problems 3-4: moles to molecules AND molecules to moles

3. How many molecules are there in 3.85 moles of carbon tetrachloride?

   Ans. 2.32×10^{24} molecules carbon tetrachloride

4. How many moles are there in 8.25×10^{26} molecules of methane?

   Ans. 1.37×10^{3} moles of methane

Problems 5-6: grams to moles to molecules AND molecules to moles to grams

5. How many molecules are there in 295 grams of ammonia?

   Ans. 1.04×10^{25} molecules of ammonia

6. How many grams are there in 8.95×10^{26} molecules of carbon disulfide?

   Ans. 1.13×10^{5} grams of carbon disulfide

Problems 7-8: moles to liters AND liters to moles

7. What would be the volume, in liters measured at STP, of 9.75 moles of carbon monoxide?
Ans. 2.18×10^2 liters of carbon monoxide

8. How many moles would there be in 5.25 liters of oxygen measured at STP?
Ans. 0.234 moles or 2.34×10^{-1} moles oxygen

Problems 9-10: grams to moles to liters AND liters to moles to grams

9. What is the volume, measured in liters at STP, of 285 grams of the gas acetylene, C_2H_2?
Ans. 245 liters of acetylene

10. How many grams are there in 512 liters (measured at STP) of propane, C_3H_8?
Ans. 1.01×10^3 grams of propane

Problems 11-12: molecules to moles to liters AND liters to moles to molecules

11. What would the volume be, measured in liters at STP, of 3.01×10^{25} molecules of fluorine?
Ans. 1.12×10^3 liters of fluorine

12. How many molecules are there in 995 liters of sulfur dioxide at STP?
Ans. 2.67×10^{25} molecules of sulfur dioxide

Problems 13-16: Mixed Problems- Think about what type of conversion you are doing!

13. How many molecules are there in 2270 g of table sugar, sucrose.
Ans. 3.99×10^{24} molecules of sucrose

14. How many molecules would there be in 1.135×10^6 g of chlorine?
Ans. 9.64×10^{27} molecules of chlorine

15. What would the mass be, in grams, of 348 liters of carbon dioxide measured at STP?
Ans. 684 grams of carbon dioxide

16. How many molecules of nitrogen are there in 200 L of nitrogen measured at STP?