The Concept of the Mole

The concept of the mole is so fundamental to Chemistry, yet so poorly understood that I am going to resort to a "fairy tale" to try to describe it. Keep in mind of course that this is NOT how the mole came about, but it serves to get the point across.

Once upon a time there was a Chemist who, having convinced himself that matter was made up of atoms, wondered what the actual masses of those atoms might be. But being a wise chemist he knew that he did not have the technology to determine the actual mass in grams of any atom. Even today, there isn't a balance delicate enough for that. So he settled for the next best thing... their relative mass.

For years, he and his friends had been decomposing water. The decomposition of water had been done so often and measurements of the results taken so carefully that he was sure of the numbers he was working with. Every time water was decomposed to hydrogen and oxygen in the laboratories of his friends and colleagues, it was found that the mass of oxygen produced was always exactly eight times the mass of the hydrogen produced. It did not matter how much water was decomposed, the results were always the same.

Water ------------>	**hydrogen gas**	**+**	**oxygen gas**
any quantity of water	1 unit of mass	**:**	8 units of mass

Now, although this chemist was very smart, he did not know the formula for water. No one did in those days. After thinking about it for a little while, he said "Let us make an assumption. Let us assume that water contains one hydrogen atom for every oxygen atom and that the formula for water is HO." All his friends said, "OK, now what?" They sat stunned as he explained his theory.

If HO is the formula for water he reasoned, then it is obvious that when it decomposes, it does so as follows...

$HO_{(l)}$ -------> $H_{(g)}$ + $O_{(g)}$

If we decompose 1 million water molecules, we will get one million hydrogen atoms and one million oxygen atoms in separate piles. Therefore, if the pile of one million oxygen atoms is 8 times heavier than the pile of 1 million hydrogen atoms, that must mean that each individual oxygen atom is 8 times heavier than each individual hydrogen atom. "So! , he said, "even though we don't know the **actual** mass of either the hydrogen or the oxygen, we can conclude that, one atom of oxygen is 8 times heavier than 1 atom of hydrogen . We have discovered their **relative** mass."

Well, his colleagues went nuts over this relative mass concept. They ran back to their laboratories and started decomposing everything in sight. Before long, using similar reasoning, they had determined the relative masses of 55 of the elements. They knew how much heavier or lighter each atom was compared to every other atom. Along the way they even discovered that the formula for water was H_2O. This meant that one oxygen atom was actually 8 times heavier than 2 hydrogen atoms. This meant in turn that an oxygen atom was actually 16 times heavier than one hydrogen atom. They were so sure of their results that they published them on periodic tables around the world, and there they remain to this day. A relative unit was invented called the atomic mass unit or a.m.u. so that the relative masses would not be reported as bare numbers. Hydrogen's relative mass is still 1 a.m.u., oxygen's is still 16 a.m.u., sulfur's is 32 a.m.u., carbon is 12 a.m.u. and so on. The Chemist then went on to specialize in magic potions where he was much less successful. He nevertheless lived happily ever after.

So what do relative masses have to do with the concept of the mole? Everything, as it turns out. A relative mass is a wonderful bit of knowledge, but it has little practical value in the lab. Lab scales give units of mass in grams. **A new unit needed to be invented that told chemists how many atoms were present in any weighed quantity.** Since they already had these wonderful relative mass numbers published all over the world, they decided to invent a unit that would use these already known relative masses. So they asked themselves 2 very important questions:

1. If the relative mass of oxygen is 16 (no units), how many oxygen atoms would we have to put in a pile to get a pile of oxygen that weighs 16 grams?...and,
2. If the relative mass of sulfur is 32 (no units), how many sulfur atoms would we have to put in a pile to get a pile of sulfur atoms that weighs 32 grams?

To make a long story very short, the answer to both questions was (and still is) the same. After considerable work it was found that 6.02×10^{23} atoms of any element would have a mass in grams numerically equal to the already known relative mass of that atom. This number of atoms (6.02×10^{23}) is known as **one mole**. The End.

To Conclude:

The mole is a unit invented to refer to a quantity of matter. Just as a dozen refers to twelve objects, and a mile refers to 5,280 feet, a mole refers to 6.02×10^{23} atoms or molecules of a substance. Of course, if you put that number of atoms or molecules in a pile on a scale, it will also have a mass, usually referred to as the **molar mass.** The molar mass will be in grams, but will be numerically equal to the relative mass of that substance. Molar masses of the elements are found on the Periodic Table.

For example:

Substance	One Atom's Relative Mass (relative mass in atomic mass units)	The mass of 6.02×10^{23} atoms (molar mass in grams)
Hydrogen (H)	1.01 a.m.u.	1.01 grams
Oxygen (O)	16.00 a.m.u.	16.00 grams
Carbon (C)	12.01 a.m.u.	12.01 grams
Sulfur (S)	32.07 a.m.u.	32.07 grams

Molecules are defined as neutral aggregates of atoms, so the relative mass of a molecule should be the sum of the relative masses of the atoms that make it up. This means also, that its molar mass will be the sum of the molar masses of the atoms that are in it. Again, you need to refer to the Periodic Table to find the numbers you need to calculate both the relative and the molar mass of any molecule or formula unit.

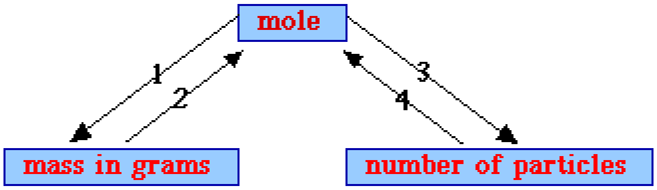
For example:

Substance	**One Molecule's Relative Mass (in atomic mass units)**	**The mass of 6.02 x 10^{23} molecules (molar mass in grams)**
Water (H_2O)	18.02 a.m.u.	18.02 grams
Calcium chloride ($CaCl_2$)	110.98 a.m.u.	110.98 grams
Carbon dioxide (CO_2)	44.01 a.m.u.	44.01 grams
Aluminum sulfate ($Al_2(SO_4)_3$)	342.17 a.m.u.	342.17 grams

When discussing a quantity of matter, a chemist must be able to express it in units of **grams**, **moles** or **number of particles**, and know how to convert from one unit to the other.

Making "Quantity of Matter" Unit Conversions

When a chemist discusses a "quantity of matter", he/she has a number of units to choose from. Primary among them is the mole, which in turn can be converted to either number of particles or mass in grams, so you need to learn how to make these conversions. The following diagram illustrates the relationship between the 3 ways a quantity of matter can be expressed.



The arrows represent the conversions you must be able to perform with ease.

1. represents the conversion of a given number of moles to a mass in grams. Using dimensional analysis, the conversion is done as follows:

What is the mass in grams of 2.60 moles of water ($H_2O_{(l)}$)?

$$2.60\ \cancel{mol\ H_2O}_{(l)} \times \frac{18.02\ g\ H_2O_{(l)}}{1\ \cancel{mol\ H_2O}_{(l)}} = 46.9\ g\ H_2O_{(l)}$$

2. represents the conversion of a given mass in grams of a substance to a number of moles of that substance. Using dimensional analysis, the conversion is done as follows:

How many moles of water are there in 112.3 g of $H_2O_{(l)}$?

$$112.3\ \cancel{g\ H_2O}_{(l)} \times \frac{1\ mol\ H_2O_{(l)}}{18.02\ \cancel{g\ H_2O}_{(l)}} = 6.232\ mol\ H_2O_{(l)}$$

3. represents the conversion of a given number of moles of a substance to the number of particles present. Using dimensional analysis, the conversion is done as follows:

How many molecules of carbon dioxide ($CO_{2(g)}$) are present in 0.850 moles of carbon dioxide?

$$0.850\ \cancel{mol\ CO}_{2(g)} \times \frac{6.02 \times 10^{23}\ molecules}{1\ \cancel{mol\ CO}_{2(g)}} = 5.12 \times 10^{23}\ molecules$$

4. represents the conversion of a number of particles of a substance to moles of that substance. Using dimensional analysis, the conversion is done as follows:

How many moles of $CH_{4(g)}$ are there in 8.20×10^{24} molecules of $CH_{4(g)}$?

$$8.20 \times 10^{24}\ \cancel{molecules\ CH}_{4(g)} \times \frac{1\ mol\ CH_{4(g)}}{6.02 \times 10^{23}\ \cancel{molecules\ CH}_{4(g)}} = 13.6\ moles\ CH_{4(g)}$$