Avogadro's Number and the Mole

Although early scientists were not able to measure the mass of individual atoms, they searched for a workable method that would allow them to define the mass of an atom. They succeeded when, through the work of **Amedeo Avogadro**, they were able to define a **relative atomic mass unit** and develop a **relative scale** to determine the atomic masses for all the elements.

The Atomic Mass Unit

Avogadro stated that equal volumes of different gases at the same temperature and pressure have equal numbers of molecules or atoms. Many experiments have shown his statement to be workable within ±2% and it is now known as **Avogadro's Law**.

Scientists, using Avogadro's Law, were able to show that the masses of carbon and hydrogen were related by a 1:12 ratio. This relationship between carbon and hydrogen was used to define a **relative atomic mass unit,** which provided a workable solution to the problem of atomic masses. Scientists arbitrarily assigned hydrogen the atomic mass of 1 amu since it is the smallest element and defined an **atomic mass unit (amu)** as 1/12 the mass of a carbon atom. Thus, carbon has a mass of 12 amu.

Now, nineteenth century scientists could work with the atomic masses of elements without specialized tools because of a **mathematical relationship** that led to the creation of a relative unit and scale. On this relative scale, if oxygen is sixteen times larger than hydrogen its atomic mass is 16 amu. If sodium is twenty three times the mass of hydrogen then its atomic mass is 23 amu and so on.

Today's scientists are able to measure the masses of atoms with a more direct approach, but their findings are very close to those proposed by the scientists of the eighteenth and nineteenth centuries.

The modern definition for an **atomic mass unit is exactly 1/12 the mass of a carbon-12 isotope**.

Avogadro's number

Armed with the relationship between hydrogen and carbon and a defined atomic mass unit, scientists set out to determine the number of atoms in 12.00 grams of carbon. If the number of atoms in 12.00 grams of carbon could be determined, it would be the same number of atoms in the relative masses of all the other known elements and scientists would have a connection between the macro and subatomic worlds.

After many experiments it was determined that 12.00 grams of carbon-12 contained **6.022 136 7 x 10^{23} atoms**. Today this value is known as **Avogadro's number (N)** in honor of his contributions to chemistry.

Since the mass relationship between carbon and hydrogen is **12 : 1** the number of atoms in 1 gram of hydrogen is also 6.022 136 7 x 10^{23} atoms. This is true for all the elements because their masses are based on the same relationship.

The Mole

A **mole** (mol) is the SI unit for the amount of a substance and is defined as an Avogadro's number of any representative particle. It is important to remember that a mole is a number just as a dozen or a gross are numbers.

Examples: 1 **dozen** eggs = **12** eggs
1 **pair** of shoes = **2** shoes
1 **gross** of pencils = **144** pencils

The mole and Avogadro's number are the links between the macro and subatomic worlds. These numbers allow chemists to construct unit factors to perform many useful calculations. Chemists can now go from particles to moles, moles to grams and even particles to grams.

IMPORTANT CONVERSION:

$$\boxed{\frac{6.022 \times 10^{23}\ \text{particles}}{1\ \text{mole of particles}}}$$

Molar Mass

Since the mole is defined as the number of atoms in 12.00 grams of carbon it can be stated that the mass of one mole of atoms is equal in grams to the numerical value of an element's atomic mass. This is called the **molar mass** and its units are grams/mole.

$$\frac{1.01\ \text{grams H}}{1\ \text{mole H atoms}} \quad \text{or} \quad \frac{16.00\ \text{grams O}}{1\ \text{mole O atoms}}$$

The molar mass is a useful unit factor for many calculations. All you need to know is the element's atomic mass and change the amu's to grams.

Formula and Molecular Weights

The **formula mass** is the sum of the atomic masses of all the atoms in the formula of a substance. The term formula mass is correct for ionic or covalent compounds, but when referring to covalent compounds only, we often use the term **molecular mass**. The molar masses are referred to as **gram molecular mass** and **gram formula mass**.

RULE #1: converting **particles to moles** or **moles to particles** think Avogadro's number.

RULE #2: converting **moles to grams** or **grams to moles** think molar mass.

RULE #3: converting **particles to grams** or **grams to particles** you must use **moles**.

PRACTICE PROBLEMS

SOLVE and show your work**.**

___________________1. Calculate the number of moles in 12 billion molecules of carbon dioxide.

___________________2. Calculate the number of grams in 2.8 mol of calcium chloride.

___________________3. Calculate the number of formula units in 2.5 grams of copper(II) chloride.

___________________4. Calculate the mass of 2.41×10^{24} molecules of hydrogen, H_2.

___________________5. Calculate the number of molecules in 260.0 grams of acrylonitrile, CH_2CHCHN.

"A day of worry is more exhausting than a week of work." -- John Lubbock