THE MOLE

The principle of the **mole** in chemistry. It's a bridge between the number of things and the mass of some known amount of things. It's that simple.

Why we need moles—an example

Take a look at this simple synthesis reaction, in which two hydrogen molecules (H_2) combine with one oxygen molecule (O_2) to produce two molecules of water:

$$2H_2 + O_2 \longrightarrow 2H_2O$$

Now suppose we want to run this reaction, but run it in such a way that we mix together just the right amount of each reactant so that at the end of the reaction there's no extra H_2 or O_2 left over, just H_2O.

The balanced equation says that we need to have two H_2 molecules for every O_2. So we need to "count out" twice as much hydrogen as oxygen. But how do we do that? These molecules are very small.

The trick, again, is to have a nice connection, our **bridge**, between numbers of atoms/ molecules, and mass; we need to know how many atoms of a certain kind are in some given mass.

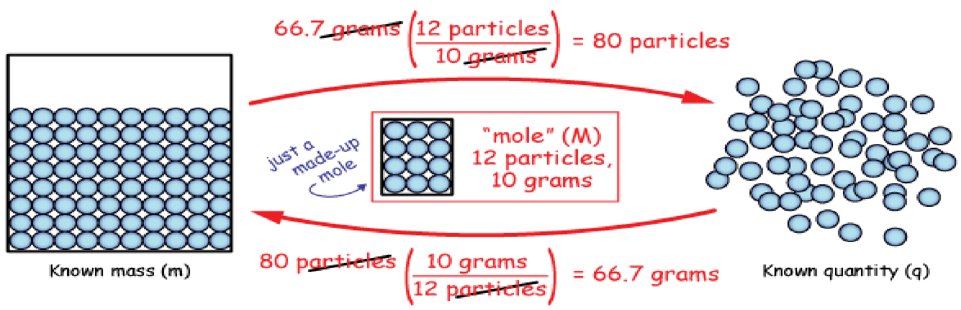
The mole concept

It's a funny name, the "**mole**." It doesn't have anything to do with the varmint that burrows underground. Much of early chemistry was developed by German chemists, and the word "mole" is the English version of the German word "mol" which is short for *molekulargewicht*, or "molecular weight." So it's not so odd after all.

We'll discuss the particulars below, but the mole is basically a known relationship between the mass of a collection of atoms and the number of atoms in that collection. For historical reasons, the mole happens to be the number of atoms in exactly 12.0 grams of pure carbon, but we'll get to that later.

The figure below shows how having the mole (this one is just made up: 12 particles has a mass of 10 g) can serve as the bridge between mass and number. If we know the mass of a known number of particles, we can divide by the mass per number (our "mole") and get the number of particles in that mass.

If we know the number of particles, we can multiply by the mass per number to get the mass. This ability, simple as it seems, will be invaluable in our study of chemistry.



Notice that in the calculations above I've carefully written out and canceled the units to make sure that the calculation represents the conversion I really want to make. You should do that, too.

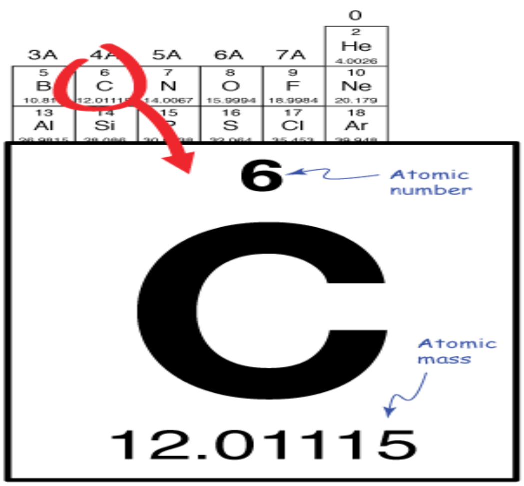
It starts with carbon

We begin, for reasons tied to the historic development of chemistry, with **carbon**.

If we measure the **mass** of one element in an instrument called a **mass spectrometer**, the result is meaningless because a mass spec. can only give us *relative* masses. That is, it can tell us how much heavier or lighter one element is than another, but nothing absolute. We don't have a **scale** for directly measuring the weight of atoms.

So early on, we made a decision: We set the mass of carbon to **12,** in units we called **atomic mass units** (**amu**s) because most carbon has six protons and six neutrons, and they constitute most of the mass of the atom.

Then when we sent other elements through the mass spectrometer, we would get their masses in *multiples* or *fractions* of the carbon mass.



For example, Lithium (Li), would have half the mass of carbon (because it has half the number of heavy particles in its nucleus). Magnesium (Mg) has a mass twice that of carbon, and so on.

In this way, the relative masses of the elements were measured and the periodic table was formed. Much later, masses were adjusted using further knowledge, thus the 12.01115 amu mass for carbon in the figure above.

Mass is proportional to number

Now it's not surprising that the mass of a group of atoms or molecules is directly proportional to the number of atoms or molecules present in the sample.

Nor is it surprising that the mass of an atom is proportional to the number of heavy particles (protons and neutrons) in its nucleus.

If one carbon atom weighs 12 amu, then two will weigh 24 amu, and so on. We'd like to be able to measure the masses of elements like carbon in grams, because amu's are very small units that we can't actually weigh with ease.

So the mass of carbon in grams has to be proportional to the number of atoms present in the sample. What if, for convenience, we made our mole be the number of atoms in 12 grams of carbon. This number was in fact measured in a number of ways around 1910 by physicist Jean Perrin, and he named the special number "Avogadro's number" after Amadeo **Avogadro**, who in about 1810 had proposed that the volume of a gas is proportional to the number of gas atoms present. **Avogadro's number** (**L**) is about **6.022 x 10^{23}** atoms per mole.

Now, here's the beauty of this number: Let's think about Lithium (Li), which, with three protons and three neutrons in its nucleus, has half the atomic mass of carbon. The same number of atoms, each of which weighs half the mass of carbon, should produce a total mass of half of our 12 grams of carbon. That means that in 6 g of Li, there are 6.022 x 10^{23} Li atoms. It turns out that there are 6.022 x 10^{23} atoms of any element in n grams of that element, where n is its atomic mass. It's a very special number.

There are 6.022 x 1023 (Avogadro's number, L) atoms of any element in a mass of that element equal to the atomic mass, but in grams. That mass is called 1 mole of the element.

There are 6.022 x 1023 of anything in one mole of that substance. For example, there are 6.022 x 1023 chickens in a mole of chickens.

Avogadro's number: $L = 6.022 \times 10^{23}$ units per mole

Molar mass

The molar mass of any element is its atomic mass, as read from the periodic table, in grams. So:

- 1 mole of **carbon** has a mass of 12.0 g.
- 1 mole of **iron** has a mass of 58.9 g
- 1 mole of **arsenic** has a mass of 74.9 g,

and so on. And there are **6.022×10^{23}** atoms in 12g of C, and in 58.9g of Fe, and in 74.9g of As.

Now we can do the same thing for **molecules**.

For example the atomic mass of methane (**CH_4**) is 12 amu for the carbon plus 4 x 1 amu for the four hydrogens, for a total of 16 amu. Therefore the **molar mass** of methane is 16g. We say that one mole of methane has a mass of 16 g, and that there are **6.022×10^{23}** atoms in that mass of methane.

Here are a few practice problems. Find the molar mass or **formula weight** (**FW**) of each compound by adding the periodic table masses of all of its elements. Roll over each problem to see the solution.

The molar mass or formula weight (FW) of an atom or molecule is the mass of one mole of that substance. It is found by adding the atomic masses in the periodic table: The FW of an atom is its atomic mass in grams; the FW of a molecule is a sum of the atomic masses of its atoms in grams.

Kr	Water: H_2O
Propane: C_3H_8	Sulfuric acid: H_2SO_4
Magnesium hydroxide: $Mg(OH)_2$	Sodium citrate: $Na_3C_6H_5O_7$

Ammonium sulfate: $(NH_4)_2SO_4$

Triphenyl phosphate: $PO_4(C_6H_5)_3$

A word about precision

You might have noticed that in the calculations above I didn't make use of all of the **precision** that most periodic tables afford in reporting atomic masses. For example, in my periodic table, the mass of oxygen is given as 15.9994 g/mol, which I rounded to 16.

In my view, for most of the "bench chemistry" that people normally do, if one really needs that kind of precision in setting up a reaction or doing some other mole calculation, the experiment probably won't work anyway.

Still, it's not a bad practice to use all of the precision available to you in any calculation, then truncate or round the final result to match the number of lowest precision upon which your result depends. With time you'll work out your own situation-appropriate approach.

Source: http://www.drcruzan.com/ChemistryTheMole.html