General Chemistry I

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Chapter 1. The Atomic Molecular Theory

Foundation

There are over 18 million known substances in our world. We will begin by assuming that all materials are made from **elements**, materials which cannot be decomposed into simpler substances. We will assume that we have identified all of these elements, and that there a very small number of them. All other pure substances, which we call **compounds**, are made up from these elements and can be decomposed into these elements. For example, metallic iron and gaseous oxygen are both elements and cannot be reduced into simpler substances, but iron rust, or ferrous oxide, is a compound which can be reduced to elemental iron and oxygen. The elements are not transmutable: one element cannot be converted into another. Finally, we will assume that we have demonstrated the **Law of Conservation of Mass**.

Law 1.1.

The total mass of all products of a chemical reaction is equal to the total mass of all reactants of that reaction.

These statements are summaries of many observations, which required a tremendous amount of experimentation to achieve and even more creative thinking to systematize as we have written them here. By making these assumptions, we can proceed directly with the experiments which led to the development of the atomic-molecular theory.

Goals

The statements above, though correct, are actually more vague than they might first appear. For example, exactly what do we mean when we say that all materials are made from elements? Why is it that the elements cannot be decomposed? What does it mean to combine elements into a compound? We want to understand more about the nature of elements and compounds so we can describe the processes by which elements combine to form compounds, by which compounds are decomposed into elements, and by which compounds are converted from one to another during chemical reactions.

One possibility for answering these questions is to assume that a compound is formed when indestructible elements are simply mixed together, as for example, if we imagine stirring together a mixture of sugar and sand. Neither the sand nor the sugar is decomposed in the process. And the mixture can be decomposed back into the original components. In this case, though, the resultant mixture exhibits the properties of **both** components: for example, the mixture would taste sweet, owing to the sugar component, but gritty, characteristic of the sand component.

In contrast, the compound we call iron rust bears little resemblance to elemental iron: iron rust does not exhibit elemental iron's color, density, hardness, magnetism, etc. Since the properties of the elements are not maintained by the compound, then the compound must not be a simple mixture of the elements.

We could, of course, jump directly to the answers to these questions by stating that the elements themselves are comprised of atoms: indivisible, identical particles distinctive of that element. Then a compound is formed by combining the atoms of the composite elements. Certainly, the Law of Conservation of Mass would be easily explained by the existence of immutable atoms of fixed mass.

However, if we do decide to jump to conclusions and assume the existence of atoms without further evidence (as did the leading chemists of the seventeenth and eighteenth centuries), it does not lead us anywhere. What happens to iron when, after prolonged heating in air, it converts to iron rust? Why is it that the resultant combination of iron and air does not maintain the properties of either, as we would expect if the atoms of each are mixed together? An atomic view of nature would not yet provide any understanding of how the air and the iron have interacted or combined to form the new compound, and we can't make any predictions about how much iron will produce how much iron rust. There is no basis for making any statements about the properties of these atoms. We need further observations.

Observation 1: Mass relationships during chemical reactions

The Law of Conservation of Mass, by itself alone, does not require an atomic view of the elements. Mass could be conserved even if matter were not atomic. The importance of the Law of Conservation of Mass is that it reveals that we can usefully measure the masses of the elements which are contained in a fixed mass of a compound. As an example, we can decompose copper carbonate into its constituent elements, copper, oxygen, and carbon, weighing each and taking the ratios of these masses. The result is that every sample of copper carbonate is 51.5% copper, 38.8% oxygen, and 9.7% carbon. Stated differently, the masses of copper, oxygen, and carbon are in the ratio of 5.3 : 4 : 1, for every measurement of every sample of copper carbonate. Similarly, lead sulfide is 86.7% lead and 13.3% sulfur, so that the mass ratio for lead to sulfur in lead sulfide is always 6.5 : 1. Every sample of copper carbonate and every sample of lead sulfide will produce these elemental proportions, regardless of how much material we decompose or where the material came from. These results are examples of a general principle known as the **Law of Definite Proportions**.

Law 1.2.

When two or more elements combine to form a compound, their masses in that compound are in a fixed and definite ratio.

These data help justify an atomic view of matter. We can simply argue that, for example, lead sulfide is formed by taking one lead atom and combining it with one sulfur atom. If this were true, then we also must conclude that the ratio of the mass of a lead **atom** to that of a sulfur **atom** is the same as the 6.5 : 1 lead to sulfur mass ratio we found for the bulk lead sulfide. This atomic explanation looks like the definitive answer to the question of what it means to combine two elements to make a compound, and it should even permit prediction of what quantity of lead sulfide will be produced by a given amount of lead. For example, 6.5g of lead will produce exactly 7.5g of lead sulfide, 50g of lead will produce 57.7g of lead sulfide, etc.

There is a problem, however. We can illustrate with three compounds formed from hydrogen, oxygen, and nitrogen. The three mass proportion measurements are given in the following table. First we examine nitric oxide, to find that the mass proportion is 8 : 7 oxygen to nitrogen. If this is one nitrogen atom combined with one oxygen atom, we would expect that the mass of an oxygen atom is 8/7=1.14 times that of a nitrogen atom. Second we examine ammonia, which is a combination of nitrogen and hydrogen with the mass proportion of 7 : 1.5 nitrogen to hydrogen. If this is one nitrogen combined with one hydrogen, we would expect that a nitrogen atom mass is 4.67 times that of a hydrogen atom mass. These two expectations predict a relationship between the mass of an oxygen atom and the mass of a hydrogen atom. If the mass of an oxygen atom is 1.14 times the mass of a nitrogen atom and if the mass of a nitrogen atom is 4.67 times the mass of a hydrogen atom, then we must conclude that an oxygen atom has a mass which is $1.14 \times 4.67 = 5.34$ times that of a hydrogen atom.

But there is a problem with this calculation. The third line of the following table shows that the compound formed from hydrogen and oxygen is water, which is found to have mass proportion 8:1 oxygen to hydrogen. Our expectation should then be that an oxygen atom mass is 8.0 times a hydrogen atom mass. Thus the three measurements in the following table appear to lead to contradictory expectations of atomic mass ratios. How are we to reconcile these results?

Table 1.1. Mass Relationships for Hydrogen, Nitrogen, Oxygen Compounds

Compound	**Total Mass**	**Mass of Hydrogen**	**Mass of Nitrogen**	**Mass of Oxygen**	**"Expected" Relative Atomic Mass of Hydrogen**	**"Expected" Relative Atomic Mass of Nitrogen**	**"Expected" Relative Atomic Mass of Oxygen**
Nitric Oxide	15.0 g	-	7.0 g	8.0 g	-	7.0	8.0
Ammonia	8.5 g	1.5 g	7.0 g	-	1.5	7.0	-
Water	9.0 g	1.0 g	-	8.0 g	1.0	-	8.0

One possibility is that we were mistaken in assuming that there are atoms of the elements which combine to form the different compounds. If so, then we would not be surprised to see variations in relative masses of materials which combine.

Another possibility is that we have erred in our reasoning. Looking back, we see that we have to assume how many atoms of each type are contained in each compound to find the relative masses of the atoms. In each of the above examples, we assumed the ratio of atoms to be 1:1 in each compound. If there are atoms of the elements, then this assumption must be wrong, since it gives relative atomic masses which differ from compound to compound. How could we find the correct atomic ratios? It would help if we knew the ratio of the atomic masses: for example, if we knew that the oxygen to hydrogen mass ratio were 8:1, then we could conclude that the atomic ratio in water would be 1 oxygen and 1 hydrogen. Our reasoning seems to circular: to know the atomic masses, we must know the **formula** of the compound (the numbers of atoms of each type), but to know the formula we must know the masses.

Which of these possibilities is correct? Without further observations, we cannot say for certain whether matter is composed of atoms or not.

Observation 2: Multiple Mass Ratios

Significant insight into the above problem is found by studying different compounds formed from the same elements. For example, there are actually three oxides of nitrogen, that is, compounds composed only of nitrogen and oxygen. For now, we will call them oxide A, oxide B, and oxide C. Oxide A has oxygen to nitrogen mass ratio 2.28 : 1. Oxide B has oxygen to nitrogen mass ratio 1.14 : 1, and oxide C has oxygen to nitrogen mass ratio 0.57 : 1.

The fact that there are three mass ratios might seem to contradict the Law of Definite Proportions, which on the surface seems to say that there should be just one ratio. However, each mass combination gives rise to a completely unique chemical compound with very different chemical properties. For example, oxide A is very toxic, whereas oxide C is used as an anesthesia. It is also true that the mass ratio is not arbitrary or continuously variable: we cannot pick just any combination of masses in combining oxygen and nitrogen, rather we must obey one of only three. So there is no contradiction: we simply need to be careful with the Law of Definite Proportions to say **that each unique compound** has a definite mass ratio of combining elements.

These new mass ratio numbers are highly suggestive in the following way. Notice that, in each case, we took the ratio of oxygen mass to a nitrogen mass of 1, and that the resultant ratios have a very simple relationship:

$$\begin{aligned} 2.28 : 1.14 : 0.57 &= 2 : 1 : 0.5 \\ &= 4 : 2 : 1 \end{aligned} \qquad ()$$

The masses of oxygen appearing in these compounds are in simple whole number ratios when we take a fixed amount of nitrogen. The appearance of these simple whole numbers is very significant. These integers imply that the compounds contain a multiple of a fixed unit of mass of oxygen. The simplest explanation for this fixed unit of mass is that oxygen is **particulate**. We call the fixed unit of mass an **atom**. We now assume that the compounds have been formed from combinations of atoms with fixed masses, and that different compounds have differing numbers of atoms. The mass ratios make it clear that oxide B contains twice as many oxygen atoms (per nitrogen atom) as does oxide C and half as many oxygen atoms (per nitrogen atom) as does oxide A. The simple mass ratios must be the result of the simple ratios in which atoms combine into molecules. If, for example, oxide C has the molecular formula $\boldsymbol{NO}$, then oxide B has the formula NO_2, and oxide A has the formula NO_4. There are other possibilities: if oxide B has molecular formula $\boldsymbol{NO}$, then oxide A has formula NO_2, and oxide C has formula N_2O. Or if oxide A has formula $\boldsymbol{NO}$, then oxide B has formula N_2O and oxide C has formula N_4O. These three possibilities are listed in the following table.

Table 1.2. Possible Molecular Formulae for Nitrogen Oxides

Assuming that:	**Oxide C is** $\boldsymbol{NO}$	**Oxide B is** $\boldsymbol{NO}$	**Oxide A is** $\boldsymbol{NO}$
Oxide A is	NO_4	NO_2	$\boldsymbol{NO}$
Oxide B is	NO_2	$\boldsymbol{NO}$	N_2O
Oxide C is	$\boldsymbol{NO}$	N_2O	N_4O

We don't have a way (from these data) to know which of these sets of molecular formulae are right. But we can assert that either one of them or one analogous to them is right.

Similar data are found for any set of compounds formed from common elements. For example, there are two oxides of carbon, one with oxygen to carbon mass ratio 1.33:1 and the other with mass ratio 2.66:1. The second oxide must have twice as many oxygen atoms, per carbon atom, as does the first. The general statement of this observation is the **Law of Multiple Proportions**.

Law 1.3.

When two elements combine to form more than one compound, the mass of element A which combines in the first compound with a given amount of element B has a simple whole number ratio with the mass of element A which combines in the second compound with the same given mass of element B.

This sounds confusing, but an example clarifies this statement. Consider the carbon oxides, and let carbon be element B and oxygen be element A. Take a fixed given mass of carbon (element B), say 1 gram. The mass of oxygen which combines with 1 gram of carbon to form the first oxide is

1.33 grams. The mass of oxygen which combines with 1 gram of carbon to form the second oxide is 2.66. These masses are in ratio 2.66 : 1.33 = 2 : 1 , a simple whole number ratio.

In explaining our observations of the Law of Multiple Proportions for the carbon oxides and the nitrogen oxides, we have concluded that the simple mass ratio arises from the simple ratio of atoms contained in the individual molecules. Thus, we have established the following postulates of the **Atomic Molecular Theory**.

Theory

- the elements are comprised of identical atoms
- all atoms of a single element have the same characteristic mass
- these number and masses of these atoms do not change during a chemical transformation
- compounds consist of identical molecules formed of atoms combined in simple whole number ratios

Review and Discussion Questions

Exercise 1.

Assume that matter does not consist of atoms. Show by example how this assumption leads to hypothetical predictions which contradict the Law of Multiple Proportions. Do these hypothetical examples contradict the Law of Definite Proportions? Are both observations required for confirmation of the atomic theory?

Exercise 2.

Two compounds, A and B, are formed entirely from hydrogen and carbon. Compound A is 80.0% carbon by mass, and 20.0% hydrogen, whereas Compound B is 83.3% carbon by mass and 16.7% hydrogen. Demonstrate that these two compounds obey the Law of Multiple Proportions. Explain why these results strongly indicate that the elements carbon and hydrogen are composed of atoms.

Exercise 3.

In many chemical reactions, mass does not appear to be a conserved quantity. For example, when a tin can rusts, the resultant rusty tin can has a greater mass than before rusting. When a candle burns, the remaining candle has invariably less mass than before it was burned. Provide an explanation of these observations, and describe an experiment which would demonstrate that mass is actually conserved in these chemical reactions.

Exercise 4.

The following question was posed on an exam:

> *An unknown non-metal element (Q) forms two gaseous fluorides of unknown molecular formula. A 3.2 g sample of Q reacts with fluorine to form 10.8 g of the unknown fluoride A. A 6.4 g sample of Q reacts with fluorine to form 29.2 g of unknown fluoride B. Using these data only, demonstrate by calculation and explanation that these unknown compounds obey the Law of Multiple Proportions.*

A student responded with the following answer:

> *The Law of Multiple Proportions states that when two elements form two or more compounds, the ratios of the masses of the elements between the two compounds are in a simple whole number ratio. So, looking at the data above, we see that the ratio of the mass of element Q in compound A to the mass of element Q in compound B is 3.2 : 6.4 = 1 : 2 , which is a simple whole number ratio. This demonstrates that these compounds obey the Law of Multiple Proportions.*

Assess the accuracy of the students answer. In your assessment, you must determine what information is correct or incorrect, provide the correct information where needed, explain whether the reasoning is logical or not, and provide logical reasoning where needed.

Solutions

Chapter 2. Relative Atomic Masses and Empirical Formulae

Foundation

We begin by assuming the central postulates of the **Atomic-Molecular Theory**. These are: the elements are comprised of identical atoms; all atoms of a single element have the same characteristic mass; the number and masses of these atoms do not change during a chemical transformation; compounds consist of identical molecules formed of atoms combined in simple whole number ratios. We also assume a knowledge of the observed natural laws on which this theory is based: the **Law of Conservation of Mass**, the **Law of Definite Proportions**, and the **Law of Multiple Proportions**.

Goals

We have concluded that atoms combine in simple ratios to form molecules. However, we don't know what those ratios are. In other words, we have not yet determined any molecular formulae. In [the second table of Concept Development Study #1](), we found that the mass ratios for nitrogen oxide compounds were consistent with many different molecular formulae. A glance back at the nitrogen oxide data shows that the oxide B could be NO , NO_2 , N_2O , or any other simple ratio.

Each of these formulae correspond to different possible relative atomic weights for nitrogen and oxygen. Since oxide B has oxygen to nitrogen ratio 1.14 : 1, then the relative masses of oxygen to nitrogen could be 1.14:1 or 2.28:1 or 0.57:1 or many other simple possibilities. If we knew the relative masses of oxygen and nitrogen atoms, we could determine the molecular formula of oxide B. On the other hand, if we knew the molecular formula of oxide B, we could determine the relative masses of oxygen and nitrogen atoms. If we solve one problem, we solve both. Our problem then is that we need a simple way to "count" atoms, at least in relative numbers.

Observation 1: Volume Relationships in Chemical Reactions

Although mass is conserved, most chemical and physical properties are not conserved during a reaction. Volume is one of those properties which is not conserved, particularly when the reaction involves gases as reactants or products. For example, hydrogen and oxygen react explosively to form water vapor. If we take 1 liter of oxygen gas and 2 liters of hydrogen gas, by careful analysis we could find that the reaction of these two volumes is complete, with no left over hydrogen and oxygen, and that 2 liters of water vapor are formed. Note that the total volume is not conserved: 3 liters of oxygen and hydrogen become 2 liters of water vapor. (All of the volumes are measured at the same temperature and pressure.)

More notable is the fact that the ratios of the volumes involved are simple whole number ratios: 1 liter of oxygen : 2 liters of hydrogen : 2 liters of water. This result proves to be general for reactions involving gases. For example, 1 liter of nitrogen gas reacts with 3 liters of hydrogen gas to form 2 liters of ammonia gas. 1 liter of hydrogen gas combines with 1 liter of chlorine gas to form 2 liters of hydrogen chloride gas. These observations can be generalized into the **Law of Combining Volumes**.

Law 2.1.

When gases combine during a chemical reaction at a fixed pressure and temperature, the ratios of their volumes are simple whole number ratios.

These simple integer ratios are striking, particularly when viewed in the light of our conclusions from the Law of Multiple Proportions. Atoms combine in simple whole number ratios, and evidently, volumes of gases also combine in simple whole number ratios. Why would this be? One simple explanation of this similarity would be that the volume ratio and the ratio of atoms and molecules in the reaction are the same. In the case of the hydrogen and oxygen, this would say that the ratio of volumes (1 liter of oxygen : 2 liters of hydrogen : 2 liters of water) is the same as the ratio of atoms and molecules (1 atom of oxygen: 2 atoms of hydrogen: 2 molecules of water). For this to be true, equal volumes of gas would have to contain equal numbers of gas particles (atoms or molecules), independent of the type of gas. If true, this means that the volume of a gas must be a direct measure of the number of particles (atoms or molecules) in the gas. This would allow us to "count" the number of gas particles and determine molecular formulae.

There seem to be big problems with this conclusion, however. Look back at the data for forming hydrogen chloride: 1 liter of hydrogen plus 1 liter of chlorine yields 2 liters of hydrogen chloride. If our thinking is true, then this is equivalent to saying that 1 hydrogen atom plus 1 chlorine atom makes 2 hydrogen chloride molecules. But how could that be possible? How could we make 2 identical molecules from a single chlorine atom and a single hydrogen atom? This would require us to divide each hydrogen and chlorine atom, violating the postulates of the atomic-molecular theory.

Another problem appears when we weigh the gases: 1 liter of oxygen gas weighs more than 1 liter of water vapor. If we assume that these volumes contain equal numbers of particles, then we must conclude that 1 oxygen particle weighs more than 1 water particle. But how could that be possible? It would seem that a water molecule, which contains at least one oxygen atom, should weigh more than a single oxygen particle.

These are serious objections to the idea that equal volumes of gas contain equal numbers of particles. Our postulate appears to have contradicted common sense and experimental observation. However, the simple ratios of the Law of Combining Volumes are also equally compelling. Why should volumes react in simple whole number ratios if they do not represent equal numbers of particles? Consider the opposite viewpoint: if equal volumes of gas do not contain equal numbers of particles, then equal numbers of particles must be contained in unequal volumes not related by integers. Now when we combine particles in simple whole number ratios to form molecules, the volumes of gases required would produce decidedly non-whole number ratios. The Law of Combining Volumes should not be contradicted lightly.

There is only one logical way out. We will accept our deduction from the Law of Combining Volumes that **equal volumes of gas contain equal numbers of particles**, a conclusion known as **Avogadro's Hypothesis**. How do we account for the fact that 1 liter of hydrogen plus 1 liter of chlorine yields 2 liters of hydrogen chloride? There is only one way for a single hydrogen particle to produce 2 identical hydrogen chloride molecules: each hydrogen particle must contain more than one atom. In fact, each hydrogen particle (or molecule) must contain an even number of hydrogen atoms. Similarly, a chlorine molecule must contain an even number of chlorine atoms.

More explicitly, we observe that

$$1 \text{ liter of hydrogen} + 1 \text{ liter of chlorine} \rightarrow 2 \text{ liters of hydrogen chloride} \tag{2.1}$$

Assuming that each liter volume contains an equal number of particles, then we can interpret this observation as

$$1\ H_2 \text{ molecule} + 1\ \mathrm{Cl}_2 \text{ molecule} \rightarrow 2\ H\mathrm{Cl} \text{ molecules} \tag{2.2}$$

(Alternatively, there could be any fixed even number of atoms in each hydrogen molecule and in each chlorine molecule. We will assume the simplest possibility and see if that produces any contradictions.)

This is a wonderful result, for it correctly accounts for the Law of Combining Volumes and eliminates our concerns about creating new atoms. Most importantly, we now know the molecular formula of hydrogen chloride. We have, in effect, found a way of "counting" the atoms in the reaction by measuring the volume of gases which react.

This method works to tell us the molecular formula of many compounds. For example,

$$2 \text{ liters of hydrogen} + 1 \text{ liter of oxygen} \rightarrow 2 \text{ liters of water} \tag{2.3}$$

This requires that oxygen particles contain an even number of oxygen atoms. Now we can interpret this equation as saying that

$$2\ H_2 \text{ molecules} + 1\ O_2 \text{ molecule} \rightarrow 2\ H_2O \text{ molecules} \tag{2.4}$$

Now that we know the molecular formula of water, we can draw a definite conclusion about the relative masses of the hydrogen and oxygen atoms. Recall from the Table that the mass ratio in water is 8:1 oxygen to hydrogen. Since there are two hydrogen atoms for every oxygen atom in water, then the mass ratio requires that a single oxygen atom weigh 16 times the mass of a hydrogen atom.

To determine a mass scale for atoms, we simply need to choose a standard. For example, for our purposes here, we will say that a hydrogen atom has a mass of 1 on the atomic mass scale. Then an oxygen atom has a mass of 16 on this scale.

Our conclusions account for the apparent problems with the masses of reacting gases, specifically, that oxygen gas weighs more than water vapor. This seemed to be nonsensical: given that water contains oxygen, it would seem that water should weigh more than oxygen. However, this is now simply understood: a water molecule, containing only a single oxygen atom, has a mass of 18, whereas an oxygen molecule, containing two oxygen atoms, has a mass of 32.

Determination of Atomic Weights for Gaseous Elements

Now that we can count atoms and molecules to determine molecular formulae, we need to determine relative atomic weights for all atoms. We can then use these to determine molecular formulae for any compound from the mass ratios of the elements in the compound.

We begin by examining data on reactions involving the Law of Combining Volumes. Going back to the nitrogen oxide data given **here**, we recall that there are three compounds formed from nitrogen and oxygen. Now we measure the volumes which combine in forming each. We find that 2 liters of oxide B can be decomposed into 1 liter of nitrogen and 1 liter of oxygen. From the reasoning above, then a nitrogen particle must contain an even number of nitrogen atoms. We assume for now that nitrogen is N_2 . We have already concluded that oxygen is O_2 . Therefore, the molecular formula for oxide B is $\boldsymbol{NO}$, and we call it nitric oxide. Since we have already determined that the oxygen to nitrogen mass ratio is 1.14 : 1, then, if we assign oxygen a mass of 16, as above, nitrogen has a mass of 14. (That is $\frac{16}{1.14} = 14$.) 2 liters of oxide A is formed from 2 liters of oxygen and 1 liter of nitrogen. Therefore, oxide A is $N\,O_2$, which we call nitrogen dioxide. Note that we predict an oxygen to nitrogen mass ratio of $\frac{32}{14} = 2.28 : 1$, in agreement with the data. Oxide C is $N_2\,O$, called nitrous oxide, and predicted to have a mass ratio of $\frac{16}{28} = 0.57 : 1$, again in agreement with the data. We have now resolved the ambiguity in the molecular formulae.

What if nitrogen were actually N_4 ? Then the first oxide would be $N_2\,O$, the second would be $N_2\,O_2$, and the third would be $N_4\,O$. Furthermore, the mass of a nitrogen atom would be 7. Why don't we assume this? Simply because in doing so, we will always find that the minimum relative mass of nitrogen in any molecule is 14. Although this might be two nitrogen atoms, there is no reason to believe that it is. Therefore, a single nitrogen atom weighs 14, and nitrogen gas particles are N_2 .

Determination of Atomic Weights for Non-Gaseous Elements

We can proceed with this type of measurement, deduction, and prediction for any compound which is a gas and which is made up of elements which are gases. But this will not help us with the atomic masses of non-gaseous elements, nor will it permit us to determine the molecular formulae for compounds which contain these elements.

Consider carbon, an important example. There are two oxides of carbon. Oxide A has oxygen to carbon mass ratio 1.33 : 1 and oxide B has mass ratio 2.66 : 1. Measurement of reacting volumes shows that we find that 1 liter of oxide A is produced from 0.5 liters of oxygen. Hence, each molecule of oxide A contains only half as many oxygen atoms as does an oxygen molecule. Oxide A thus contains one oxygen atom. But how many carbon atoms does it contain? We can't determine this yet because the elemental carbon is solid, not gas. This means that we also cannot determine what the mass of a carbon atom is.

But we can try a different approach: we weigh 1 liter of oxide A and 1 liter of oxygen gas. The result we find is that oxide A weighs 0.875 times per liter as much as oxygen gas. Since we have assumed that a fixed volume of gas contains a fixed number of particles, then 1 liter of oxide A contains just as many particles as 1 liter of oxygen gas. Therefore, each **particle** of oxide A weighs 0.875 times as much as a particle of oxygen gas (that is, an O_2 molecule). Since an O_2 molecule weighs 32 on our atomic mass scale, then a particle of oxide A weighs $0.875 \times 32 = 28$. Now we know the molecular weight of oxide A.

Furthermore, we have already determined from the combining volumes that oxide A contains a single oxygen atom, of mass 16. Therefore, the mass of carbon in oxide A is 12. However, at this point, we do not know whether this is one carbon atom of mass 12, two atoms of mass 6, eight atoms of mass 1.5, or one of many other possibilities.

To make further progress, we make additional measurements on other carbon containing gas compounds. 1 liter of oxide B of carbon is formed from 1 liter of oxygen. Therefore, each oxide B molecule contains two oxygen atoms. 1 liter of oxide B weighs 1.375 times as much as 1 liter of oxygen. Therefore, one oxide B molecule has mass $1.375 \times 32 = 44$. Since there are two oxygen atoms in a molecule of oxide B, the mass of oxygen in oxide B is 32. Therefore, the mass of carbon in oxide B is 12, the same as in oxide A.

We can repeat this process for many such gaseous compounds containing carbon atoms. In each case, we find that the mass of carbon in each molecule is either 12 or a multiple of 12. We never find, for examples, 6 or 18, which would be possible if each carbon atom had mass 6. The simplest conclusion is that a carbon atom has mass 12. Once we know the atomic mass of carbon, we can conclude that the molecular formula of oxide A is $\boldsymbol{CO}$, and that of oxide B is $C\,O_2$.

Therefore, the atomic masses of non-gaseous elements can be determined by mass and volume measurements on gaseous compounds containing these elements. This procedure is fairly general, and most atomic masses can be determined in this way.

Moles, Molecular Formulae and Stoichiometric Calculations

We began with a circular dilemma: we could determine molecular formulae provided that we knew atomic masses, but that we could only determine atomic masses from a knowledge of molecular formulae. Since we now have a method for determining all atomic masses, we have resolved this dilemma and we can determine the molecular formula for any compound for which we have percent composition by mass.

As a simple example, we consider a compound which is found to be 40.0% carbon, 53.3% oxygen, and 6.7% hydrogen by mass. Recall from the Law of Definite Proportions that these mass ratios are independent of the sample, so we can take any convenient sample to do our analysis. Assuming that we have 100.0g of the compound, we must have 40.0g of carbon, 53.3g of oxygen, and 6.7g of hydrogen. If we could count or otherwise determine the number of atoms of each element represented by these masses, we would have the molecular formula. However, this would not only be extremely difficult to do but also unnecessary.

From our determination of atomic masses, we can note that 1 atom of carbon has a mass which is 12.0 times the mass of a hydrogen atom. Therefore, the mass of N atoms of carbon is also 12.0 times the mass of N atoms of hydrogen atoms, no matter what N is. If we consider this carefully, we discover that 12.0g of carbon contains exactly the same number of atoms as does 1.0g of hydrogen. Similarly, we note that 1 atom of oxygen has a mass which is $\frac{16.0}{12.0}$ times the mass of a carbon atom. Therefore, the mass of N atoms of oxygen is $\frac{16.0}{12.0}$ times the mass of N atoms of carbon. Again, we can conclude that 16.0g of oxygen contains exactly the same number of atoms as 12.0g of carbon, which in turn is the same number of atoms as 1.0g of hydrogen. Without knowing (or necessarily even caring) what the number is, we can say that it is the same number for all three elements.

For convenience, then, we **define** the number of atoms in 12.0g of carbon to be 1 **mole** of atoms. Note that 1 mole is a specific number of particles, just like 1 dozen is a specific number, independent of what objects we are counting. The advantage to defining the mole in this way is that it is easy to determine the number of moles of a substance we have, and knowing the number of moles is equivalent to counting the number of atoms (or molecules) in a sample. For example, 24.0g of carbon contains 2.0 moles of atoms, 30.0g of carbon contains 2.5 moles of atoms, and in general, x grams of carbon contains $\frac{x}{12.0}$ moles of atoms. Also, we recall that 16.0g of oxygen contains exactly as many atoms as does 12.0g of carbon, and therefore 16.0g of oxygen contains exactly 1.0 mole of oxygen atoms. Thus, 32.0g of oxygen contains 2.0 moles of oxygen atoms, 40.0g of oxygen contains 2.5 moles, and x grams of oxygen contains $\frac{x}{16.0}$ moles of oxygen atoms. Even more generally, then, if we have m grams of an element whose atomic mass is M, the number of moles of atoms, n, is