1.1 Stoichiometry

Stoichiometry is the quantitative study of the composition of compounds and mixtures, and of the amounts of reactants or products involved in a chemical reaction.

compound $\leftrightarrow$

mixture $\leftrightarrow$

In short, stoichiometry is the study and use of quantitative relationships involving amounts of matter. In the lab, we are (usually) restricted to using samples of matter that are large enough to see and manipulate using laboratory equipment. Such samples typically contain uncountable numbers of atoms or molecules. Consequently, counting atoms and molecules directly is simply out of the question. We are forced to deduce the number of atoms or molecules indirectly from some macroscopic measure of amount (e.g. mass).

1.2 The Mole

The SI unit for the amount of substance is the **mole**.

Related units:

1 mmol =

1 μmol =

Why is the mole defined as such a funny number?

The formal definition of the mole is that it is the number of atoms in exactly 12 grams of ^{12}C. Scientists have measured the mass of one ^{12}C, using a mass spectrometer, and its mass is 1.9927×10^{-23} g. Therefore:

Other important quantities

atomic mass unit
(amu or u)

average atomic mass

Avogadro's number (N_A)

Molar mass
(g mol^{-1})

The masses given in a periodic table have two equivalent interpretations.
For example, consider the entry for carbon.

6
C
12.011

The molar mass allows us to calculate the number of moles, *n*, in a weighed sample.

1.2 The Mole Method

The mole method for solving stoichiometry problems involves three basic steps.

(1) convert to moles

(2) convert between moles

(3) convert from moles

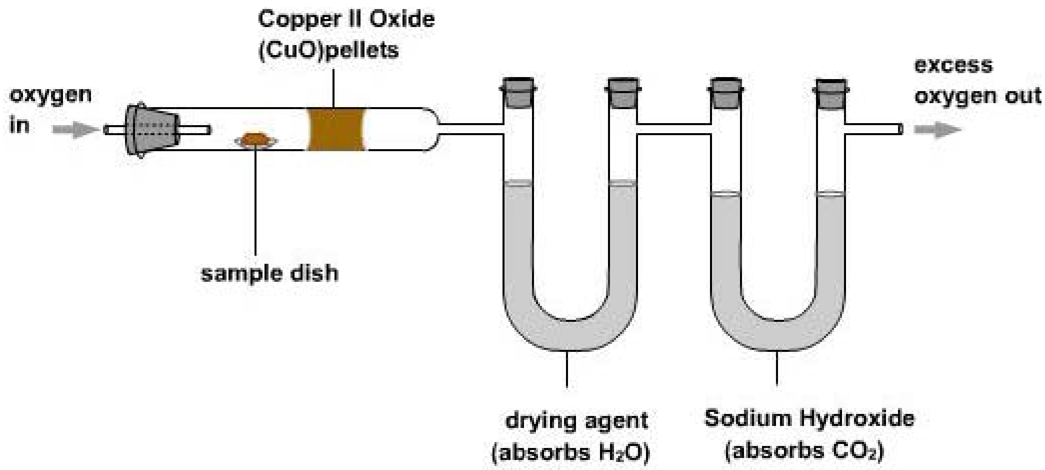
Example: What is the maximum mass of Al_2O_3 that can be obtained from 5.0 g Al and excess O_2, given that Al and O_2 react according to the chemical equation below?

$$4\ Al(s) + 3\ O_2(g) \rightarrow 2\ Al_2O_3(s)$$

Example: (Combustion analysis)

A compound contains only C, H, and O. A 0.875-g sample of the compound is burned in excess O_2 yielding 2.21 g CO_2 and 0.387 g H_2O.

See also problems 3-60, 3-61 and 3-62 from the text.



(a) What is the % composition of the compound?

(b) What is the empirical formula?

(c) In a separate experiment, the molar mass of the compound is found is found to be 245 g/mol. What is the molecular formula of the compound?

The molecular formula tells us exactly how many atoms of each type there are in each molecule.

The empirical formula provides only the ratio in which atoms combine to form the compound.

Example: **(limiting and excess reagents; theoretical, actual and percent yields)**

Small amounts of N_2, for lab use, can be made via the reaction below.

$2\ NH_3(g) + 3\ CuO(s) \rightarrow N_2(g) + 3\ Cu(s) + 3\ H_2O(l)$

A reaction mixture contains 18.1 g NH_3 and 90.4 g CuO. If 6.63 g N_2 is obtained, then what is the percentage yield for this experiment?

Example: (analysis of a mixture, simultaneous reactions)

A 2.00-g sample of a mixture of $CaCl_2$ and RbCl is treated with excess $AgNO_3$(aq), causing AgCl(s) to precipitate from the solution. If 3.45 g AgCl is obtained, then what is the percentage by mass of RbCl in the original mixture?

This is problem B3 from the list of problems on page 9 of the course information booklet. Problem B4 is similar. Try problems 3-96 and 3-100 from the text too.

When faced with two or more reactions, it is very tempting to add together the chemical equations to obtain a single chemical equation for the overall reaction. When considering adding chemical equations together, it is important to distinguish between simultaneous reactions and consecutive reactions.

consecutive reactions =

simultaneous reactions =

Important
For consecutive reactions, you can add the chemical equations together.

For independent reactions, never add the chemical equations together!