General Information

- A chemical reaction is a process in which one or more substances are converted into new substances that have different physical and chemical properties.
- In a chemical reaction, those substances present before the reaction are called the reactants. Substances produced in a chemical reaction are called the products.
- A chemical reaction takes place when an atom can find a more stable way to be joined to other atoms.

Chemical Equations

- Scientists represent chemical reactions in chemical equations. Chemical equations can be in words or in formulas.
- In writing equations, many times it is important to know the state of the atom or molecule you are working with.
- To show this, we use the following abbreviations:
  - (l) - liquid (g) - gaseous
    (s) - solid (aq) - aqueous
- Aqueous means dissolved in water.

Balancing Equations

- In all reactions, the Law of Conservation of Matter is adhered to. Therefore, we must have the same number of atoms of each type of element on both sides of the reaction.
- We accomplish this by placing necessary coefficients in front of the atoms or molecules in order to balance the number of each atom or group on each side of the arrow.

Examples

Balance:

__ $Cu_{(s)}$ + __ $HCl_{(aq)}$ --> __ $CuCl_{2\,(aq)}$ + __ $H_{2\,(g)}$

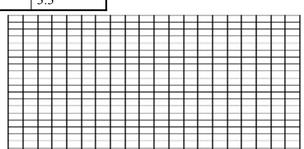
__ $SrF_{2\,(aq)}$ + __ $Cu_3PO_{3\,(aq)}$ --> __ $CuF_{(aq)}$ + __ $Sr_3(PO_3)_{2\,(s)}$

Balance:

Solid calcium is mixed with aqueous nickel (III) sulfate and makes aqueous calcium sulfate and solid nickel.

Solid copper (II) oxide mixes with carbon dioxide gas (CO_2) to make solid copper (II) carbonate.

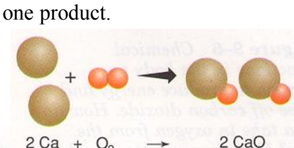
Experiment	mL $AgNO_3$	mL K_2CrO_4	Precipitate (g)
1	5.0	45.0	1.7
2	15.0	35.0	5.0
3	25.0	25.0	8.3
4	30.0	20.0	10.0
5	35.0	15.0	9.9
6	40.0	10.0	6.6
7	45.0	5.0	3.3

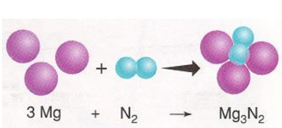


Types of Reactions

In this chemistry course, we will focus on five basic types of reactions:

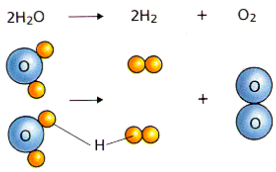
1) Combination or Synthesis - two or more reactants come together to form one product.





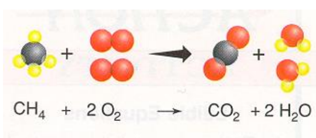
Types of Reactions

2) Decomposition - a single compound is broken down into two or more smaller compounds or elements.



Types of Reactions

5) Combustion - reactions in which a hydrocarbon is combined with O_2 and burned to always form carbon dioxide and water.



Atomic & Formula Mass

As you may recall, protons and neutrons are roughly the same mass, and their mass was given a special unit, the amu.

Protons and neutrons were approximately 1 amu each, while an electron was so small it was held to be 0 amu.

Using this information, we can conclude that the atomic mass of an element is equal to the atomic mass number in amu.

Formula mass is a little different. Formula mass is the sum of the atomic masses in a compound (ex. H_2O = 18.02 amu))

Molar Mass

We know how to find the formula and atomic masses. Molar mass is the next step.

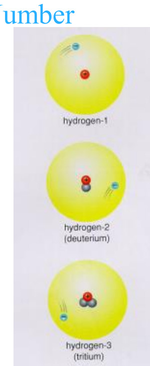
The molar mass of a compound is the mass of one mole of a compound. The unit for this is g/mol. It is directly related to atomic and formula masses.

Example: What is the molar mass of the following compound?: **NH_3**

Answer: mass of 1 N = 14.01 g, mass of 3 H (3 x 1.01 g) = 3.03 g **total mass = 17.04 g/mol**

Isotopes and Mass Number

- Every atom of a given element has the same number of protons. However, like the electrons, the amount of neutrons in any atom of that element can be different.
- The chemical properties of each isotope are identical, except the number of neutrons within the nucleus.



Isotopes and Mass Number

- To identify the different isotopes, scientist add a mass number after the element's name. The mass number is the sum of the number of protons and neutrons (rounded off to the nearest whole number).
- The average mass (in amu) of a group of atoms of the same element is used to find the average atomic mass.

The Mole

The mole is a unit of measure, just like a dozen, a gross or a ream.

To use the atom as the standard unit of measure for the mass is difficult. It is impossible to directly measure the mass of an atom. So, we would like to use a different measure, the gram.

It would make sense to have the mass of a certain number of atoms the same as the atomic mass number, only in grams instead of amu.

Mole Determination

To find the equivalent number of atoms to make the amu# equal to the mass #, we can use the fact that 1 amu = 1.66×10^{-24} g, and determine how many atoms are in a mole.

Through a simple calculation, Amadeo Avogadro, in the early 1800's determined what this number would be. It is also termed Avogadro's number (N).

This number is called the mole. In every mole of a substance, there are 6.02×10^{23} particles.

Mole Conversions

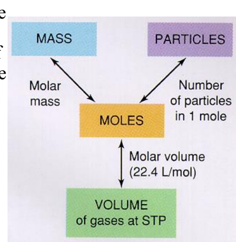
Throughout the rest of the year, we will need to know how many moles are in a certain sample of a compound. We may know the mass, moles, volume of gas (at STP) or # of particles.

Conversion factors

1 mole = 6.02×10^{23} atoms or molecules

1 mole = 22.4 L of gas

1 mole = molar mass



Percent Composition

The percent of the mass made up by each element in a compound.

$$\% \text{ comp.} = \frac{\text{Mass of Element}}{\text{Mass of Compound}} \times 100\%$$

Suppose you had a mole of CO_2. What is the percent composition of each element?

Empirical Formula

A formula that gives the simplest whole-number ratio of the atoms of the elements in a compound.

To determine the empirical formula:

1) Determine the moles of each element in the compound.
2) Set-up a mole ratio for the elements.
3) Divide each number by the smallest mole value.
4) Round numbers off to nearest whole number, and write the formula.

Molecular Formulas

The formula that gives the actual number of atoms of each element in a compound.

In order to find the molecular formula:

1) Determine the molar mass of the molecular compound.
2) Determine the empirical formula mass.
3) Divide (Step1/Step2): The number you get is how many times more atoms are in the molecular formula than the empirical.

Balanced Equations

We already know how to balance an equation, and can induce from that what the equation's coefficients mean:

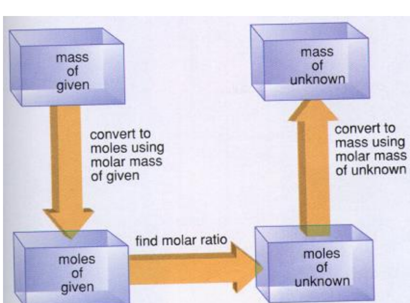
Now, we could multiply every coefficient by the same number, say 20, and still have a balanced equation.

As a result, the coefficients not only tell us the molecule ratio in the reaction, it also tells us the mole ratio.

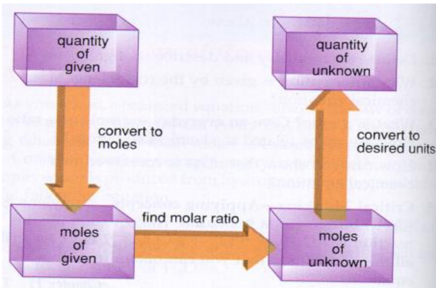
Mole to Mole Stoichiometry

Knowing these facts, it is now possible to predict the amount of chemicals used or produced in a reaction from knowing just one amount of any one of the materials.

Mass to Mass Stoichiometry



Quantity to Quantity Stoichiometry



Limiting Reactant

To determine which reactant is the limiting reactant and which is in excess, it is necessary to calculate the yield of both reactants assuming stoichiometric proportions.

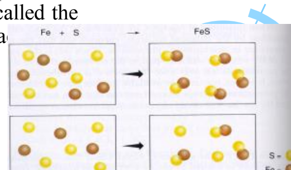
The reactant that produces the smaller yield will be identified as the limiting reactant, since it will run out first.

Do not assume the reactant in smaller amount will automatically be the limiting reactant.

Limiting Reactant

Sometimes chemicals are combined in non-stoichiometric proportions. When this happens, one reactant is completely used up, but there is a excess of the other reactant.

The reactant that limits the amount of product that can be formed is called the limiting rea



Percent Yield

Many times, in performing an experiment, the actual yield (amount you actually produce) will be different from the expected yield (amount you calculate it should produce).

A percent yield is helpful to determine the percent of the calculated yield obtained:

$$\% \text{ yield} = \frac{\text{actual yield}}{\text{expected yield}} \times 100\%$$

Example

Using the reaction:

__ $CaCl_2$ (aq) + __ Na_3PO_4 (aq) --> __$Ca_3(PO_4)_2$ (s) + __ NaCl (aq)

a) If 30.0 g of calcium chloride is mixed with 30.0 grams of sodium phosphate, how many grams of precipitate will be produced?

b) By performing this experiment, it is found that 24.7 g of precipitate form. What is the percent yield?