Introduction

We don't like to deal with small things individually. We buy eggs and donuts by the dozen, and firecrackers by the gross. Individual eggs, donuts, and firecrackers - from a store's perspective - are not usually worth selling individually. They're not profitable enough and are inconvenient to handle.

Atoms, likewise, are so **small** that handling individual atoms is impractical - so we don't. We work with **groups of atoms.**

The mole concept and the formula weight

Since atoms are so small, we cannot use small groups like dozens to work with atoms. We need a much larger group. The group we use for atoms and molecules is called a **mole**, abbreviated **mol**.

The mole is:

- *6.022×10^{23} objects*
- *The number of atoms in exactly 12 grams of carbon-12*
- *The number of atoms in the atomic weight (expressed in grams) of an element.*

Instead of dealing with individual atoms or molecules, we deal with **moles** of atoms or molecules.

When we work with chemicals, we need to know how many atoms we're handling. Chemical reactions, after all, depend on the **number of atoms** present. This presents us with a small problem: the number of atoms in a sample is not easy to directly measure using only tools readily available in a freshman-level chemistry lab. We normally measure out chemicals by **mass** (for solids and some liquids) or **volume** (gases, some liquids) in our labs.

We will discuss the volume measurements later in the course. For right now, let's focus on mass measurements. We can easily measure the mass of a sample using a balance. Since different kinds of atoms have different masses, we need a way to relate mass to the number of atoms present. Also, since individual atoms are troublesome to deal with, then we would like to **relate the mass of substances we weigh to the moles of atoms or molecules present**.

How? Use the formula weight. The **formula weight** (or **molar mass**) of a substance is the mass of one mole of that substance. For **elements**, the **formula weight is numerically equal to the atomic weight**. The unit, though, is grams instead of atomic mass units. For **compounds**, the **formula weight is equal to the sum of the atomic weights of every atom in the compound**.

We typically express atomic weights and formula weights as “grams per mole”, or g/mol. For example, the formula weight of iron (Fe) is 55.85 g/mol. The formula weight of iron(III) oxide is calculated in the following chart.

Element	***Number of atoms***	***Atomic weight***	***Mass***
Fe	2	55.85	111.70
O	3	16.00	48.00
Total:			**159.70**

So the formula weight of iron(III) oxide, Fe_2O_3, is 159.70 g/mol.

Calculations with the formula weight

The formula weight is used as a **conversion factor** to convert **grams to moles** or **moles to grams**. Here are a few examples.

Calculate the mass of 0.0957 moles of iron(III) oxide.

$$0.0957\ \text{moles}\ Fe_2O_3 \times \frac{159.70\ \text{g}\ Fe_2O_3}{\text{mol}\ Fe_2O_3} = \mathbf{15.3\ g\ Fe_2O_3}$$

Calculate the number of moles of iron(III) oxide in 250. grams of iron(III) oxide.

$$250.\ \text{g}\ Fe_2O_3 \times \frac{\text{mol}\ Fe_2O_3}{159.70\ \text{g}\ Fe_2O_3} = \mathbf{1.57\ mol\ Fe_2O_3}$$

These calculations will crop up in nearly any laboratory work you do in your chemistry courses, so you should practice simple mass/mole conversions until they become automatic. You can find suggested practice problems in the study guides for the course and on the web site.

Application: Using the mole concept to determine the chemical formula

For some compounds, the chemical formula can be experimentally determined in the laboratory using simple techniques and an understanding of the mole concept. For **ionic** compounds, the chemical formula represents the **smallest whole-number ratio of the ions** in the compound. (This is also called the **empirical formula**.)

For **oxides**, which contain only oxygen and one other element, it's possible to determine the formula of the compound by weighing the element, reacting it with oxygen to form the oxide, and then weighing the oxide. For metal oxides, which are solids at room temperature, this is an easy experiment to perform. The metal is burned in a closed container with oxygen, and the metal oxide ash is collected and weighed.

Here's an example to show how it works.

We burn an 0.2000 g sample of titanium (Ti, 47.88 g/mol). We obtain 0.3337 g of a titanium oxide ash. What is the formula of this oxide?

To find the formula, we need to obtain the ratio of titanium atoms to oxygen atoms. This is the **same thing** as the ratio of moles of titanium to moles of oxygen. Finding the moles of titanium is simple. We assume that all the original titanium is now part of the titanium oxide ash.

$$0.2000\,\mathrm{g\,Ti} \times \frac{\mathrm{mol\,Ti}}{47.88\,\mathrm{g\,Ti}} = \mathbf{4.177 \times 10^{-3}\ mol\ Ti}$$

To find the moles of oxygen, we first realize that the titanium oxide ash contains titanium **and** oxygen. So, we can subtract out the titanium to find the mass of oxygen in the compound.

$$0.3337\,\mathrm{g\ titanium\ oxide} - 0.2000\,\mathrm{g\,Ti} = \mathbf{0.1337\,g\,O}$$

Now, find the moles of oxygen in the compound.

$$0.1337\,\mathrm{g\,O} \times \frac{\mathrm{mol\,O}}{16.00\,\mathrm{g\,O}} = \mathbf{8.356 \times 10^{-3}\ mol\,O}$$

The molar ratio of **titanium to oxygen** is **4.177×10^{-3} to 8.356×10^{-3}**.

To write a usable formula, we need to express this ratio in terms of **small whole numbers**. Start simplifying the ratio by dividing both numbers of moles by the smaller of the two numbers. (Mathematically, you can multiply or divide both sides of a ratio by any nonzero number without changing the ratio.)

$$\frac{4.177 \times 10^{-3}\ \mathrm{mol\,Ti}}{4.177 \times 10^{-3}} = \mathbf{1.000\,mol\,Ti}$$

$$\frac{8.356 \times 10^{-3}\ \mathrm{mol\,O}}{4.177 \times 10^{-3}} = \mathbf{2.000\,mol\,O}$$

So the ratio of titanium to oxygen is **1 to 2**, and the formula of the oxide must be **TiO_2**. (The oxide would be named titanium(IV) oxide.)

Keep in mind that in a real experiment, experimental error caused by loss of titanium oxide ash or some other random factor might make the ratio come out a little less exact than the example above.

For some ratios, you may need to multiply by a factor **after** dividing by the smallest number of moles. A compound that actually has a 2:3 ratio of elements would come out of the previous math with a 1:1.5 ratio. Multiplying both sides of the ratio by 2 will give you the correct whole-number ratio.

Summary

This note pack presented the mole concept and the simple calculations required to relate the mass and moles of any substance. A simple application of the concept is the experimental determination of the formula of a metallic oxide. More applications of the mole concept will be discussed later in the course.