Mole Concept and Stoichiometry

The mole concept and the closely related concept of stoichiometry are the hardest concepts for most students to fully understand.

##### The Mole

The “mole” is the concept of measuring the amount of a substance and it is a fundamental unit in the système international d'unités (International System of Units – the modern form of the metric system). It is defined as the ***number of representative particles in exactly 12.0 grams of C-12***. So by this definition, the mass of one mole of carbon-12 must be exactly 12.0 grams.

**Representative Particles**

The term “representative particles” is used to describe the particles or entities that represent the smallest unit that has all of the properties of that “group” – examples would be:

|  |  |
| --- | --- |
| Large Group | Representative particle |
| geese | goose |
| mice | mouse |
| people | person |
| VW Beetle  | a single car |

When counting VW Beetles, you do not count wheels, or windows, because they do not have the same “properties” of a whole VW Beetle. You must count a whole car, same with counting geese, you do not count feathers or beaks – you must count the whole bird (and that is a goose).

For chemistry, there are three representative particles that we normally use in mole conversions and one other unit that we use in electrochemistry and oxidation-reduction problems. Again, the representative particles depend on what is being “counted” and you have to be absolutely sure as to the term used – it helps with the AP questions.

* + for elements the representative particles 🡺 atoms
	+ for ionic compounds the representative particles 🡺 formula units
	+ for covalent compounds (also called molecular compounds) the representative particles 🡺 molecules

When dealing with electrochemistry and oxidation-reduction problems, the moles of electrons are used in the calculations; the representative particle for electrons is (drum roll please) an electron.

|  |
| --- |
| **Representative particles for various substances** |
| **Substance** | **Representative particle** |
| Iron | Atom |
| Water | Molecule |
| Table salt | Formula unit (and a formula unit is one grouping of NaCl) |
| Oxygen gas | Molecule (remember that oxygen is diatomic and is O2, therefore must be a molecule and not an atom. |

The number of representative particles in exactly 12.0 grams of carbon-12 (in this case, the representative particles are atoms, since carbon is an element) is 6.0221415 x 1023 (on the AP Chemistry formula charts, it is written as 6.022 x 1023). That number is referred to as **Avogadro’s Number (*NA*)** in honor of the nineteenth century Italian scientist [Amedeo Avogadro](http://en.wikipedia.org/wiki/Amedeo_Avogadro). It was Avogadro, in the early 1800’s that first proposed that the volume of a gas is proportional to the number of particles of gas (either [atoms](http://en.wikipedia.org/wiki/Atom) or [molecules](http://en.wikipedia.org/wiki/Molecule) – again that idea of representative particles and what you are counting). Since gas volume changes with temperature and pressure, the gases have to be at the same pressure and temperature for the comparison to be made.

**Molar Mass**

There are two definitions that are closely related in chemistry, atomic mass units and that of a mole. They are related so that the following connections can be made. The two definitions are:

* + Atomic mass unit is exactly 1/12th the mass of a carbon-12 nucleus.
	+ A mole is the number of particles in exactly 12.0 grams of C-12.

**If you want to see how they are related, read the table below. If not, skip below the table to the paragraph at the bottom of this page.** Let’s see how they are related.

|  |  |
| --- | --- |
| **Atomic Mass Units** | **Moles** |
| Definition: An atomic mass unit is exactly 1/12th the mass of a carbon-12 nucleus | Definition: A mole is the number of particles in exactly 12.0 grams of C-12 |
| An atom of carbon-12 has:* 6 protons
* 6 neutrons
* 6 electrons
 | An atom of carbon-12 has:* 6 protons
* 6 neutrons
* 6 electrons
 |
|  | A mole of carbon-12 has:* 6 moles protons
* 6 moles neutrons
* 6 moles electrons
 |
| The atomic mass of C-12 is 12 amu. | What is the mass of one mole of C-12? By the definition, one mole of carbon-12 has to have a mass of 12 grams. |
| Atomic mass = #p+ + #no (since electrons have relatively no mass) | Molar mass = #p+ + #no (since electrons have relatively no mass) |
| So: 12 amu = 6p+ + 6no | So: 12 grams = 6 moles p+ + 6 moles no |
| Since the mass of a proton and the mass of a neutron are the same, we can substitute protons for neutrons (or vice versa) in the equation above and get: | Since the mass of a proton and the mass of a neutron are the same, we can substitute protons for neutrons (or vice versa) in the equation above and get: |
| 12 amu = 6p+ + 6p+ | 12 grams = 6 moles p+ + 6 moles p+ |
|  | The mass of a proton is 1 amu, so |
|  | 12 grams = 6 mol (1 amu) + 6 mol (1 amu) |
|  | 12 grams = 12 moles of amu’s |
| **1 amu = 1 p+** | **1 gram = 1 mole of amu’s** |

So what does this all mean? The definitions of atomic mass units (amu) and moles are related, so as to allow the periodic table to show both the average atomic mass and the molar mass of every element. The average atomic mass of chlorine is 35.45 amu, while the molar mass is 35.45 grams. Everything that is done with isotopes and atomic masses, applies to moles and molar masses.

The definition of molar mass is the mass of one mole of a substance.

* + Units are g/mol
	+ Numerically equal to the average atomic weight in atomic mass units.
	+ Specific names depend on the representative particles.
		- atoms = gram atomic mass
		- ionic compounds = gram formula mass
		- covalent compounds = gram molecular mass

It is common for AP free response questions to ask for the students to find molar mass of an unknown compound. This is especially true with the lab based question.

|  |  |  |
| --- | --- | --- |
| **Year** | **How was problem stated?** | **What does this tell you?** |
| 1983 | The molecular weight of a monoprotic acid HX was to be determined. | Molecular weight = covalent compound |
| 1978 | Calculate the molecular weight of the acid HA. | Molecular weight = covalent compound |
| 1989 | In an experiment to determine the molecular weight and the ionization constant for ascorbic acid (vitamin C) | Molecular weight = covalent compound |
| 1979 | In a laboratory determination of the atomic weight | Atomic weight = element |
| 1973 | What minimum data are needed to determine the molecular weight | Molecular weight = covalent compound |

**Molar Volume**

As stated above, it was Avogadro which first proposed that the volume of a gas is proportional to the number of particles of gas at constant temperature and pressure. Since gas volume changes with temperature and pressure, the pressure and temperature have to be stated for the comparison to be made. The temperature and pressure that has been agreed upon is called **S**tandard **T**emperature and **P**ressure (STP). Standard temperature is 0OC (273.15 K) and 1atmosphere of pressure (1 atm). ***If and only if you are at those conditions can you use the value for molar volume!*** Molar volume is the volume occupied by one mole of **ANY** gas at STP; that volume is 22.4 liters.

Avogadro used his law with equal volumes of gases under the same conditions and equal numbers of molecules to determine how many particles were in a mole.

**Students often see “volume” or “liters” in a problem and jump to using 22.4 liters! You must have two things in order to use molar volume.**

**Molarity and molar does not mean a gas! You cannot use 22.4 liters in the problem!**

1. **Must have a gas!**
2. **Must be at STP!**

**Mole Diagram**

The mole diagram is designed to help students visualize how to do the conversions. If you are asked to go from grams to moles, that is a one step conversion – going from the “mass in grams” to the “moles” in the middle. If you are asked to convert liters @STP to atoms of He, then that is a two step problem: first is to convert from liters (at the top) to moles in the middle, then convert from the moles to the representative particles at the bottom.

**Examples**

**Find the number of moles in 35.78 grams of (NH4)2SO4.**

 To do this problem, you need to first find the molar mass of (NH4)2SO4:

|  |  |  |  |
| --- | --- | --- | --- |
| **Element in the formula** | **Number of moles of the element in the formula** | **Molar Mass (periodic table)** | **Final value** |
| Nitrogen | 2 | 14.01 | 28.02 |
| Hydrogen | 8 | 1.01 | 8.08 |
| Sulfur | 1 | 32.07 | 32.07 |
| Oxygen | 4 | 16 | 64 |
|  |  | Add them up | 132.17 grams |

**Find the volume of 4.65 x 1024 atoms of He @ STP.**

**Find the number of representative particles in 2.76 moles of HCl.**

**Find the moles of 38.9 grams of oxygen gas.**

**Hardest problem: Find the number of atoms of oxygen in 343 grams of ammonium oxalate.**

**Determination of the formula of a compound**

Whenever you need to break up a formula into smaller parts, always place the number “1” in front of the most complicated formula and then determine the other coefficient.

**Percent Composition**



Percent composition is a very useful tool to determine the amount of an element in a compound. Since it is based on the total mass (and we use the molar mass for the total mass) the percent composition of an element in a compound is independent of the amount of that compound.

**Example: Find the percent composition for nitrogen in (NH4)2SO4.**

 To do this problem, you need to first find the molar mass of (NH4)2SO4:

|  |  |  |  |
| --- | --- | --- | --- |
| **Element in the formula** | **Number of moles of the element in the formula** | **Molar Mass (periodic table)** | **Final value** |
| Nitrogen | 2 | 14.01 | 28.02 |
| Hydrogen | 8 | 1.01 | 8.08 |
| Sulfur | 1 | 32.07 | 32.07 |
| Oxygen | 4 | 16 | 64 |
|  |  | Add them up | 132.17 grams |



If the problem asked for the amount of grams of nitrogen in 402 grams of (NH4)2SO4 you can do that problem two ways. The first way is using mole conversions.



This problem can also be done using percent composition. 21.20% of (NH4)2SO4 is nitrogen, so 21.20% of 402 grams of (NH4)2SO4 must give you grams of nitrogen.

21.20% x 402 grams (NH4)2SO4 = 0.2120 x 402 = 85.2 grams

Notice that the numbers are exactly the same if you look at how the 21.20% was calculated:



Look at the mole conversion, all of the numbers are located in the exact same place, 402, 2 and 14.01 on the top and the 132.17 on the bottom.

**Empirical and Molecular Formulas**

There are two formulas that are asked for on the AP exam, empirical and molecular formulas. They are related to each other and sometimes the problems ask for one or the other or both. There are two ways to determine the molecular formulas and only one way to get the empirical formula. Since a formula, such as (NH4)2SO4 relates the number of moles of each element present, the molecular and empirical formulas deal with the moles of each element in the compound.

* **empirical formula**
	+ the lowest whole number ratio of atoms in a formula
	+ generally used to determine ratio of elements in a compound
	+ can be used to determine the molecular formula
	+ two common problems for determining the data needed to find the empirical formula
		- from percent composition
		- combustion analysis
* **Molecular formula**
	+ the formula showing the actual ratio of atoms in a formula
	+ usually found from the empirical formula
	+ the molar mass of the actual compound is usually determined by laboratory data
		- Find the grams of the compound (given to you in the problem) and use the other information in the problem to do stoichiometry to determine the moles of the compound.
		- Molar mass is grams per mole so take the grams and divide by the moles.
		- There are three common methods in the laboratory to determine the molar mass:
			* colligative properties
			* titration (usually acid-base)
			* gas laws
* **Finding empirical and molecular formulas:**
	+ - find empirical formulas
			* if given grams of each element, start at step three of the percentage problems
			* for percentage problems
				+ step one: assume a 100 gram sample
				+ step two: for each element multiply the percentage by the mass (100g) to give you grams
				+ step three: convert all grams into moles
				+ step four: divide by smallest number of moles
				+ step five: clear fractions by multiplication each moles by the denominator of the fraction (if needed)
		- find molecular formulas
			* If given the molecular formula’s molar mass
				+ Use the empirical formula’s molar mass and see how many times it goes into the molecular formula’s molar mass. Must be a whole number of times (2x, 3x etc.)
				+ Can use the molecular formula’s molar mass to find molecular formula directly.
			* May have to determine the molecular formula’s molar mass from laboratory data

There are two common ways to find the molecular formula…find it directly or use the empirical formula. Many times on the AP free response questions, you will have to solve a problem to determine the molecular formula’s molar mass. If you want to solve this first and then find the molecular formula that is fine or you can find the empirical formula and then find the molecular formula.

**Examples from AP Free Response Questions**

**Problem involving combustion analysis:**

 (a) A 0.7549 g sample of the compound burns in O2*(g)* to produce 1.9061 g of CO2*(g)* and 0.3370 g of H2O*(g)*.

 (i) Calculate the individual masses of C, H, and O in the 0.7549 g sample.

 (ii) Determine the empirical formula for the com­pound.

**Answer:**

You start out with an equation: CxHyOz + O2 🡪 CO2 + H2O

Our job is to find the x, y and z. Using the Law of Conservation of Mass, we know that all of the carbon in CO2 came from the original compound and that all of the H in H2O came from the original compound. So we need to find the C in the CO2 and the H in the H2O and we can do this by either mole conversions or by percent composition. As good review, we are going to find the amount of H by mole conversion and the amount of C by percent composition.

(a i) 

 

 0.7549 g sample – (0.5198 g C + 0.0374 g H) = 0.1977 g of Oxygen

(a ii)  

Divide by the smallest number of moles

Since 3.5 moles of C is not a whole number, take the denominator of the fraction (0.5 = ½) and multiple each moles by 2 and get 7 C : 6 H : 2 O, thus an empirical formula of C7H6O2. Let’s say that the problem said that the molecular formula’s molar mass was determined to be 244 g/mol (more likely is that you are going to have to solve for this). The empirical formula has to be a “reduced” version of the molecular formula, so we can go back and put in the factor that was taken out to get the “reduced” empirical formula.

C7H6O2 = EF = 122 g/mol

First find the empirical formula’s molar mass (in this case 122 g/mol).

X 2

X 2

Next, how many times does the EF’s mass go into the MF’s mass?

Answer here is two times. So multiple the EF by two for the number of atoms and get the molecular formula to be C14H12O4.

C?H?O? = EF = 244 g/mol

**Problem involving percentages:**

An unknown compound contains only the three elements C, H, and O. A pure sample of the compound is analyzed and found to be 65.60 percent C and 9.44 percent H by mass. Determine the empirical formula of the compound.

**Answer:**

All three percentages have to add up to be a 100, so 65.60 %C + 9.44 %H + *X* %O = 100, so 24.96% O.

Step one: Assume a 100 gram sample.

Step two: Then 65.60% of 100 grams = 65.60 grams of C (likewise you have 24.96 g O and 9.44 g H)

Step Three: Convert to moles



 Step Four: divide by smallest number of moles



 Step Five: clear fractions by multiplication of denominator of fraction (if needed)

 Empirical formula = C7H12O2

**Finding the molecular formula**

We are going to use the information above to illustrate the two methods of finding the molecular formula. Again, we will assume that the problem said that the molecular formula’s molar mass was determined to be 256 g/mol. So to find it from the empirical formula:

C7H12O2 = EF = 128 g/mol

First find the empirical formula’s molar mass (in this case 128 g/mol).

X 2

X 2

Next, how many times does the EF’s mass go into the MF’s mass?

Answer here is two times. So multiple the EF by two for the number of atoms and get the molecular formula to be C14H24O4.

C?H?O? = EF = 256 g/mol

**NOTE: If the EF’s mass does not go evenly into the MF’s mass, you made a mistake.**

To find the molecular formula directly you use the same process you did for finding the empirical formula. However, to find the molecular formula directly, you will have to do more work at the beginning of the problem to get the grams of each element, but there will be no work at the end of the problem: you will not get a fraction and you do not have to divide by the smallest number.

The original problem for the last example was a four part problem…only the first part was shown above. Part (a) and part (b) are below.

An unknown compound contains only the three elements C, H, and O. A pure sample of the compound is analyzed and found to be 65.60 percent C and 9.44 percent H by mass.

(a) Determine the empirical formula of the compound.

(b) A solution of 1.570 grams of the compound in 16.08 grams of camphor is observed to freeze at a temperature 15.2 Celsius degrees below the normal freezing point of pure camphor. Determine the molar mass and apparent molecular formula of the compound. (The molal freezing-point depression constant, *Kâ*, for camphor is 40.0 kg•K•mol-1.)

We have already done the work for part (a) above and shown how to find part (b) from part (a) by given you the molar mass that would be found from part (b).

However, on the AP exam, you can answer the question in any order you would like, so the second method to finding the molecular formula involves doing part (b) first (finding the molar mass), then using the percentages to find the molecular formula and finally using that to find part (a).

Again, the answer for the molar mass from part (b) is 256 grams/mol.

We are going to assume that we have a one mole sample of this compound – that means that we have a sample mass of 256 grams. So using that information, we need to find the amount of each element in that compound.

65.60% of 256 grams = 167.9g C 9.44% of 256 grams = 24.2g H 24.96% of 256 grams = 63.9g O

Now convert the grams into moles.

}



This will give you the molecular formula automatically, so there is no dividing by the smallest number. The 13.98 moles of C is 14 moles, 23.96 moles H becomes 24 moles and the 3.99 moles of O becomes 4 moles. So the molecular formula is C14H24O4.

Once you have the molecular formula, C14H24O4, just reduce it to lowest terms and get the empirical formula. Since the numbers are 14, 24 and 4, we have a common factor of 2, so they would become 7, 12 and 2 and the empirical formula becomes C7H12O2.

**HYDRATES**

**A hydrate** is a compound that has water attached to it. It is common for the AP exam to ask questions about hydrates, usually it is to determine the amount of water that the hydrate has. It is nothing more than an empirical formula problem. A dot is used to state how many water molecules are attached (but not chemically combined) to the crystal. A formula might look like CuSO45H2O.

* Terms
	+ How to name them
		- uses prefixes to state number of water then put it in the form of compound’ name followed by the prefix for the amount of water attached to the word “hydrate”
		- copper(II) sulfate pentahydrate
	+ **anhydrous** – *without* water
	+ **water of hydration** – amount of water attached to the hydrate

Usually the problem will give enough information to determine the grams of the water of hydration and the grams of the anhydrous. Once you have the grams of each one, follow the steps to find the empirical formulas. The only difference is that you will be using formulas and not elements.

**Example:**

When 1.357 grams ofCoCl2*x* H2O is heated, 1.062 grams of cobalt(II) chloride is left behind. Find the value for “x”.

1.357 grams of CoCl2*x* H2O – 1.062 grams CoCl2= 0.295 grams H2O.

** **

Divide by the smallest number of moles

Empirical formula is CoCl22 H2O

##  STOICHIOMETRY

Stoichiometry is the study of the quantitative relationship between the amount of reactants and products. There is a simple graphical approach to setting up the information in the problem and planning your approach to solving the problem. First you must start with a balanced chemical equation.

Graphical approach to solving stoichiometry problems

Remember that the coefficients of a balanced equation relate the number of moles of each substance that reacts together and the moles of the products that are formed. It is important that you **MUST** start with a balanced equation.

* Place the information in the problem.
	+ Balance equation on first line.
	+ Moles go on the second line.
	+ Anything that can be changed into moles on the third line.
		- Grams
		- molarity
		- molality
		- representative particles
		- liters @ STP
	+ Use a question mark (?) to identify the unknown.
* Plan how to solve the problem.
	+ Use arrows to plan to do the problem and to guide you in your conversions.
	+ Draw from you starting information
	+ Must be in moles to go from side to side in a balance equation.
	+ Draw arrows down if need be.

**Balanced equation here**

2 H2 + O2 🡪 2 H2O

Moles are placed on this line.

Anything that can be changed into moles goes here.

**A typical problem should look like this:**

Problem: When 150.0 grams of TiO2 is reacted with BrF3, find the number of liters of oxygen gas at STP, the number of molecules of bromine, the number of moles of titanium(IV) fluoride and the grams of BrF3 required for complete reaction.

Your setup should look like this:

TiO2 + BrF3 🡪 TiF4 + Br2 + O2

 ?

150.0 g ? g ? ?

 molecules L@STP

Draw the arrows to help you with the problem, we will draw only one set at a time, however, on your problems, you can draw all of the sets at the same time. The only time that you can go side to side is when you are in moles. A balanced chemical equation relates the moles of each species in the equation. **DO NOT FORGET TO BALANCE THE EQUATION.**

3 TiO2 + 4 BrF3 🡪 3 TiF4 + 2 Br2 + 3 O2

 ?

150.0 g ? g ? ?

 molecules L@STP

Anytime that you are going up or down in the diagram, the number in front of moles will ALWAYS be one. You are using the three definitions of (i) mole, (ii) molar mass and (iii) molar volume. Since all three of these definitions are based on one mole, the number in front of moles has to be one.

Your problem would be solved like this:



3 TiO2 + 4 BrF3 🡪 3 TiF4 + 2 Br2 + 3 O2

 ?

150.0 g ? g ? ?

 molecules L@STP



3 TiO2 + 4 BrF3 🡪 3 TiF4 + 2 Br2 + 3 O2

 ?

150.0 g ? g ? ?

 molecules L@STP



3 TiO2 + 4 BrF3 🡪 3 TiF4 + 2 Br2 + 3 O2

 ?

150.0 g ? g ? ?

 molecules L@STP



**Practice Problem #1**

How many grams of water can be made from 6.0 grams of hydrogen, and how many moles of oxygen are needed for the reaction to go to completion?

H2 + O2 🡪 H2O

## Now…set up the problem and then solve the problem – without a calculator (the numbers are easy!).LIMITING REAGENTS

* limiting reagent
	+ The limiting reagent determines or limits the amount of product formed.
	+ At least two amounts are given in the problem (may be more than two, but on the AP test, it has only been two – does not mean that it could not be more than two on the AP exam).
	+ Once identified, ALL calculations MUST be based on this reactant.
* excess reagent
	+ The reagent that is left over…have more than enough of it.
	+ Uses up all of the limiting reagent.
* Easiest way to determine this type of problem is to change each starting amount to moles (or grams – depends on what the problem is asking) of the product that you are looking for.
	+ Smaller amount comes from the LIMITING REAGENT and is your answer.
	+ **MAKE SURE THAT YOU ANSWER THE QUESTION CORRECTLY…THE PRODUCT IS NOT THE LIMITING REAGENT.**
	+ Read the problem before you start and see what product they are asking about…and then go find that product.

You can always tell this type of problem by setting up the problems with the three lines. Once you have two starting amounts on the diagram…you know that it is a limiting reagent problem – see example below.

4 Fe*(s)* + 3O2*(g)* 🡪 2 Fe2O3*(s)*

 42 g 75 g ? g

The two starting amounts tell you that it is a limiting reagent problem. To solve this problem, talk both starting amounts and find the grams of the Fe2O3. The one with the smallest answer will be the correct value for the amount of Fe2O3.



Since the 60. grams is smaller than the 250 grams, the 60. grams ofFe2O3 is the amount that is produced. Since the 60. grams of Fe2O3 is produced from the 42 grams of Fe, the Fe must be the limiting reagent and the 75 grams of oxygen gas must be the excess reagent.

**Part of an AP question – Practice Problem #2**

CH4*(g)* + 2 Cl2*(g)* → CH2Cl2*(g)* + 2 HCl*(g)*

Methane gas reacts with chlorine gas to form dichloromethane and hydrogen chloride, as represented by the equation above.

(a) A 25.0 g sample of methane is placed in a reaction vessel containing 2.58 mol of Cl2*(g)*.

(i) Identify the limiting reactant when the methane and chlorine gases are combined. Justify your answer with a calculation.

(ii) Calculate the total number of moles of CH2Cl2*(g)* in the container after the limiting reactant has been totally consumed.

CH4 + 2 Cl2 → CH2Cl2 + 2 HCl

2.58 ?????

 25.0 g

**Part of an AP question – Practice Problem #3**

**NOTE:** You can so the questions in any order that you want so this approach answer part (ii) first and then you answer part (i).

4 Fe*(s)* + 3O2*(g)* 🡪 2 Fe2O3*(s)*

Iron reacts with oxygen to produce iron(III) oxide as represented above. A 75.0 g sample of Fe*(s)* is mixed with 40.0 grams of oxygen gas.

(a) Calculate the number of moles of each of the fol­lowing before the reaction occurs.

 (i) Fe*(s)*

(ii) O2*(g)*

(b) Identify the limiting reactant when the mixture is heated to produce Fe2O3. Support your answer with calculations.

(c) Calculate the number of moles of Fe2O3 produced when the reaction proceeds to completion.

**Percent Yield**

Percent yield is a measurement of the efficiency of a chemical reaction.

* theoretical yield
	+ This is what you calculate from stoichiometry.
	+ This is fantasy land! Most reactions will not give you 100% of the product that you expect.
* actual yield
	+ What you get in the experiment or laboratory investigation (what happens in the real world).
	+ Remember, no chemical process is 100% efficient.
		- Therefore no yield will be 100% (useful for lab reports!)
		- if more than 100% 🡺 contamination of some kind

Percent yield = actual yield (real) x 100

 theoretical yield (ideal)

You can use a magic circle if it helps you.

**A Y**

**TY**

**%Y**

Laboratory free response questions on the AP exam may give you the information in a data table. The data table should give you information that will allow you to determine the actual results and the starting amount of the limiting reactant – usually in the lab it is clearly stated which reagent is the limiting reagent. From that information, you can do the stoichiometry to determine the theoretical yield (the maximum amount possible) and then determine the percent yield.

**Reagent versus reactant**

Although most people consider them interchangeable (and in most cases they are) there is a very subtle difference between the two. The AP exam does not care which one you use.

**Reactant:** Any substance that participates in a chemical reaction, especially one that is present at the initiation of that reaction.

**Reagent:** A chemical that generally is added to a substance to produce a specific chemical change or to detect a specific chemical.

**Answers to Practice Problems**

**Problem 1**

Equation is: 2 H2 + O2 🡪 2 H2O



**Problem 2**



Since 1.29 moles is the smallest answer, that is the amount of CH2Cl2 that is formed and the limiting reagent has to be the Cl2.

**Problem 3**

(a) 

(b and c) 

Since the 0.833 moles of Fe2O3 is the smallest amount – that must be the amount produced [answer to part (c)]. Since the O2 produces that amount, the O2 must be the limiting reagent [the answer to part (b)].

**Problems**

Directions: Put all of your work on separate sheets of paper. DO NOT DO YOUR WORK ON THESE PAGES.

Find the number of moles in each of the following:

1) 1.204 x 1024 atoms of He

2) 24.6 grams of Sr

3) 5004 grams of Na2CO3

4) 456 liters of O2 (g) @ STP

Find the number of grams in each of the following

5) 4.57 moles of NH­3

6) 2.35 moles of FeCO3

7) 2.20 moles of Sn

Find the number of representative particles in each of the following

8) 3.45 moles of Ar

9) 23.4 moles of Fe2(CO3)3

10) 6.35 grams of Cu

11) 34.5 liters of H2O(g) @STP

Make the following conversions:

12) 5.75 grams of H2 gas to liters @ STP

13) 6.35 moles of oxygen gas to grams

14) 2.35 x 1027 atoms of Np to grams

15) 3.54 x 1022 molecules of H2O to grams

16) 45.7 liters of H2 (g) @ STP to atoms of H (be careful on this one!!)

17) 4.5 moles of Ne to grams

18) 6.75 x 1026 molecules of CH4 (g) to liters @ STP

19) 245 grams of SrCO3 to formula units

20) 24.2 liters of chlorine gas to grams @ STP

21) Calculate the empirical formula if a compound is 27.59% C, 1.15% H, 16.09% N, and 55.17% O.

22) Find the empirical and molecular formulas of an unknown compound. Its percent composition is 58.8% C, 9.8% H, and 31.4% oxygen. The molecular compound’s molar mass is 306 grams/mol.

23) What is the molar mass of Fe2(CO3)3?

24) What is the percent composition of O in Fe2(CO3)3?

25) What is the grams of oxygen in 867 grams of Fe2(CO3)3?

26) Convert 23.4 moles of Fe2(CO3)3 to grams.

27) Convert 437 grams of Fe2(CO3)3 to atoms of carbon.

28) Styrene is a compound contains only C and H. If 0.438 grams of styrene is burned in oxygen, 1.481 grams of CO2 and 0.303 grams of water are produced, what is the E.F.

29) A hydrate of magnesium iodide has the formula MgI2 • *X* H2O. To determine a value for *X* , the hydrate sample is heated until all the water of hydration is removed. A 1.628 g sample is heated, leaving 1.072 grams of magnesium iodide. What is the value of *X* ?

30) 2 NH3 + 5 F2 🡪 N2F4 + 6 HF

 ?????? ???????

 45.75 grams ??? Liters @STP

31) 3 (NH4)2PtCl6 🡪 3 Pt + 2 NH4Cl + 2 N2 + 16 HCl

 ? ?

 758 g ? L @STP ? grams