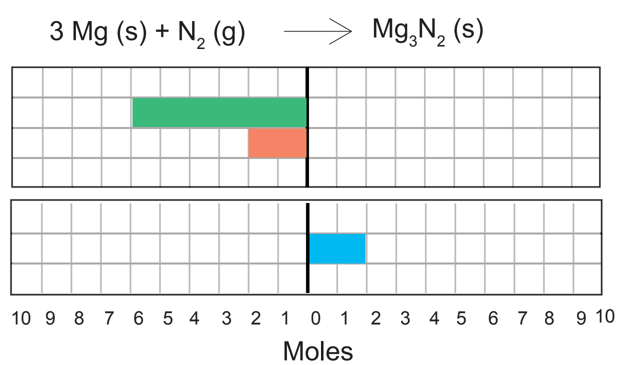
STOICHIOMETRY via ChemLog



by

Dr. Stephen Thompson
Mr. Joe Staley
Ms. Mary Peacock

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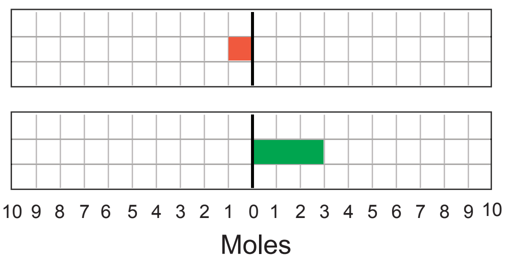
ChemLog STOICHIOMETRY

CONTENTS

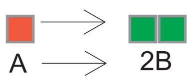
CHEMICAL REACTIONS

In the following reaction, one mole of reactant A goes to 2 moles of product B.

A → 2B



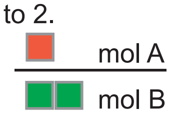
This can also be shown using "blocks" from the ChemLog.



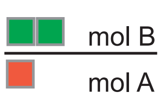
▶ From the ChemLog, how many moles of B are formed for each mole of A that reacts? To answer this question, fill in the appropriate numbers in this sentence: _____ mole(s) of A react(s) to form _____ mole(s) of B.

▶ From the ChemLog, how many moles of A are needed to decompose for each mole of B produced? To answer this question, fill in the appropriate number in this sentence: For every _____ mole(s) of B that is formed, _____ mole(s) of A decomposed.

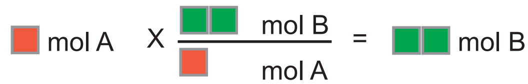
Another way to say this is that the ratio of moles of A to B is 1 to 2.



The ratio of moles of B to A is 2 to 1.



▶ We can use this ratio to answer the following question: when starting with one mole of A, how many moles of B will you obtain? Let's set this up mathematically.

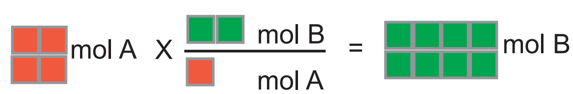


You might be wondering why we chose this ratio (with moles of B on top) rather than the other ratio. The trick is to remember to put the units you want (in our case, we want to get to moles of B) on top. In this example, moles of A cancels, and we're left with moles of B.

▶ We can answer this question simply by using numbers also. Notice that we get the same answer.

$$\underline{\ 1\ }\text{ mol A} \times \frac{\underline{\ 2\ }\text{ mol B}}{\underline{\ 1\ }\text{ mol A}} = \underline{\ 2\ }\text{ mol B}$$

▶ Now try this one, when starting with four moles of A, how many moles of B will you obtain?



▶ Fill in the appropriate numbers for this reaction.

$$\underline{\quad}\text{ mol A} \times \frac{\underline{\quad}\text{ mol B}}{\underline{\quad}\text{ mol A}} = \underline{\quad}\text{ mol B}$$

▶ Now try this one, when starting with 3 moles of A, how many moles of B will you obtain?

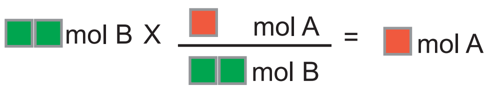
▶ When starting with one half a mole of A, how many moles of B will you obtain?

CHEMICAL REACTIONS

Let's continue working with the following reaction.

$$A \longrightarrow 2B$$

▶ Here's a slightly different, but similar question: when you get 2 moles of B from this reaction, how many moles of A did you start with? We'll go about answering it in the same manner.

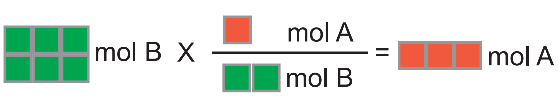


Notice that we're using a different ratio here. Our answer needs to be in units of "mol A", so we use the appriopriate ratio with moles of A on top.

▶ Fill in the appropriate numbers for this reaction.

___ mol B X (___ mol A / ___ mol B) = ___ mol A

▶ When you get 6 moles of B from this reaction, how many moles of A did you start with?



▶ Fill in the appropriate numbers for this reaction.

___ mol B X (___ mol A / ___ mol B) = ___ mol A

▶ When you get 3 moles of B from this reaction, how many moles of A did you start with?

▶ How many moles of B are produced in this reaction when you start with 3 moles of A?

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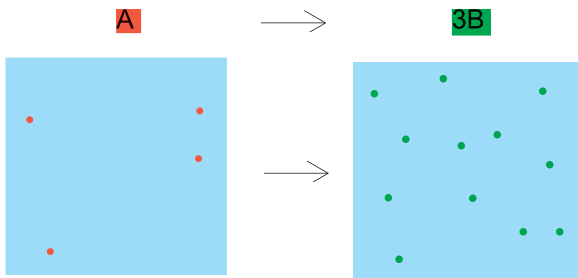
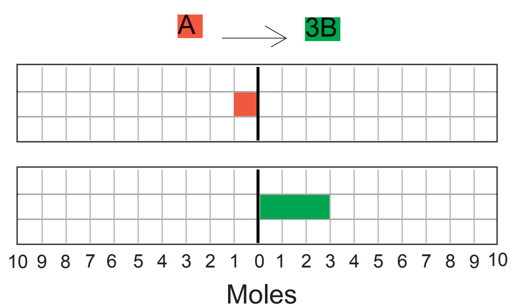
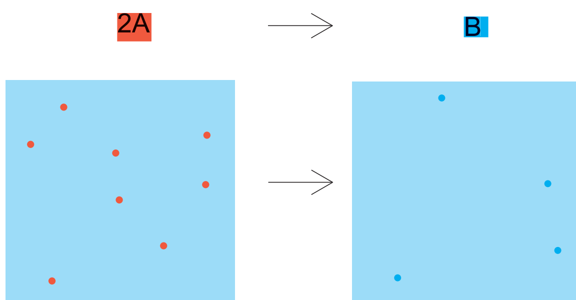
CHEMICAL REACTIONS

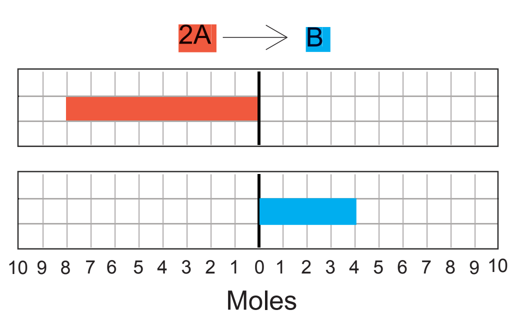
Here are some more practice questions.



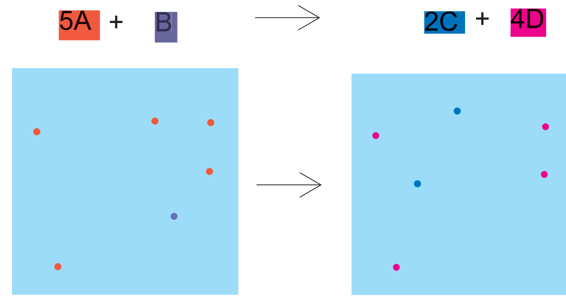


▶ How many moles of A will react to give 6 moles of B? Five moles of B?

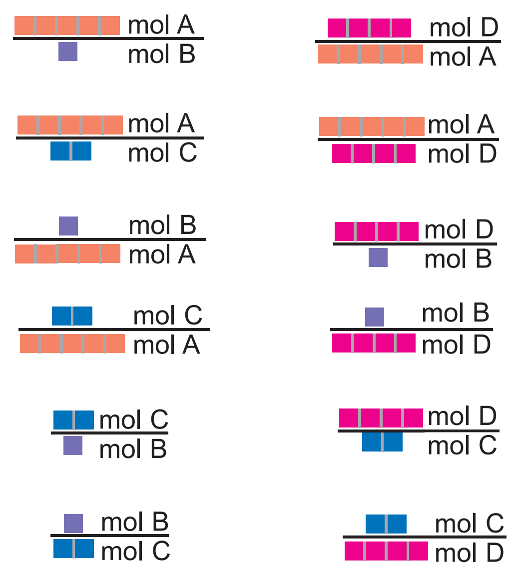


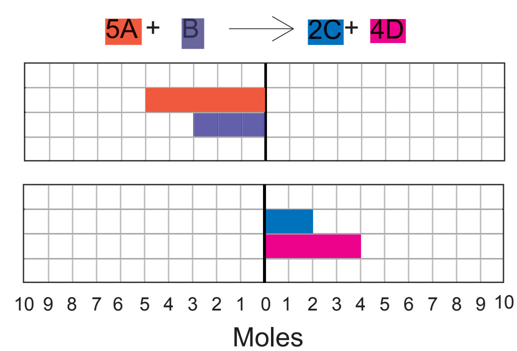


▶ How much of the reactant is left after the reaction?
When 7 moles of A react, how many moles of product will be obtained? Three moles of B?



To help with answering the following questions, all the possible ratios are given.





▶ How much of the reactant(s) is left after the reaction?
When 2 moles of C are formed, how many moles of D are formed?
How many moles of C will form from 10 moles of A and 2 moles of B?
Challenge: How many moles of C will form from 10 moles of A and 1 mole of B?

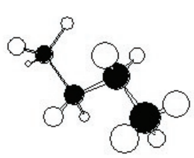
MASS PERCENT

Determining the mass percent composition of a compound refers to the proportion of one element expressed as a percentage of the total mass of the compound.
Knowing the mass percent composition of a compound can help determine environmental effects from that compound. For example, carbon dioxide (CO_2) from burning fossil fuels may contribute to global warming. Methane (CH_4) and butane (C_4H_{10}) are both fossil fuels that, when burned, produce CO_2. Which one will produce less CO_2? Well, it's the one that contains the least amount of carbon as a percentage of the total compound. Let's use mass percent calculations to determine this.

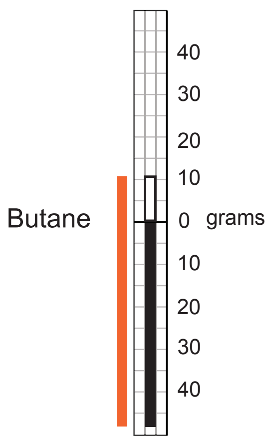
$$\text{Mass Percent of Carbon (C)} = \frac{\text{Mass of Carbon (C)}}{\text{Total Mass of Compound}} \times 100\ \%$$

First, let us determine the mass percent composition of carbon in butatne.

total mass of butane
mass of hydrogen in butane
mass of carbon in butane



Butane
C_4H_{10}



$$\text{Mass Percent of Carbon (C)} = \frac{4 \times 12.011\text{ g C}}{(4 \times 12.011\text{ g C}) + (10 \times 1.008\text{ g H})} \times 100\ \%$$

$$= \frac{48.044\text{ g C}}{48.044\text{ g C} + 10.080\text{ g H}} \times 100\ \%$$

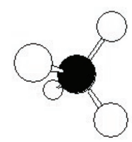
$$= \frac{48.044\text{ g C}}{58.124\text{ g } C_4H_{10}} \times 100\ \%$$

$$= \frac{48.044\text{ g C}}{58.124\text{ g } C_4H_{10}} \times 100\ \%$$

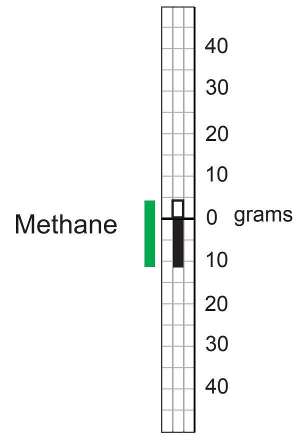
$$= 82.66\ \%$$

Now, let's consider methane.

total mass of methane
mass of hydrogen in methane
mass of carbon in methane



Methane
CH_4



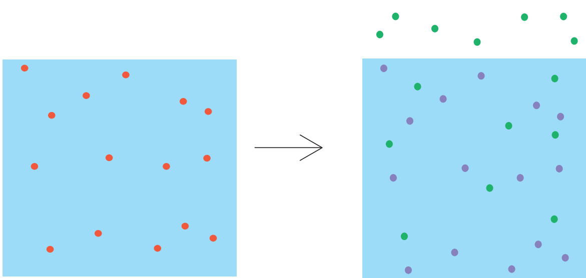
▶ Use the same technique to calculate the mass percent of carbon in methane.

▶ Which compound, methane or butane, contains a higher mass percent of carbon?
Which one will produce more CO_2 when burned?

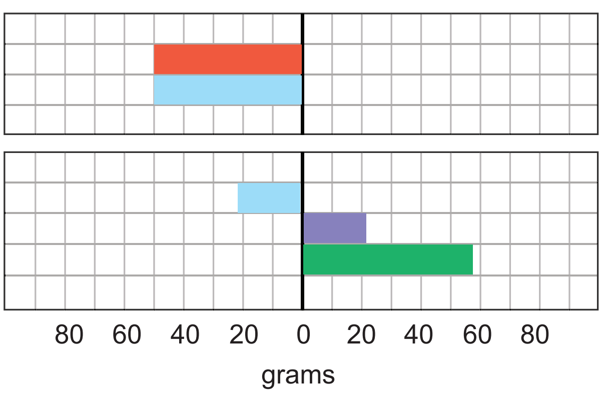
LIMITING REACTANTS

When carrying out a chemical reaction, we may use the exact amount of each reactant needed. Or, we may use an excess of some reactants and a limited amount of others. We may do this if one reactant is very expensive and others are inexpensive so that we can use all of the expensive compound. It can be more cost effective, even if we are wasting money on the excess reactants. The reactant that governs the maximum yield of a product is the limiting reactant.

$CaC_2(s) + 2\ H_2O\ (l) \longrightarrow Ca(OH)_2(aq) + C_2H_2\ (g)$



$CaC_2(s) + 2\ H_2O\ (l) \longrightarrow Ca(OH)_2(aq) + C_2H_2\ (g)$



▶ Which is the limiting reactant when 100 g of water reacts with 100 g of calcium carbide?

First, determine the moles of each reactant that we start with.

$$\text{Moles of } CaC_2(s) = 100\text{ g } CaC_2(s) \times \frac{1\text{ mol } CaC_2(s)}{64.10\text{ g } CaC_2(s)}$$

$$= 1.56\text{ mol } CaC_2(s)$$

$$\text{Moles of } H_2O\ (l) = 100\text{ g } H_2O\ (l) \times \frac{1\text{ mol } H_2O\ (l)}{18.02\text{ g } H_2O\ (l)}$$

$$= 5.55\text{ mol } H_2O\ (l)$$

▶ How many moles of H_2O react with 1 mole CaC_2 in this reaction?

▶ Next, what is the amount of CaC_2 that is needed to react with 100 g of H_2O? Convert moles to grams.

$$5.55\text{ mol } H_2O\ (l) \times \frac{1\text{ mol } CaC_2(s)}{2\text{ mol } H_2O\ (l)} = 2.78\text{ mol } CaC_2\ (s)$$

▶ Compare this answer with what we actually start with.

$$2.78\text{ mol } CaC_2\ (s) > 1.56\text{ mol } CaC_2(s)$$

▶ What is the amount of H_2O that is needed to react with 100 g of CaC_2? Convert moles to grams.

$$1.56\text{ mol } CaC_2(s) \times \frac{2\text{ mol } H_2O\ (l)}{1\text{ mol } CaC_2(s)} = 3.12\text{ mol } H_2O\ (l)$$

▶ Compare this answer with what we actually start with.

$$3.12\text{ mol } H_2O\ (l) < 5.55\text{ mol } H_2O\ (l)$$

Because 3.12 mol H_2O is required and 5.55 mol H_2O is supplied, there is an excess of H_2O. So CaC_2 (s) is the limiting reactant and all of it can react.

▶ Why is it always important to work in the unit of moles when determining limiting reactants? Why not grams or milliliters?

LIMITING REACTANTS

There's another way to determine the limiting reactant using the number of moles of a product that can be made from each reactant.

▶ How many moles of C_2H_2 is formed from one mole of CaC_2? How many moles of C_2H_2 is formed from one mole of H_2O? Use this information to determine the number of moles of product that can be made from our starting quantities.

Moles of C_2H_2 (aq) from CaC_2(s) =

$$1.56 \text{ mol } CaC_2(s) \times \frac{1 \text{ mol } C_2H_2 (aq)}{1 \text{ mol } CaC_2(s)} = 1.56 \text{ mol } C_2H_2$$

Moles of C_2H_2 (aq) from H_2O (l) =

$$5.55 \text{ mol } H_2O (l) \times \frac{1 \text{ mol } C_2H_2 (aq)}{2 \text{ mol } H_2O (l)} = 2.78 \text{ mol } C_2H_2$$

▶ From this calculation, which reactant is the limiting reactant? Why?
Is it the limiting reactant that was determined previously?

▶ Determine which reactant is the limiting reactant when 40 g of magnesium and 20 g of nitrogen react in the following reaction:

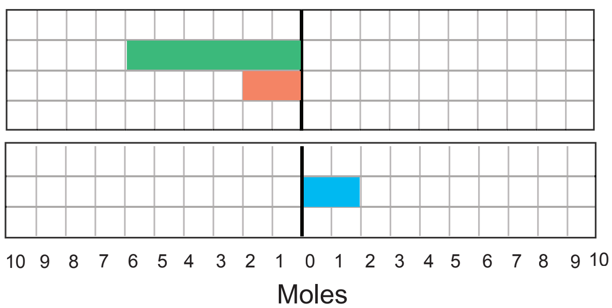
$$3 \text{ Mg (s)} + N_2 \text{ (g)} \longrightarrow Mg_3N_2 \text{ (s)}$$

▶ How many moles of the limiting reactant are consumed by the reaction?
How many grams of the excess reactant are left after the reaction?

YIELDS

The theoretical yield is the maximum product that can be obtained from the amount (mass, moles, volume) of reactant(s) used. Calculate the maximum number of moles of product that can be obtained from the following reaction, when 13.45 g of N_2 reacts with 35 g of Mg.
*Always remember to check for limiting reactants!

$$3\ Mg\ (s) + N_2\ (g) \longrightarrow Mg_3N_2\ (s)$$



The actual yield of a reaction is the amount (moles, volume, mass) of product obtained at the end of the reaction.
The percentage yield can be calculated by:

$$\text{Percentage Yield} = \frac{\text{Actual Yield}}{\text{Theoretical Yield}} \times 100\%$$

▶ In the reaction that forms magnesium nitride, the actual yield was 39.8 g. The theoretical yield was not obtained because the Mg was impure, meaning that when 35 g was weighed out, it was not all Mg (s). What is the percentage yield for this reaction?

▶ Name three reasons that a chemist could obtain a yield less than the theoretical yield.

▶ Determine the theoretical yield for the reaction magnesium nitride reaction when 35 g of Mg reacts with 16 g of N_2.

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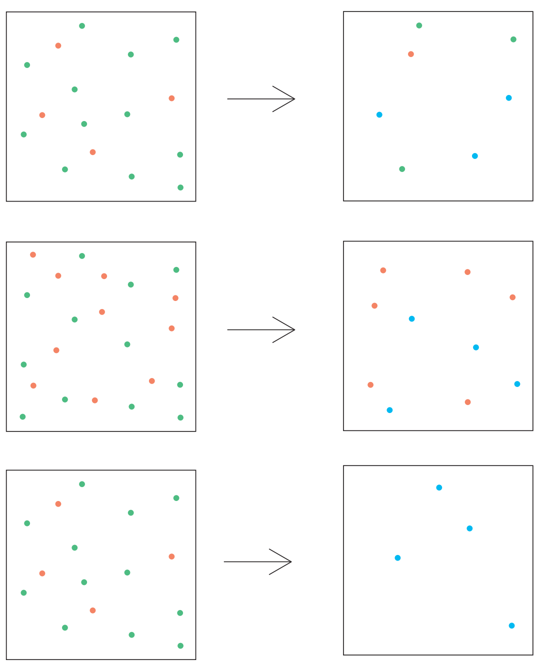
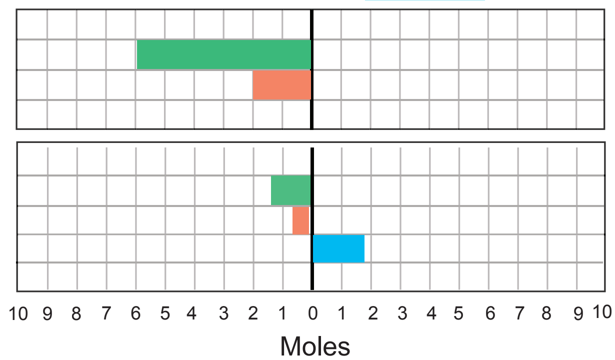
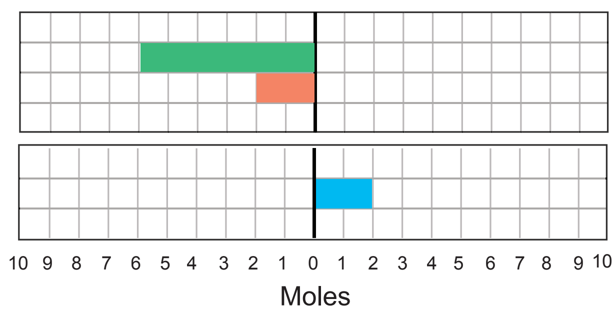
YIELDS

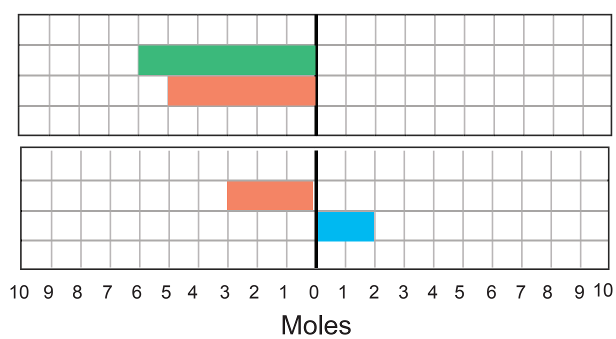
▶ Mark each set of pictures and each ChemLog as either showing a reaction that proceeded to the theoretical yield, the reaction which shows an actual yield, or the reaction which shows there was a limiting reactant.



$3\ Mg\ (s) + N_2\ (g) \rightarrow Mg_3N_2\ (s)$







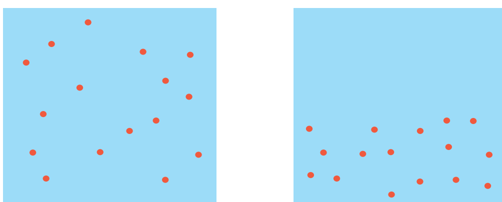
▶ Connect the reaction pictures above with the ChemLog that each refers to.

MOLARITY

A solution is a homogeneous (uniform in composition) mixture of two chemicals. The solute in a solution is the substance that is being dissolved. The solvent in a solution is the substance doing the dissolving.

Supposed we poured sugar into water to form a solution.

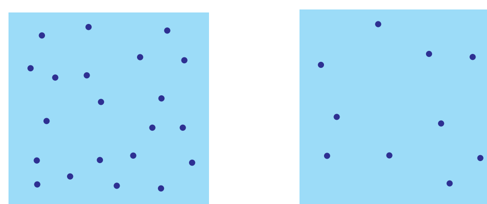
▶ Which of the following pictures would be an accurate representation of the sugar solution? Circle the solution

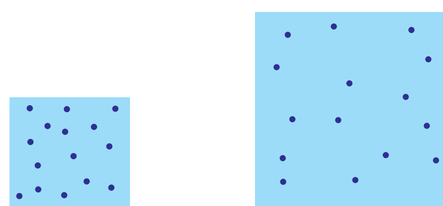


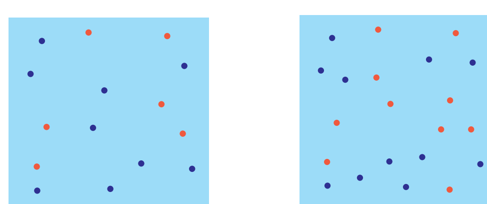
Chemists talk about the concentrations of solutions. The concentration of solution is the amount of solute per solvent. One way to state the concentration of a liquid solution is to state its molarity. Molarity is defined as the number of moles of solvent per liter of solvent.

$$\text{Molarity (M)} = \frac{\text{moles of solute}}{\text{liters of solvent}}$$

▶ Of the following pairs of pictures. Circle the picture which is most concentrated (the picture with the highest molarity).



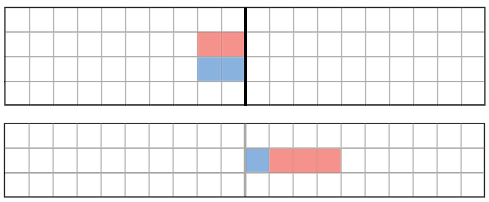




BALANCING CHEMICAL REACTIONS

The numbers to the left of entire chemical formulas in a reaction are called the stoichiometric coefficients. A coefficient of one, as seen in the reaction below, is not written explicity but is implied. When the number of atoms of each element on each side of the arrow are the same, the reaction is said to be balanced.

$N_2 + H_2 \rightarrow NH_3$ 100 C



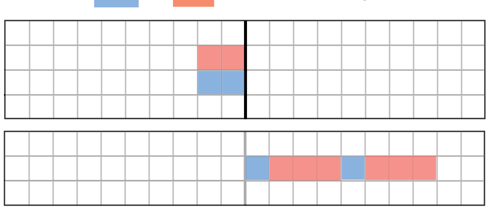
▶ Each block in ChemLog desgnates an element. Count the number of red elements before the reaction. Count the number of red elements after the reaction. Count the number of blue elements before the reaction. Count the number of blue elements after the reaction. Compare the number of red elements before and after the reaction. Compare the number of blue elements before and after the reaction.

Before Reaction	After Reaction
2	1
2	3

In order to balance the chemical reaction, we must make the number of blue elements before the reaction equal to the number of blue elements after the reaction. The same must be done for the red elements. Lets first begin with the products of the reaction. In order to make the number of blue blocks before and after the reaction equal, we need to double the blue blocks after the reaction. In doing that, we double the red blocks too.

▶ Notice that the amount of product has been doubled. Complete the following table using the ChemLog below.

$N_2 + H_2 \rightarrow 2\ NH_3$ 100 C



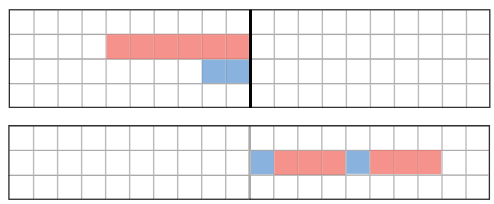
Before Reaction		After Reaction
2	=	2 x 1
2		

$N_2 + H_2 \rightarrow 2\ NH_3$

Notice the number in front of the product; this is called a stoichiometric coefficient. Since we doubled the number amount of product, we must show this in the chemical reaction using coefficients.

▶ The number of blue elements before and after the reaction is now equal. What can we do to make the number of red elementss equal before and after the reaction?
Complete the following table using the ChemLog below.

$N_2 + 3\ H_2 \rightarrow 2\ NH_3$ 100 C



Before Reaction		After Reaction
2	=	2
2 x ____	=	6

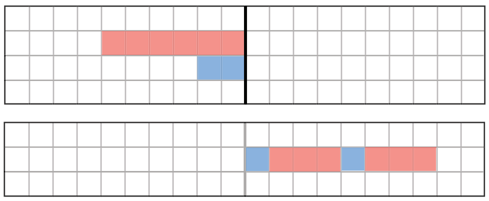
$N_2 + 3\ H_2 \rightarrow 2\ NH_3$

Notice that the coefficients are the numbers we multiplied by in order to gain the same number of blue boxes before and after the reaction and the same number of red boxes before and after the reaction.

MOLE-TO-MOLE CALCULATIONS

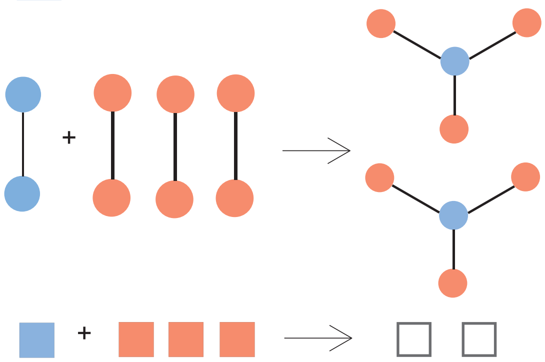
Using the chemical reaction we balanced previously, let's do some simple calculations with that information.

$N_2 + 3\,H_2 \rightarrow 2\,NH_3$



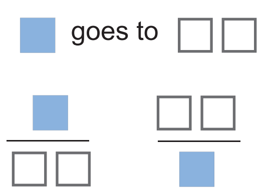
Here are a few ways to think about and visualize this reaction.

$N_2 \quad + \quad 3\,H_2 \quad \rightarrow \quad 2\,NH_3$



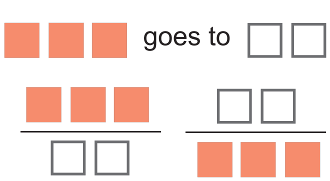
▶ How many moles of product can come from one mole of N_2?

To answer, just look at the fact give in the reaction. One mole of N_2 (along with something not important to answer the question) goes to two moles of NH_3, the product.



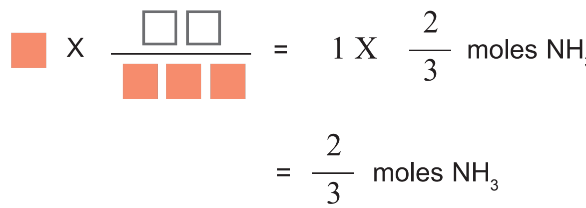
So, 2 moles of product can come from 1 mole of N_2.

▶ How many moles of product can come from three moles of H_2?



▶ How many moles of product can come from one mole of H_2?

Since this question is more difficult and cannot be answered just be looking at the reaction, we'll use some simple math to figure it out.



Hint: Remember that the ratio we use has the number of moles of product (the thing we need to answer the question) on the top.

▶ How many moles of NH_3 can be made from four moles of N_2?

▶ Why must moles be used instead of molecules when answering this question? Try to figure out the number of molecules of product you would make when starting with one molecule of H_2. Explain your answer.

AVOGADRO'S NUMBER

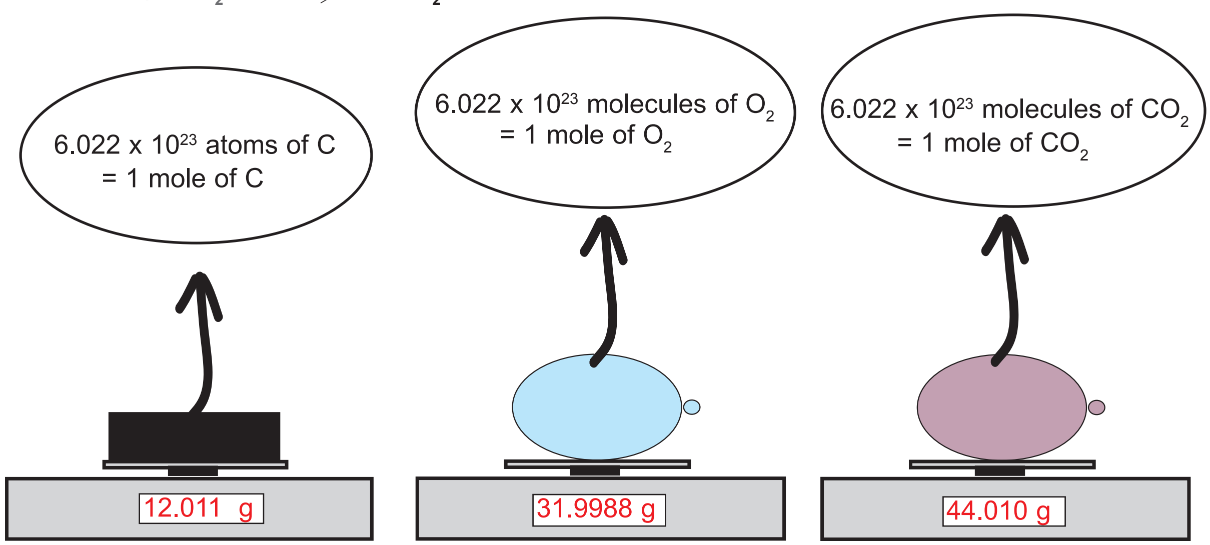
Avogadro's number is the number of particles of substance in a mole. Just as 1 dozen means 12 of something, regardless of what it is - eggs, diamonds, molecules, 1 mole means 6.022×10^{23} of something. There are Avogardo's number, 6.022×10^{23}, of particles in every mole. That's a lot!

▶ How many atoms are in a mole of atoms?
How many molecules are in a mole of molecules?

▶ Since there are the same number of items in a mole, does a mole have a specific mass, say 5 grams? Why or why not?

Let's look at the following reaction of carbon oxygen to form carbon dioxide.

$$C + O_2 \longrightarrow CO_2$$



C	+	O_2	$\longrightarrow$	CO_2
12.011 g	+	31.9988 g	=	44.010 g

▶ What is the mass of 2 moles of C?
What is the mass of 0.5 moles of CO_2?
What is the mass of 1 mole of ozone, O_3?

You may be wondering, how scientists came up with Avogadro's number - 6.022×10^{23} isn't a number one normally thinks of off the top of one's head! Well, Amadeo Avogardro was not the first scientist who realized this number. However, he was the first scientist to sense the significance of the mole, so the number is named after him. Technically, a mole is an amount of substance that contains as many elementary entities as there are atoms in exactly 12 g of the carbon - 12 isotope.

Avogadro's Numbert = $N_A = 6.022 \times 10^{23}$