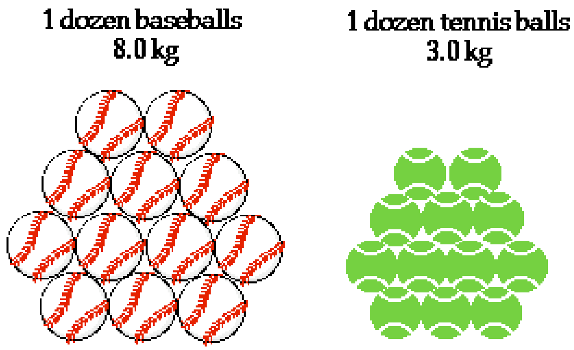
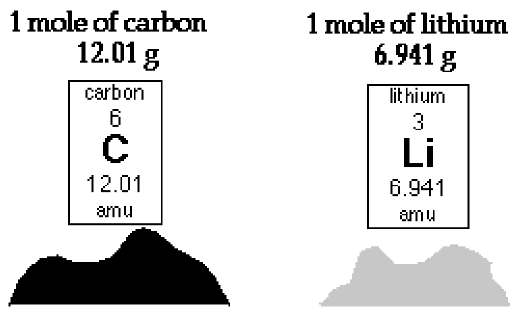
Moles

I. **Lab**: Rice Counting
II. Counting atoms and molecules
    I. When doing reactions chemists need to count atoms and molecules. The problem of actually counting individual atoms and molecules comes from the fact that they are so small. Instead of physically counting each one we can measure the mass of a large group of them. If we know the mass of just one atom or molecule then we can figure out how many we have.
    II. Because we will always be dealing with groups of atoms and molecules it would be convenient to give a name to a certain number of them. Just as we can talk about groups of objects like a dozen (12) eggs or a gross (144) of pencils, we need a unit for a large number of atoms or molecules. This unit is the **mole.**
    III. **1 mole = 6.02×10^{23}. This is also known as Avagadro's number**, in honor of his contributions toward the development of atom counting experiments and theories.
    IV. To give you some idea of how large a mole of items is think about the following:
        i. Spreading one mole of rice grains over the surface of the earth would create a layer one meter deep.
        ii. If you had a mole of pennies and decided to give every person on Earth $1,000,000 a day, then it would take about 3000 years to spend all of your pennies.
        iii. Even with today's high speed technology we still need to count by weighing. A computer counting 1 million atoms a second would take 20 years to count out one mole of atoms or molecules.
    V. **Demo**: Tray of moles.
    VI. Weighing atoms
        i. So, it's too hard to count atoms but we can measure the mass of a large group and know how many we have. To do this we need to know the mass of an individual atom or the mass of a specific number of atoms.
        ii. The mass of one carbon-12 atom is very tiny - 1.99×10^{-23}g.
        iii. This mass is a bit cumbersome to deal with so a unit of mass was invented to make it easier to discuss the mass of atoms and molecules - the atomic mass unit (amu). It was defined such that the carbon-12 atom was exactly equal to 12.000 **amu** The mass of every other atom was then measured relative to this standard. If an atom weighed twice as much as a carbon-12 atom it had a mass of 24.000 **amu**
        iv. So someone thought, "How many atoms of carbon-12 would there be in 12.000g of those atoms?". The answer is 6.02×10^{23} or 1.00 mol.
        v. Because 12.000g of carbon is an easily manageable amount of substance to use, the mole became our standard unit for counting the numbers of atoms or molecules.
        vi. Every atomic mass on the periodic table has the unit **amu**, and if we have enough of an element measured out in grams to equal its atomic mass then we have 1.00 mol of that substance.

vii. Remember the term mole is just like the term dozen, it just represents a certain number of things. See the illustration bellow for a clarification of the relationship between atomic mass and the mole.



Notice the same number of baseballs and tennis balls can be referred to as a group (a dozen) but they can have different masses. The same is true for substances, except we refer to moles instead of dozens.



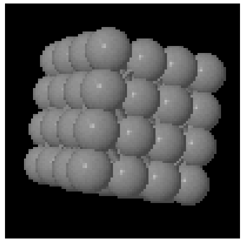
III. Molar Mass and % Composition

I. Molar Mass

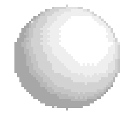
i. The molar mass of a substance is defined as the mass of one mole of particles of that substance.

ii. A particle or unit of a substance is defined by its formula. When doing calculations we need to relate the mass of a substance to the mass of one mole of that substance. For monatomic elements this is easy. The mass of one mole is equivalent to the atomic mass in grams. This is commonly referred to as the **molar mass.** However, to calculate the molar mass of any other type of substance we use the formula to determine how many moles of atoms are contained in one mole of the substance. For example:

a. Copper as a general substance looks like this --->

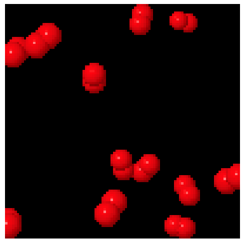


Its formula is **Cu**, so a mole of copper means a mole of these -->

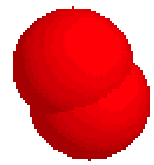


1 mol of copper atoms =
1 x 63.55g = **63.55g**

b. Oxygen as a general substance looks like this --->

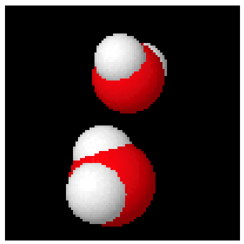


Its formula is **O_2**, so a mole of oxygen means a mole of these -->

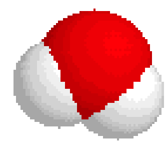


1 mol of oxygen molecules =
2 x 16.00g = **32.00g**

c. Water as a general substance looks like this --->

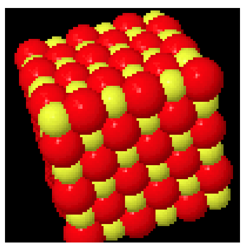


Its formula is **H_2O**, so a mole of water means a mole of these -->

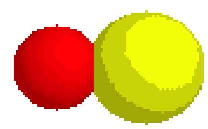


1 mol of water molecules =
2 x 1.008g = 2.016g
1 x 16.00g = <u>16.00g</u>
18.02g

d. Sodium Chloride as a general substance looks like this --->



Its formula is **NaCl**, so a mole of sodium chloride means a mole of these -->



1 mol of sodium chloride =
1 x 22.99g = 22.99g
1 x 35.54g = <u>35.54g</u>
58.44g

iii. What is the molar mass of calcium phosphate - $Ca_3(PO_4)_2$?

II. Percent Composition
  i. One way to identify compounds is to determine what percentage each element is by mass in a compound.
  ii. See the example of water from the above table. The molar mass of sodium chloride is 58.44g. 22.99g of this is sodium and 35.45g is chlorine. If I want to calculate the percent composition of sodium chloride, I would do the following:

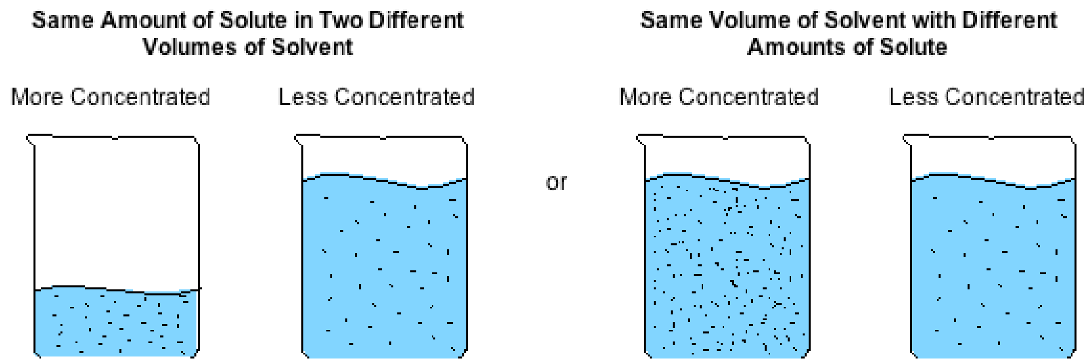
$$Na = \frac{22.99g}{58.44g} \times 100 = 39.34\%$$

$$Cl = \frac{35.45g}{58.44g} \times 100 = 60.66\%$$

  iii. What is the percent composition of calcium phosphate?

i. **Homework**: Molar Mass and % Composition sheet;
ii. **Lab**: Copper Oxide Lab

IV. Molarity (moles in solutions)
  I. When working with solutions it is important to know how concentrated the solution is.
  II. Obviously, dropping just a pinch of sugar in some tea will be much less sweet than dumping three heaping table-spoons into the drink. The difference is the second drink is a much more concentrated sugar solution.



  I. If you are working with acids, the concentration of the solution can tell you the difference between a substance that can be handled safely or one that can only be opened in a ventilated hood.

II. Our unit for solution concentration is **MOLARITY** and uses the abbreviation: **M**. Molarity describes how many **moles** are dissolved **per liter of of solution** not per liters of solvent.

III. It is calculated using the following formula:

$$\text{Molarity} = \frac{\text{moles of solute}}{\text{liter of solution}} = \text{M or (mol/L)}$$

IV. Try the following examples:

i. What was the molarity of a solution if you took 600.0 mL of it, evaporated the water, and got 3.4 mols of solute?

ii. What is the molarity of a solution in which 1.00 grams of NaCl are dissolved in 10.0 mL of the solution?

V. Notice that molarity is moles per **liter of solution, not liter of solvent.** This has implications for how you will make a solution. If I handed you a pile of salt and asked you to make one liter of aqueous solution with this salt, how would you do it?

VI. Let's be more specific: If I asked you to make 500.0 mL of a 3.0 M solution of NaOH. How would you do it?Click here for answer.

V. Conversions

I. Making conversion factors

i. Whenever we have an equality between two different units we can make a conversion factor. These factors can be multiplied by a number we wish to convert from one unit to another.

ii. 1 inch = 2.54 cm so we get two conversion factors:

$$\frac{2.54\ \text{cm}}{1\ \text{in}} \text{ and } \frac{1\ \text{in}}{2.54\ \text{cm}}$$

iii. To convert 12.0 in to cm, we would multiply by the first conversion factor:

$$12.0\ \text{in} \times \frac{2.54\ \text{cm}}{\text{in}} = 30.5\ \text{cm}$$

iv. To convert 45.8 cm to in, we would multiply by the second conversion factor:

$$45.8\ \text{cm} \times \frac{\text{in}}{2.54\ \text{cm}} = 19.1\ \text{in}$$

v. During this year we will be making frequent conversions between grams and moles.

vi. Fortunately, the molar mass gives us an equality between these two units. For all substances 1 mol = molar mass. For example, for carbon, 1 mol = 12.01g. We therefore get the two conversion factors:

$$\frac{12.01\ \text{g}}{1\ \text{mol}} \text{ and } \frac{1\ \text{mol}}{12.01\ \text{g}}$$

vii. What is the mass of 0.438 mol of carbon?

viii. How many moles of carbon do you have if you have 2.56g of carbon?

II. For solutions, you get conversion factors from the concentration. Molarity is the same as mol/L, so you can use molarity to convert between volumes of soultion and moles of solute.

$$6.53M = 6.53\frac{mol}{L} \quad \text{so you get conversion factors of} \quad \frac{6.53mol}{1L} \quad \text{and} \quad \frac{1L}{6.53mol}$$

III. How many liters of 2.45 M solution would you need to have 0.75 mols of solute?

IV. How many mols of solute would you have if you had 35.0L of a 0.65 M solution?

i. **Homework**: Converting between mols, grams, and volume of solution.

VI. Empirical Formulas

I. An empirical formula states the simplest whole number ratio of atoms in a substance.

II. All ionic formulas are empirical (for example we write NaCl instead of Na_2Cl_2 or Na_3Cl_3) and some molecular formulas are empirical.

III. For example:
   i. Aluminum sulfide (ionic) is Al_2S_3 and is empirical.
   ii. Water (molecular) is H_2O and is empirical.
   iii. Benzene (molecular) is C_6H_6 which is not empirical. Its empirical formula is CH.

IV. If we have a sample of a substance and can determine the mass of each element present in the sample then we can determine its empirical formula.

V. To understand how to do this, let's take water for example. In one mole of water we have 2.016g of H and 16.00g of O. If we convert each of these masses into moles we get the familiar ratio: 2.000 moles of H to 1.000 mole of O. See calculation below:

$$H = 2.016g \times \frac{1 \text{ mol}}{1.008 \text{ g}} = 2.000 \text{ mol}$$

$$O = 16.00g \times \frac{1 \text{ mol}}{16.00 \text{ g}} = 1.000 \text{ mol}$$

VI. If we only had one quarter of a mole of water then the H mass would be 0.504g and the O mass would be 4.00g. If we convert these masses to moles then we get:

$$H = 0.504g \times \frac{1 \text{ mol}}{1.008 \text{ g}} = 0.500 \text{ mol}$$

$$O = 4.00g \times \frac{1 \text{ mol}}{16.00 \text{ g}} = 0.250 \text{ mol}$$

VII. By simplifying the mole ratio we still get a nice whole number ratio, from which we can construct the correct empirical formula - H_2O:

$$H = \frac{0.500 \text{ mol}}{0.250 \text{ mol}} = 2.00 \text{ mol}$$

$$O = \frac{0.250 \text{ mol}}{0.250 \text{ mol}} = 1.00 \text{ mol}$$

VIII. If you had a sample of aluminum chloride that contained 20.23g of aluminum and 79.77g of chlorine, what would be its empirical formula? Click here to see answer below.

IX. Given a percent composition you should also be able to calculate the empirical formula. If you know something is 30% carbon and 70% oxygen, just hypothesize a theoretical sample of any arbitrary size and determine what the mass of carbon and oxygen would be. For example, if you had a sample of the above compound with a mass of 100g, then 30g would be carbon and 70g would be oxygen. You can now start the empirical formula calculations.

i. **Homework**: Calculating Emperical Formulas

ii. **Lab**: Empirical Formula Lab

VII. Molecular Formulas

I. In order to calculate the molecular formula for molecular substances you do the same procedure as finding the empirical formula and add one more step that requires one more piece of information - the molar mass of the molecular formula.

II. To understand how the last step works consider the relationship between the empirical and molecular formula of benzene:

III. Notice that if you divide the molecular formula mass by the empirical formula mass you get a factor by which you should multiply each atom of the empirical formula. So CH becomes C_6H_6.

IV. Try the following. If a compound with the empirical formula CH_2O was found to have a molar mass of 90g, what would the molecular formula be?

Empirical Formula	Molecular Formula
CH	C_6H_6
Molar Mass	Molar Mass
13.02g	78.12g

$$\frac{\text{Molar mass of molecular formula}}{\text{Molar mass of empirical formula}} = \frac{78.12g}{13.02g} = 6.000$$

i. **Homework**: Calculating Molecular Formulas

VIII. Stoichiometry

i. Moles-Moles

I. Stoichiometry describes the relationships between reactants and products in chemical equations.

II. A chemical equation can be interpreted in several different ways. The most basic way to think about a balanced chemical equation is to think of it as the simplest recipe for a particular reaction. Take the following equation as an example: CH_4 + 3 CuO ---> 3 Cu + 2 H_2O + CO. The simplest way of interpreting this is in the following way: 1 molecule of methane reacts with 3 units of copper(II) oxide to form 3 atoms of copper, 2 molecules of water, and 1 molecule of carbon monoxide.

III. But this is just the basic recipe. We can multiply the recipe many times if we wish, which can lead to new interpretations of the equation:

	CH_4 +	3 CuO	--->	3 Cu +	2 H_2O +	CO
basic	**1 molecule**	**3 units**	--->	**3 atoms**	**2 molecules**	**1 molecule**
x2	2 molecules	6 units	--->	6 atoms	4 molecules	2 molecules
x12	12 molecules	36 units	--->	36 atoms	24 molecules	12 molecules
	1 dozen	3 dozen	--->	3 dozen	2 dozen	1 dozen
x6.02e+23	6.02e+23	1.81e+24	--->	1.81e+24	1.20e+24	6.02e+23
	1 mole	**3 moles**	--->	**3 moles**	**2 moles**	**1 mole**

IV. Chemical equations are, in fact, usually interpreted as a representation of the molar ratio of reactants and products in an equation as represented by the last line of the table above.

V. As long as we keep the ratio of one substance to another we can perform a chemical reaction using any amount of materials that we want. We can use the chemical equation as a guide for adjusting our recipe to match any particular experiment that we might do.

VI. Moles - Moles calculations

i. Consider the following problem: Given the reaction below, how many moles of **aluminum oxide** will you produce if you start with 5.40 mol of **aluminum**?

$$4\,Al + 3\,O_2 \longrightarrow 2\,Al_2O_3$$

ii. First think about what the chemical equation states. The question asks about the relationship between aluminum oxide and aluminum. We can see from the reaction that if 4 moles of Al are used then 2 moles of Al_2O_3 will be produced. This ratio is fixed, so if we start with 5.40 moles of Al instead, how do we predict the amount of Al_2O_3 produced?

iii. To do this we must set up a ratio between the two substances asked about in the question.

$$\begin{array}{ccccccc} \mathbf{5.40} & & & & \mathbf{x} \\ \mathbf{4}\,Al & + & 3\,O_2 & \longrightarrow & \mathbf{2}\,Al_2O_3 \end{array}$$

so

$$\frac{5.40\text{ mol}}{4\text{ mol}} = \frac{x}{2\text{ mol}}$$

iv. By solving for x, you will determine the moles of aluminum oxide produced.

v. So, the balanced chemical equation can be used to set up a **MOLE RATIO** between any two substances in the equation.

vi. It is important to remember that you can only set up a

MOLE RATIO!

VII. This is emphasized because you will try to do this with numbers containing other units, like grams. **Do not try to set up a ratio using the mass of a substance!** This will not work because **the coefficients in a chemical equation represent molar quantities (numbers of things), not the mass of things**.

ii. **Lab**: Moles Verification

iii. Grams-Grams

I. Grams-Grams

i. When we are working with substances in the lab we use balances to measure out how much substance to work with. So, we need to use the reading from the balance to make similar predictions about the mass of substances produced or needed to complete a chemical reaction.

ii. To do this, we need only do the appropriate conversions from grams to moles and back again to utilize the **MOLE RATIO** we can derive from the balanced chemical equation.

iii. Consider the following problem: Given the reaction between aluminum and oxygen to produce aluminum oxide, what mass of aluminum oxide would be produced if we started with 3.26g of aluminum?

iv. In order to solve this, we will need to use the chemical equation to tell us the relationship between the moles of Al and the moles of Al_2O_3. However, our Al quantity is in grams. **First we must convert the grams to moles.**

$$3.26 \text{ g of Al} \cdot \frac{\text{mol}}{26.98 \text{ g}} = 0.121 \text{ mol of Al}$$

v. Then, **set up a mole ratio**:

$$\overset{\mathbf{0.121}}{\mathbf{4}\,\text{Al}} + 3\,O_2 \longrightarrow \overset{\mathbf{x}}{\mathbf{2}\,Al_2O_3}$$

$$\frac{0.121 \text{ mol}}{4} = \frac{x}{2} \quad \text{so} \quad x = 0.0604 \text{ mol of } Al_2O_3$$

vi. Then, **convert the moles to grams**.

$$0.0604 \text{ mol of } Al_2O_3 \cdot \frac{101.96 \text{ g}}{\text{mol}} = 6.16 \text{ g of } Al_2O_3$$

vii. Now we have our answer: If we start with 3.26g of Al, we should expect to produce 6.16g of Al_2O_3.

iv. Limiting Reactants

I. Limiting Reactants

i. All of the above problems assume you are using only the precise amounts necessary to perform the reaction. Often you measure one of the reactants precisely and react it with an excess of another substance to make sure all of the first substance is consumed in the reaction.

ii. Consider the following problem. If you start with 12.0 mol of Al and 10.0 mol of O_2 which would run out first and how much would you have left over of the other substance?

iii. To solve this problem do two calculations, one in which you assume all of the Al is used and one in which all of the O_2 is used.

Assuming all of the Al is used	Assuming all of the O_2 is used
12.0 **x**	**x** **10.0**
4 Al + **3** O_2 ---> 2 Al_2O_3	**4** Al + **3** O_2 ---> 2 Al_2O_3
$\frac{12.0 \text{ mol}}{4} = \frac{x}{3}$	$\frac{x}{4} = \frac{10.0 \text{ mol}}{3}$
x = 9.00 mol of O_2 needed	x = 13.3 mol of Al needed

IV. As you can see from either calculation, we have more oxygen than can be used in this reaction. **Aluminum is the limiting reactant**. 1.00 mol of oxygen (10.0 mol - 9.00 mol) will be left over.

V. To do this given grams you will have to do the g->mol conversions before the above mole ratio calculations.

v. **Homework**: Problems Using Moles in Equations sheet

vi. **Homework**: Stoichiometry Grams/Grams sheet

vii. Stoichiometry with Solutions

I. As in all other stoichiometry problems, solutions problems also boil down to determining the moles of one substance, using the chemical equation to get the mole ratio between two substances, and determining the number of moles of the other substance.

II. For example, we might need to calculate the concentration of silver thiosulfate in used photographic solutions to determine if it should be considered hazardous waste. To do that we could react the silver thiosulfate with sodium chloride in the following reaction:

$$Ag_2S_2O_{3(aq)} + 2\ NaCl_{(aq)} \text{ ---> } 2\ AgCl_{(s)} + Na_2S_2O_{3(aq)}$$

i. If you got 15.8 g of AgCl from 250.0 mL of $Ag_2S_2O_3$ solution, what was the concentration of that solution:

ii. **First we must convert the grams to moles.**

$$35.0 \text{ g of AgCl} \times \frac{1 \text{ mol}}{143.32 \text{ g}} = 0.244 \text{ mol}$$

iii. Then, **set up a mole ratio**:

$$\overset{\mathbf{x}}{Ag_2S_2O_3} \quad + \quad 2\ NaCl \quad \text{--->} \quad \overset{\mathbf{0.244}}{\mathbf{2}\ AgCl} \quad + \quad Na_2S_2O_3$$

$$\frac{x}{1} = \frac{0.244 \text{ mol}}{2} \quad \text{so} \quad x = 0.122 \text{ mol of } Ag_2S_2O_2$$

iv. Then, **convert the moles to molarity** .

$$\text{concentration} = \frac{n}{V} = \frac{0.122 \text{ mol}}{0.250 \text{ L}} = 0.488 \frac{\text{mol}}{\text{L}} = 0.488 \text{ M}$$

III. There are lots of different ways molarity could be used in in stoichiometry. Try these problems as well:

IV. $PbNO_{3(aq)} + 2\ Cl^{-1}{}_{(aq)} \text{ ----> } PbCl_{2(s)} + NO_3^{-1}{}_{(aq)}$

i. Determine the concentration of chloride ion in a water sample if the sample volume was 25.0 mL and was reacted with 85.0 mL of 0.10M $PbNO_3$ solution. Click here for answer.

ii. What volume of 0.50 M $PbNO_3$ would be necessary to produce 1.34 grams of $PbCl_2$?

a. **Pre-Lab**: Making solutions
b. **Homework**: Solutions Stoichiometry problems
c. **Lab**: Titration

viii. Stoichiometry with Gasses

I. Just as in every other stoichiometry problem that you have done, they all boil down to knowing some information about one substance so that you can use the mole ratio depicted in the chemical equation to determine some information about another substance.

II. Try the following examples:

Given the chemical equation for Butane burning in a lighter:

$$2\ C_4H_{10(l)} + 13\ O_{2(g)} \text{--->} 8\ CO_{2(g)} + 10\ H_2O_{(g)}$$

i. How many liters of oxygen are needed to combine with 5.00g of butane if the pressure is 1.07atm and the temperature is 25.0°C? **12.8 L**

ii. If 2.0L of water are produced at a temperature of 500°C and a pressure of 734 mmHg, then how many grams of butane were burned? **0.35 g**

a. **Homework**: Gasses Stoichiometry Problems
b. **Lab**: Measureing Gasses

IX. **Homework**: Moles Review