Atomic Mass

By international agreement, **atomic mass** is *the mass of the atom in atomic mass units (amu)*. One **atomic mass unit** is defined *as a mass exactly equal to one-twelfth the mass of one carbon-12 atom*. Carbon-12 is the carbon isotope that has six protons and six neutrons. Setting the atomic mass of carbon-12 at 12 amu provides the standard for measuring the atomic mass of the other elements. For example, experiments have shown that, on average, a hydrogen atom is only 8.400 percent as massive as the carbon-12 atom. Thus, if the mass of one carbon-12 atom is exactly 12 amu, the atomic mass of hydrogen must be 0.084×12.00 amu or 1.008 amu. Similar calculations show that the atomic mass of oxygen is 16.00 amu and that of iron is 55.85 amu. Thus, although we do not know just how much an average iron atom's mass is, we know that it is approximately 56 times as massive as a hydrogen atom.

Average Atomic Mass

Since many of the naturally occurring elements have different isotopes, taking for example, a sample of carbon we will find that the mass is not exactly 12.00 amu, but rather 12.01 amu. This is due to the averaging of the mass of all the naturally occurring isotopes of carbon.

For example, the natural abundances of carbon-12 and carbon-13 are 98.90 percent and 1.10 percent, respectively. The atomic mass of carbon-13 has been determined to be 13.00335 amu. Thus, the average atomic mass of carbon can be calculated as follows:

average atomic mass of natural carbon = (0.9890)(12.00000 amu) + (0.0110)(13.00335 amu)
= 12.01 amu

Avogadro's Number and the Molar Mass of an Element

In the SI system the **mole (mol)** is *the amount of a substance that contains as many elementary entities (atoms, molecules, or other particles) as there are atoms in exactly 12 g (or 0.012 kg) of the carbon-12 isotope*. The actual number of atoms in 12 g of carbon-12 is determined experimentally. This number is called **Avogadro's number (N_A)**. The currently accepted value is

$$N_a = 6.0221415 \times 10^{23}$$

Generally, we round Avogadro's number to 6.022×10^{23}.

We have seen that 1 mole of carbon-12 atoms has a mass of exactly 12 g and contains 6.022×10^{23} atoms. This mass of carbon-12 is its **molar mass**, defined as *the mass (in grams or kilograms) of 1 mole of units (such as atoms or molecules) of a substance*. Note that the molar mass of carbon-12 (in grams) is numerically equal to its atomic mass in amu. Likewise, the atomic mass of sodium (Na) is 22.99 amu and its molar mass is 22.99 g; the atomic mass of phosphorus is 30.97 amu and its molar mass is 30.97 g; and so on. If we know the atomic mass of an element, we also know its molar mass.

Knowing the molar mass and Avogadro's number, we can calculate the mass of a single atom in grams. For example, we know the molar mass of carbon-12 is 12.00 g and there are 6.022×10^{23} carbon-12 atoms in 1 mole of the substance; therefore, the mass of one carbon-12 atom is given by

$$\frac{12.00\ g}{6.022 \times 10^{23}} = 1.991 \times 10^{-23} g$$

We can use the preceding result to determine the relationship between atomic mass units and grams. Because the mass of every carbon-12 atom is exactly 12 amu, the number of atomic mass units equivalent to 1 gram is

$$\frac{\text{amu}}{\text{gram}} = \frac{12\ \text{amu}}{1\ \cancel{\text{carbon-12 atom}}} \times \frac{1\ \cancel{\text{carbon-12 atom}}}{1.993 \times 10^{-23}\ \text{g}}$$
$$= 6.022 \times 10^{23}\ \text{amu/g}$$

Thus, $1g = 6.022 \times 10^{23} amu$

And $1amu = 1.661 \times 10^{-24} g$

This example shows that Avogadro's number can be used to convert from the atomic mass units to mass in grams and vice versa.

The notions of Avogadro's number and molar mass enable us to carry out conversions between mass and moles of atoms and between moles and number of atoms. We will employ the following formulas.

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$$Amount\ (Number\ of\ moles) = \frac{mass}{molar\ mass}$$

These calculations can also be done using unitary method.
We can find the number of molecules (Particles) by the formula

$$No\ of\ particles\ =\ amount\ of\ substance\ (in\ mol) \times\ Avogadro's\ constant$$

Molecular Mass

The **molecular mass** *is the sum of the atomic masses (in amu) in the molecule*. For example, the molecular mass of H_2O is

$$2(atomic\ mass\ of\ H) + atomic\ mass\ of\ O$$
$$2(1.008) + 16.00 = 18.02$$

In general, we need to multiply the atomic mass of each element by the number of atoms of that element present in the molecule and sum over all the elements.
From the molecular mass we can determine the molar mass of a molecule or compound. The molar mass of a compound (in grams) is numerically equal to its molecular mass (in amu). For example, the molecular mass of water is 18.02 amu, so its molar mass is 18.02 g. Note that 1 mole of water weighs 18.02 g and contains 6.022×10^{23} H_2O molecules, just as 1 mole of elemental carbon contains 6.022×10^{23} carbon atoms.
Finally, note that for ionic compounds like NaCl and MgO that do not contain discrete molecular units, we use the term formula mass instead. The formula unit of NaCl consists of one Na1 ion and one Cl_2 ion. Thus, the formula mass of NaCl is the mass of one formula unit:

$$formula\ mass\ of\ NaCl =\ 22.99\ amu +\ 35.45\ amu = 58.44\ amu$$

and its molar mass is 58.44 g.

Molar volume

For any gaseous substances, the volume of that gas at room temperature and pressure (**RTP**) (25° C, a atm) is 24 dm^3.

$$Gas\ volume = no\ of\ nole \times 24\ dm^3$$

Mass Percentages from the Formula

The **percent composition** *is the percent by mass of each element in a compound*.
Suppose that A is a part of something—that is, part of a whole. It could be an element in a compound or one substance in a mixture. We define the mass percentage of A as the parts of A per hundred parts of the total, by mass. That is,

$$Mass\ \%\ A = \frac{mass\ of\ A\ in\ the\ whole}{Mass\ of\ whole} \times 100\%$$

You can look at the mass percentage of A as the number of grams of A in 100 g of the whole.

THE MASS SPECTROMETER

The mass spectrometer can be used to determine all the isotopes present in a sample of an element and to therefore identify elements.
It needs to be under a vacuum otherwise air particles would ionise and register on the detector Path of lighter ion.
The following are the essential 4 steps in a mass spectrometer.

1. **Ionization**
   - A Vaporized sample is injected at low pressure
     - If the sample is not vaporized then vaporizing it would be the first step.
   - An electron gun fires high energy electrons at the sample
   - This Knocks out an (outer) electron
   - Forming positive ions with different charges E.g. Ti → Ti^+ + e^-
2. **Acceleration**
   - A negative electric field accelerates the positive ions and makes them into a beam
3. **Deflection**
   - The beam of positive ions is deflected by a strong magnetic field.
   - The degree of deflection depends on the mass-to-charge ratio, m/z.

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- The heavier an ion the less it is deflected
- The smaller this ratio the larger the deflection.
- By varying the magnetic field ratio, ions of different m/z ratios pass through the center.

4. **Detection**
   - The ions reach the detector and generate a small current, which is fed to a computer for analysis. The current is produced by electrons transferring from the detector to the positive ions. The size of the current is proportional to the abundance of the species

The figure below summarizes a basic setup of mass spectroscopy. It should be also noted that there is also a step between acceleration and deflection, which is velocity selection, which ensures that the ions passing through the deflectors are all moving at a same velocity. Keeping the velocity constant is very important, as changing the velocity can mess up with detection steps, as curvature of deflection is dependent on the velocity of the ions.

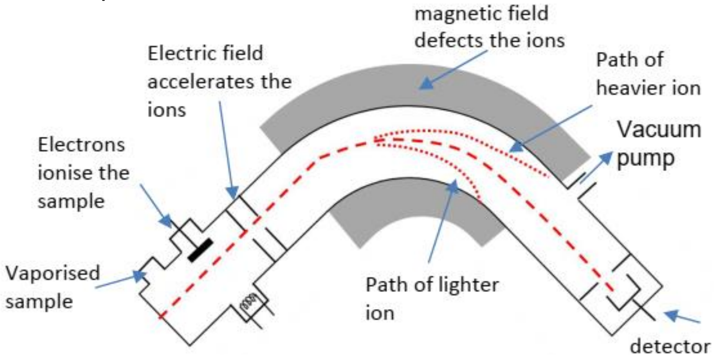
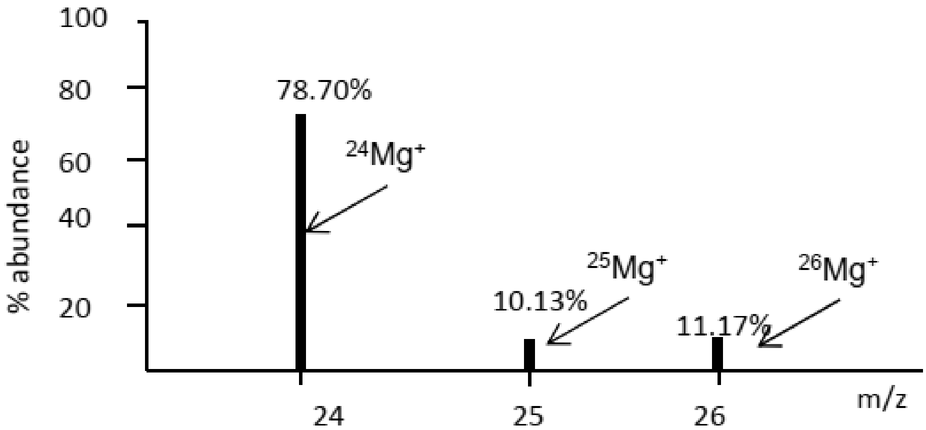


Figure 1 Schematic of a basic mass spectroscopy setup.

Calculating relative atomic mass from mass spectra

The relative atomic mass quoted on the periodic table is a weighted average of all the isotopes. The figure below shows a mass spectra for magnesium.



For each isotope the mass spectrometer can measure a **m/z (mass/charge ratio)** and an **abundance**. Sometimes two electrons may be removed from a particle forming a 2+ ion. $^{24}Mg^{2+}$ with a 2+ charge would have a m/z of 12

The formula for Relative atomic mass is

$$R.A.M. = \frac{\sum(Isotopic\ mass \times \%\ abundance)}{100}$$

For above example of Mg

$$R.A.M\ =\ [(78.7\ x\ 24)\ +\ (10.13\ x\ 25)\ +\ (11.17\ x\ 26)]\ /100\ =\ 24.3$$

If instead of percentage abundance, the spectrum shows relative abundance the formula becomes.

$$R.A.M. = \frac{\sum(Isotopic\ mass \times \%\ abundance)}{total\ relative\ abundance}$$

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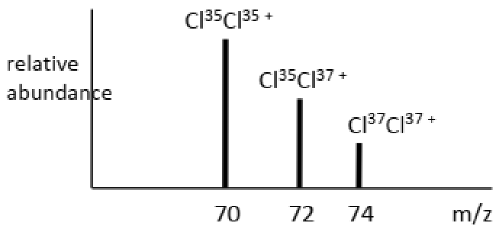
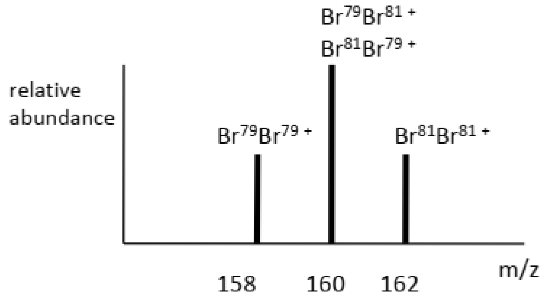
Mass spectra of molecules

Below are mass spectra of Cl_2 and Br_2.

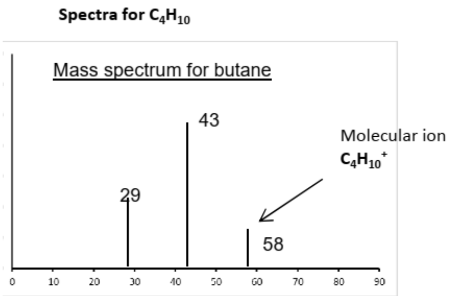
Cl has two isotopes Cl^{35} (75%) and Cl^{37}(25%) Br has two isotopes Br^{79} (50%) and Br^{81}(50%). These lead to the following spectra caused by the diatomic molecules





Measuring the Mr of a molecule

If a molecule is put through a mass spectrometer it will often break up and give a series of peaks caused by the fragments. The peak with the largest m/z, however, will be due to the complete molecule and will be equal to the M_r of the molecule. This peak is called the parent ion or molecular ion.



Uses of Mass spectrometers

- Mass spectrometers have been included in planetary space probes so that elements on other planets can be identified. Elements on other planets can have a different composition of isotopes.
- Drug testing in sport to identify chemicals in the blood and to identify breakdown products from drugs in body
- quality control in pharmaceutical industry and to identify molecules from sample with potential biological activity
- radioactive dating to determine age of fossils or human remains

Radioactive Carbon Dating

All living things have small amounts of the radioactive Carbon-14 isotope. When a living thing dies no more 14C is produced and its starts to decay. The object becomes less radioactive over time.

Measure the abundance of 14C in the material to be tested. By use of the half-life of 14C work out how old the object is by working out how much it has decayed.

Examples

1. What is the amount, in mol, in 35g of $CuSO_4$?

$$amount = \frac{mass}{Mr} = \frac{35}{63.5 + 32 + 16 \times 4} = 0.219\, mol$$

2. What is the volume in dm3 at room temperature and pressure of 50g of Carbon dioxide gas ?

$$Amount = \frac{mass}{Mr} = \frac{50}{(12 + 16 \times 2)} = 1.136$$

$$Gas\ volume = 1.136 \times 24\, dm^3 = 27.26\, dm^3$$

3. How many atoms of Tin are there in a 6.00 g sample of Tin metal?

$$amount = \frac{mass}{Mr} = \frac{6}{118.7} = 0.05055\, mol$$

$$Number\ of\ atoms = amount \times 6.022 \times 10^{23} = 0.05055 \times 6.022 \times 10^{23} = 3.04 \times 10^{22}$$

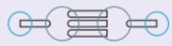
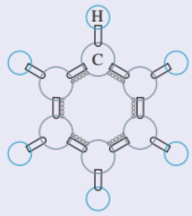
4. Do example **3.1, 3.2, 3.3, 3.4, 3.6, 3.7, 3.8, 3.9, 3.10** from Chang using the formula shown here

Determining Formulas

The percentage composition of a compound leads directly to its empirical formula. An **empirical formula** (or simplest formula) for a compound *is the formula of a substance written with the smallest integer (whole number) subscripts.* For most ionic substances, the empirical formula is the formula of the compound.

The empirical formula tells you the ratio of numbers of atoms in the compound. The empirical formula of hydrogen peroxide (H_2O_2) is HO

Compounds with different molecular formulas can have the same empirical formula, and such substances will have the same percentage composition. An example is acetylene (Ethyne), C_2H_2, and benzene, C_6H_6.

Compound	Empirical Formula	Molecular Formula	Molecular Model
Acetylene	CH	C_2H_2	
Benzene	CH	C_6H_6	

The general method for finding empirical formula from percentage composition employs the following steps.

Step 1. Divide each mass (or % mass) by the atomic mass of the element

Step 2. For each of the answers from step 1 divide by the smallest one of those numbers.

Step 3. Step 3: sometimes the numbers calculated in step 2 will need to be multiplied up to give whole numbers.

These whole numbers will be the empirical formula.

Example

4. Calculate the empirical formula for a compound that contains 1.82g of K, 5.93g of I and 2.24g of O

Step1: Divide each mass by the atomic mass of the element

K = 1.82 / 39.1	I = 5.93/126.9	O = 2.24/16
= 0.0465 mol	= 0.0467mol	= 0.14 mol

Step 2 For each of the answers from step 1 divide by the smallest one of those numbers.

K = 0.0465/0.0465	I = 0.0467/0.0465	O = 0.14 / 0.0465
=1	= 1	= 3

$\boxed{\text{Empirical formula} = KIO_3}$

Molecular Formula from Empirical Formula

The molecular formula of a compound is a multiple of its empirical formula. For example, the molecular formula of acetylene, C_2H_2, is equivalent to $(CH)_2$, and the molecular formula of benzene, C6H6, is equivalent to (CH)6. Therefore, the molecular mass is some multiple of the empirical formula mass, which is obtained by summing the atomic masses of the atoms in the empirical formula. For any molecular compound, you can write

$$Molecular\ mass = n \times empirical\ formula\ mass$$

where n is the number of empirical formula units in the molecule. You get the molecular formula by multiplying the subscripts of the empirical formula by n, which you calculate from the equation

$$n = \frac{molecular\ mass}{empirical\ formula\ mass}$$

Once you determine the empirical formula for a compound, you can calculate its empirical formula mass. If you have an experimental determination of its molecular mass, you can calculate n and then the molecular formula.

Examples

5. work out the molecular formula for the compound with an empirical formula of C_3H_6O and a M_r of 116

   C_3H_6O has a mass of 58

   The empirical formula fits twice into M_r of 116

   So molecular formula is $C_6H_{12}O_2$
6. Do example **3.11** from Chang using the methods shown in this lecture.

Chemical Reactions and Chemical Equations

A **chemical reaction** *is a process in which a substance (or substances) is changed into one or more new substances*. To communicate with one another about chemical reactions, chemists have devised a standard way to represent them using chemical equations. A **chemical equation** *uses chemical symbols to show what happens during a chemical reaction*.

This lecture is assuming that you know how to write and balance equations. If you have lacking, kindly consult section 3.7 of Chang (pg 75-79)

Amounts of Reactants and Products

Stoichiometry (pronounced “stoy-keyom -e-tree”) *is the calculation of the quantities of reactants and products involved in a chemical reaction*. It is based on the chemical equation and on the relationship between mass and moles. Such calculations are fundamental to most quantitative work in chemistry. Whether the units given for reactants (or products) are moles, grams, liters (for gases), or some other units, we use moles to calculate the amount of product formed in a reaction. This approach is called the **mole method**, which means simply that *the stoichiometric coefficients in a chemical equation can be interpreted as the number of moles of each substance*. For example, industrially ammonia is synthesized from hydrogen and nitrogen as follows:

$$N_2(g) + 3H_2(g) \rightarrow 2NH_3(g)$$

The stoichiometric coefficients show that one molecule of N_2 reacts with three molecules of H_2 to form two molecules of NH_3. It follows that the relative numbers of moles are the same as the relative number of molecules:

$N_2(g)$	+	$3H_2(g)$	$\longrightarrow$	$2NH_3(g)$
1 molecule		3 molecules		2 molecules
6.022×10^{23} molecules		$3(6.022 \times 10^{23}$ molecules$)$		$2(6.022 \times 10^{23}$ molecules$)$
1 mol		3 mol		2 mol

Thus, this equation can also be read as “1 mole of N_2 gas combines with 3 moles of H_2 gas to form 2 moles of NH_3 gas.” In stoichiometric calculations, we say that 3 moles of H_2 are equivalent to 2 moles of NH_3, that is,

$$3\,mol\,H_2 \cong 2\,mol\,NH_3$$

Where, the symbol ≅ represents “stoichiometrically equivalent to” or simply “equivalent to.” This relationship enables us to write the conversion factors

$$\frac{3\,mol\,H_2}{2\,mol\,NH_3} \quad and \quad \frac{2\,mol\,NH_3}{3\,mol\,H_2}$$

Similarly, we have 1 mol N= ≅ 2 mol NH= and 1 mol N_2 ≅ 3 mol H_2

Let’s consider a simple example in which 6.0 moles of H_2 react completely with N_2 to form NH_3. To calculate the amount of NH_3 produced in moles, we use the conversion factor that has H_2 in the denominator and write

$$\begin{aligned}\text{moles of NH}_3\text{ produced} &= 6.0\,\cancel{\text{mol H}_2} \times \frac{2\text{ mol NH}_3}{3\,\cancel{\text{mol H}_2}} \\ &= 4.0\text{ mol NH}_3\end{aligned}$$

Now suppose 16.0 g of H_2 react completely with N_2 to form NH_3. How many grams of NH_3 will be formed? To do this calculation, we note that the link between H_2 and NH_3 is the mole ratio from the balanced equation. So we need to first convert grams of H_2 to moles of H_2, then to moles of NH_3, and finally to grams of NH_3. The conversion steps are

$$grams\ of\ H_2 \rightarrow\ moles\ of\ H_2 \rightarrow\ moles\ of\ NH_3 \rightarrow\ grams\ of\ NH_3$$

First, we convert 16.0 g of H_2 to number of moles of H_2, using the molar mass of H_2 as the conversion factor:

$$\text{moles of H}_2 = 16.0\ \cancel{\text{g H}_2} \times \frac{1\ \text{mol H}_2}{2.016\ \cancel{\text{g H}_2}}$$
$$= 7.94\ \text{mol H}_2$$

Next, we calculate the number of moles of NH_3 produced:

$$\text{moles of NH}_3 = 7.94\ \cancel{\text{mol H}_2} \times \frac{2\ \text{mol NH}_3}{3\ \cancel{\text{mol H}_2}}$$
$$= 5.29\ \text{mol NH}_3$$

Finally, we calculate the mass of NH_3 produced in grams using the molar mass of NH_3 as the conversion factor:

$$\text{grams of NH}_3 = 5.29\ \cancel{\text{mol NH}_3} \times \frac{17.03\ \text{g NH}_3}{1\ \cancel{\text{mol NH}_3}}$$
$$= 90.1\ \text{g NH}_3$$

These three separate calculations can be combined in a single step as follows:

$$\text{grams of NH}_3 = 16.0\ \cancel{\text{g H}_2} \times \frac{1\ \cancel{\text{mol H}_2}}{2.016\ \cancel{\text{g H}_2}} \times \frac{2\ \cancel{\text{mol NH}_3}}{3\ \cancel{\text{mol H}_2}} \times \frac{17.03\ \text{g NH}_3}{1\ \cancel{\text{mol NH}_3}}$$
$$= 90.1\ \text{g NH}_3$$

Similarly, we can calculate the mass in grams of N_2 consumed in this reaction. The conversion steps are

$$grams\ of\ H_2 \rightarrow\ moles\ of\ H_2 \rightarrow\ moles\ of\ N_2 \rightarrow\ grams\ of\ N_2$$

By using the relationship 1 mol $N_2 \cong$ 3 mol H_2, we write

$$\text{grams of N}_2 = 16.0\ \cancel{\text{g H}_2} \times \frac{1\ \cancel{\text{mol H}_2}}{2.016\ \cancel{\text{g H}_2}} \times \frac{1\ \cancel{\text{mol N}_2}}{3\ \cancel{\text{mol H}_2}} \times \frac{28.02\ \text{g N}_2}{1\ \cancel{\text{mol N}_2}}$$
$$= 74.1\ \text{g N}_2$$

The figure below summarizes the steps described above.

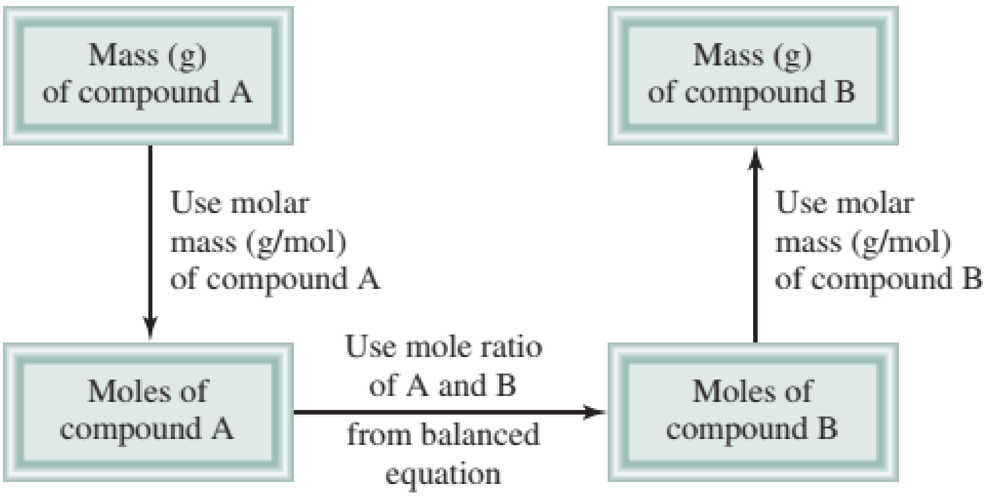


Figure 2 The mole method. First convert the quantity of reactant A (in grams or other units) to number of moles. Next, use the mole ratio in the balanced equation to calculate the number of moles of product B formed. Finally, convert moles of product to grams of product.

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Limiting Reactant; Theoretical and Percentage Yields

When a chemist carries out a reaction, the reactants are usually not present in exact **stoichiometric amounts**, that is, in the *proportions indicated by the balanced equation*. Because the goal of a reaction is to produce the maximum quantity of a useful compound from the starting materials, frequently a large excess of one reactant is supplied to ensure that the more expensive reactant is completely converted to the desired product. Consequently, some reactant will be left over at the end of the reaction. *The reactant used up first in a reaction* is called the **limiting reagent**, because the maximum amount of product formed depends on how much of this reactant was originally present. When this reactant is used up, no more product can be formed. **Excess reagents** are *the reactants present in quantities greater than necessary to react with the quantity of the limiting reagent*.

The concept of limiting reagent is summarized by the example below.

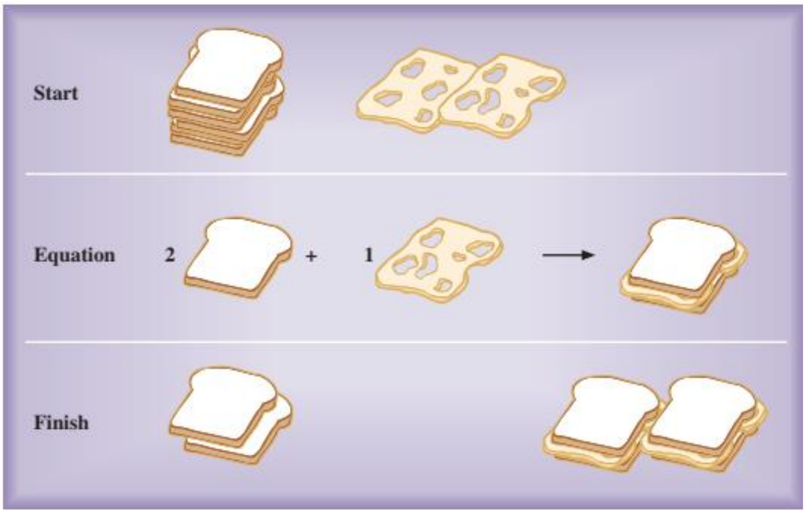


Figure 3 Limiting reactant analogy using cheese sandwiches: Start with six slices of bread, two slices of cheese, and the sandwich-making equation. Even though you have extra bread, you are limited to making two sandwiches by the amount of cheese you have on hand. Cheese is the limiting reactant.

The moles of product are always determined by the starting moles of limiting reactant.

Suppose you put 1 mol H_2 and 1 mol O_2 into a reaction vessel. How many moles of H_2O will be produced? First, you note that 2 mol H_2 produces 2 mol H_2O and that 1 mol O_2 produces 2 mol H_2O. Now you calculate the moles of H2O that you could produce from each quantity of reactant, assuming that there is sufficient other reactant.

$$1\ \cancel{\text{mol H}_2} \times \frac{2\ \text{mol H}_2\text{O}}{2\ \cancel{\text{mol H}_2}} = 1\ \text{mol H}_2\text{O}$$

$$1\ \cancel{\text{mol O}_2} \times \frac{2\ \text{mol H}_2\text{O}}{1\ \cancel{\text{mol O}_2}} = 2\ \text{mol H}_2\text{O}$$

Comparing these results, you see that hydrogen, H_2, yields the least amount of product, so it must be the limiting reactant. By the time 1 mol H_2O is produced, all of the hydrogen is used up; the reaction stops. Oxygen, O_2, is the excess reactant.

The **theoretical yield** of product is *the maximum amount of product that can be obtained by a reaction from given amounts of reactants*. It is the amount that you calculate from the stoichiometry based on the limiting reactant.

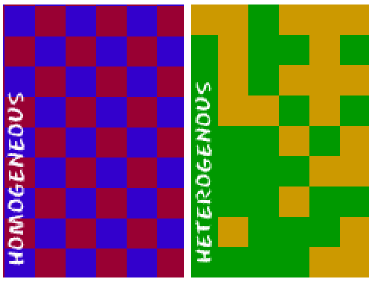
The **percentage yield** of product is *the actual yield (experimentally determined) expressed as a percentage of the theoretical yield (calculated).*

$$\%\ yield = \frac{actual\ yield}{theoretical\ yield} \times 100\%$$

Another important formula is % Atom economy, which shows the efficiency of any reaction theoretically.

$$\%Atom\ economy = \frac{Mass\ of\ useful\ product}{Mass\ of\ all\ reactant} \times 100\%$$

Solution



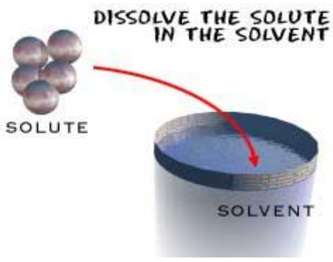
Solution is a type of mixture. Before we dive into solutions, let's separate solutions from other types of mixtures. Solutions are groups of molecules that are mixed and evenly distributed in a system. Scientists say that solutions are **homogenous systems**. Everything in a solution is evenly spread out and thoroughly mixed. Heterogeneous mixtures have a little more of one thing (higher concentration) in one part of the system when compared to another. Let's compare sugar in water (H2O) to sand in water. Sugar dissolves and is spread throughout the glass of water. The sand sinks to the bottom. The sugar-water is a homogenous mixture while the sand-water is a heterogeneous mixture. Both are mixtures, but only the sugar-water can also be called a solution.

Types of solution

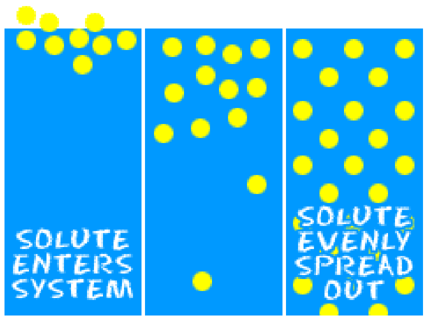
Solutions can be solids dissolved in liquids. Solutions can also be gases dissolved in liquids, such as carbonated water. There can also be gases in other gases and liquids in liquids. If you mix things up and they stay at an even distribution, it is a solution. You probably won't find people making solid-solid solutions. They usually start off as solid/gas/liquid-liquid solutions and then harden at room temperature. Alloys with all types of metals are good examples of solid solutions at room temperature.

SOLUTION	EXAMPLE
Gas-Gas	Air
Gas-Liquid	Carbon Dioxide (CO_2) in Soda
Gas-Solid	Hydrogen (H_2) in Palladium (Pd) Metal
Liquid-Liquid	Gasoline
Liquid-Solid	Dental Fillings
Solid-Solid	Metal Alloys Such as Sterling Silver

Making Solutions



A simple solution is basically two substances that are evenly mixed together. One of them is called the solute and the other is the solvent. A solute is the substance to be dissolved (sugar). The solvent is the one doing the dissolving (water). As a rule of thumb, there is usually more solvent than solute. Be patient with the next sentence as we put it all together. The amount of solute that can be dissolved by the solvent is defined as solubility. That's a lot of "sol" words.



So, How do you make that solution? Mix the two liquids and stir. It's that simple. Science breaks it into three steps. When you read the steps, remember...

Solute=Sugar, Solvent=Water, System=Glass.

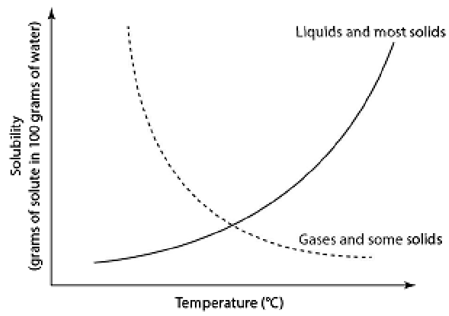
1. The solute is placed in the solvent and the concentrated solute slowly breaks into pieces. Stirring the liquid, the mixing process happens much faster.
2. The molecules of the solvent begin to move out of the way and they make room for the molecules of the solute. Example: The water has to make room for the sugar molecules to spread out.
3. The solute and solvent interact with each other until the concentration of the two substances is equal throughout the system. The concentration of sugar in the water would be the same from a sample at the top, bottom, or middle of the glass.

Colloids

A **colloid** is *a dispersion of particles of one substance (the dispersed phase) throughout another substance or solution (the continuous phase)*. Fog is an example of a colloid: it consists of very small water droplets (dispersed phase) in air (continuous phase). A colloid differs from a true solution in that the dispersed particles are larger than normal molecules,

Solubility

Sure. All sorts of things can change the concentrations of substances in solution.. **Solubility** is *the ability of the solvent (water) to dissolve the solute (sugar)*. You may have already seen the effect of temperature in your classes. Usually when you heat up a solvent, it can dissolve more solid materials (sugar) and less gas (carbon dioxide). If your friend was mixing sugar and water, she would be able to dissolve a lot more sugar into hot water rather than cold.



Next on the list of factors is pressure. When you increase the surrounding pressure, you can usually dissolve more gases in the liquid. Think about your soda can. It is able to keep the fizz inside, because the contents of the can are under higher pressure. Think about a bottle of soda. The first time you open the bottle, a lot of bubbles come out. If you open and close it over a few hours, fewer and fewer bubbles will come out of the solution. When you opened the bottle the first time, you lost the high pressure that was keeping the carbon dioxide (CO_2) gas in solution.

Last is the structure of the substances. Some things dissolve easier in one kind of substance as opposed to another. Sugar dissolves easily in water and oil does not. Water has a low solubility when it comes to oil. Since oil is not soluble in water, it will never truly dissolve. Oil and water is a mixture, not a solution. The two types of molecules (oil and water) are not evenly distributed in the system.

Saturation

Saturation is the point at which a solution can dissolve no more of that substance and any additional amount of solute will appear as undissolved particles. There are three degrees of saturation.

Unsaturated: *a solution not in equilibrium with respect to a given dissolved substance and in which more of the substance can dissolve.*

Saturated: *If the liquid has dissolved the maximum amount of solute it can dissolve at a specific temperature it is called to be saturated*.

Supersaturated: *The liquid contains more solute than it can theoretically dissolve at a given temperature.*

Weak and strong solution (concentration)

A solution can be described as dilute (Weak) or concentrated (Strong). Dilute means that a small amount of solute is dissolved in the solvent. Concentrated means a large amount of solute is dissolved in the solvent. This strength of a solution is known as the concentration of the solution.
Concentration of a solution refers to the amount of solute dissolved in a solution.

Ways of Expressing Concentration

The **concentration** of a solute is t*he amount of solute dissolved in a given quantity of solvent or solution*. The quantity of solvent or solution can be expressed in terms of volume or in terms of mass or molar amount. Thus, there are several ways of expressing the concentration of a solution.
The **molarity** of a solution *is the moles of solute in a liter (dm^3) of solution*.

$$Molarity = \frac{Moles\ of\ solute}{Volume\ of\ solution\ in\ dm^3}$$

For example, 0.20 mol of ethylene glycol dissolved in enough water to give 2.0 dm^3 (2.0 L) of solution has a molarity of

$$\frac{0.20\ mol\ ethylene\ glycol}{2.0\ dm^3\ solution} = 0.1\ moldm^{-3} = 0.1M$$

It is common to use $moldm^{-3}$ instead of M as unit of molarity.
Solution concentration is sometimes expressed in terms of the **mass percentage of solute**—that is, *the percentage by mass of solute contained in a solution*.

$$Mass\ percentage\ of\ solute = \frac{mass\ of\ solute}{mass\ of\ solution} \times 100\%$$

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For example, an aqueous solution that is 3.5% sodium chloride by mass contains 3.5 g of NaCl in 100.0 g of solution. It could be prepared by dissolving 3.5 g of NaCl in 96.5 g of water (100.0 - 3.5 = 96.5).

The **molality** of a solution is *the moles of solute per kilogram of solvent.*

$$Molality = \frac{moles\ of\ solute}{kilogram\ of\ solvent}$$

For example, 0.20 mol of ethylene glycol dissolved in 2.0×10^3 g (= 2.0 kg) of water has a molality of

$$\frac{0.20\ mol\ ethylene\ glycol}{2.0\ kg\ solvent} = 0.10\ m\ ethylene\ glycol$$

That is, the solution is 0.10 molal (denoted m). The units of molality and molarity are sometimes confused. Note that molality is defined in terms of mass of solvent, and molarity is defined in terms of volume of solution.

The **mole fraction** of a component substance A (X_A) in a solution is defined as *the moles of component substance divided by the total moles of solution* (that is, moles of solute plus solvent).

$$X_A = \frac{moles\ of\ substance\ A}{total\ moles\ of\ solution}$$

For example, if a solution is made up of 1 mol of ethylene glycol and 9 mol of water, the total moles of solution are 1 mol + 9 mol = 10 mol. The mole fraction of ethylene glycol is 1/10 = 0.1, and the mole fraction of water is 9/10 = 0.9. Multiplying mole fractions by 100 gives mole percent. Hence, this solution is 10 mole percent ethylene glycol and 90 mole percent water. You can also say that 10% of the molecules in the solution are ethylene glycol and 90% are water. The sum of the mole fractions of all the components of a solution equals 1.

Parts per million (ppm) and Parts per Billion (PPB)

Concentrations can be given also in **parts per million**. This is often used for gases in the atmosphere or in exhausts, and pollutants in water.

$$PPM = \frac{mass\ of\ substance\ in\ mixture}{total\ mass\ of\ mixture} \times 10^6$$

Concentrations can be given also in **parts per billion**. This is often used for gases in the atmosphere or in exhausts, and pollutants in water with concentration much lower to be expressed by PPM.

$$PPB = \frac{mass\ of\ substance\ in\ mixture}{total\ mass\ of\ mixture} \times 10^9$$

Examples

7. What is the concentration of solution made by dissolving 5g of Na_2CO_3 in 250 cm^3 water?

$$amount\ = \frac{mass}{Mr} = \frac{5}{23.1 \times 2\ +\ 12\ + 16\ \times 3}\ = 0.0472\ mol$$

$$conc = \frac{mol}{volume} = \frac{0.0472}{0.25} = 0.189\ moldm^{-3}$$

8. How many chloride ions are there in a 25.0 cm^3 of a solution of magnesium chloride of concentration 0.400 $moldm^{-3}$?

$$amount\ MgCl_2 =\ concentration \times\ Volume = 0.4 \times 0.025 = 0.01\ mol$$

$$Amount\ of\ chloride\ ions = 0.0100 \times 2 = 0.0200\ mol$$

$$Number\ of\ ions\ of\ Cl^- = mole \times 6.02 \times 10^{23} = 0.02 \times 6.02 \times 10^{23} = 1.204 \times 10^{22}$$

9. Blood plasma typically contains 20 parts per million (ppm) of magnesium, by mass. Calculate the mass of magnesium, in grams, present in 100 g of plasma.

$$PPM = \frac{mass\ of\ substance\ in\ mixture}{total\ mass\ of\ mixture} \times 10^6$$

$$20 = \frac{mass\ of\ substance\ in\ mixture}{100} \times 10^6$$

$$mass\ of\ substance\ in\ mixture = 20 \times \frac{100}{10^6} = 2 \times 10^{-3} g$$

10. What mass of Carbon dioxide would be produced from heating 5.5 g of sodium hydrogencarbonate?

$2NaHCO_3 \rightarrow Na_2CO_3 + CO_2 + H_2O$

Step 1: work out amount, in mol, of sodium hydrogencarbonate

amount = mass / Mr

= 5.5 /84

= 0.0655 mol

Step 2: use balanced equation to give amount in mol of CO_2

2 moles $NaHCO_3$: 1 moles CO_2

So 0.0655 HNO_3 : 0.0328mol CO_2

Step 3: work out mass of CO_2

Mass = amount x Mr

= 0.0328 x 44.0

=1.44g

11. What mass of Copper would react completely with 150 cm3 of 1.6M nitric acid?

$3Cu + 8HNO_3 \rightarrow 3Cu(NO_3)_2 + 2NO + 4H_2O$

Step 1: work out moles of nitric acid

amount = conc x vol

= 1.6 x 0.15

= 0.24 mol

Step 2: use balanced equation to give moles of Cu

8 moles HNO_3 : 3 moles Cu

So 0.24 HNO_3 : 0.09 (0.24 x $^3/_8$) mol Cu

Step 3: work out mass of Cu

Mass = amount x Mr

= 0.09 x 63.5

=5.71g

12. What volume in cm^3 of oxygen gas would be produced from the decomposition of 0.532 g of potassium chlorate(V)?

$2KClO_3 \rightarrow 2KCl + 3O_2$

Step 1: work out amount, in mol, of potassium chlorate(V)?

amount = mass / Mr

= 0.532 /122.6

= 0.00434 mol

Step 2: use balanced equation to give amount in mol of O_2

2 moles $KClO_3$: 3 moles O_2

So 0.00434 HNO_3 : 0.00651moles O_2

Step 3: work out volume of O_2

Gas Volume (dm^3)= amount x 24

= 0.00651 x 24

= 0.156 dm^3

= 156 cm^3

13. If one burnt 500 cm^3 of methane at 1atm and 300K what volume of Oxygen would be needed and what volume of CO_2 would be given off under the same conditions?

$$CH_4(g) + \mathbf{2}\, O_2(g) \rightarrow CO_2(g) + 2\, H_2O(l)$$

1 mole   **2** mole   1 mole

500cm^3 1dm^3 500cm^3

Simply multiply gas volume x2

14. An important reaction which occurs in the catalytic converter of a car is
$2CO(g) + 2NO(g) \rightarrow 2CO_2(g) + N_2(g)$
In this reaction, when 500 cm^3 of CO reacts with 500 cm^3 of NO at 650 °C and at 1 atm.
Calculate the **total** volume of gases produced at the same temperature and pressure ?

$2CO(g) + 2NO(g) \rightarrow 2CO_2(g) + N_2(g)$
500cm^3 500cm^3 **500cm^3**   **250cm^3**

total volume of gases produced = 750cm^3

15. 25g of Fe_2O_3 was reacted and it produced 10g of Fe. What is the percentage yield?
$Fe_2O_3 + 3CO \rightarrow 2Fe + 3\, CO_2$

First calculate maximum mass of Fe that could be produced

Step 1: work out amount in mol of Iron oxide
amount = mass / Mr
=25 / 159.6
= 0.1566 mol

Step 2: use balanced equation to give moles of Fe
1 moles Fe_2O_3 : 2 moles Fe
So 0.1566 Fe_2O_3 : 0.313moles Fe

Step 3: work out mass of Fe
Mass = amount x Mr
= 0.313 x 55.8
=17.48g

% yield = (actual yield/theoretical yield) x 100
= (10/ 17.48) x 100
=57.2%

16. What is the % atom economy for the following reaction where Fe is the desired product assuming the reaction goes to completion?

$$Fe_2O_3 + 3CO \rightarrow 2Fe + 3\, CO_2$$

$$\%\ \text{atom economy} = \frac{(2 \times 55.8)}{(2 \times 55.8 + 3\times16) + 3 \times (12+16)} \times 100$$

$$=45.8\%$$

17. Do example **3.12, 3.13, 3.14, 3.15, 3.16** from Chang using the methods shown in this lecture.

Most of the examples in this lecture is taken from chemrevise.org

1. Chang Chapter 3 questions: 3.5, 3.6, 3.7, 3.8, 3.11, 3.13, 3.14, 3.15, 3.20, 3.21, 3.23, 3.24, 3.33, 3.34, 3.41, 3.54, 3.59, 3.65, 3.67, 3.81, 3.89.

Reference books

The following Books are main textbooks that will be followed throughout the course

1. General Chemistry – The Essential Concepts, Raymond Chang and Jason Overby, 6th edition.
2. General Chemistry, Darrell D. Ebbing and Stephen D. Gammon, 9th edition.

If you find the above books difficult, your HSC chemistry books may be handy at times.

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