Section - 2 CHEMICAL FORMULAE AND STOICHIOMETRY

The language that chemists use to describe the forms of matter and the changes in its composition appears throughout the scientific world. Chemical symbols, formulas, and equations are used in such diverse areas as agriculture, home economics, engineering, geology, physics, biology, medicine, and dentistry. In this section we shall describe the simplest atomic theory. We shall use it as we represent chemical formulas of elements and compounds. Later this theory will be expanded when we discuss chemical changes.

The word "stoichiometry" is derived from the Greek *stoicheion*, which means "first principle or element," and *metron*, which means "measure." **Stoichiometry** describes the quantitative relationships among elements in compounds (composition stoichiometry) and among substances as they undergo chemical changes (reaction stoichiometry). In this section we shall be concerned with chemical formulas and composition stoichiometry. In next section we shall discuss chemical equations and reaction stoichiometry.

Atoms and Molecules

Around 400 BC, the Greek philosopher Democritus suggested that all matter is composed of tiny, discrete, indivisible particles that he called atoms. His ideas, based entirely on philosophical speculation rather than experimental evidence, were rejected for 2000 years.

(The term "atom" comes form the Greek language and means "not divided" or indivisible.)

By the late 1700s, scientists began to realise that the concept of atoms provided an explanation for many experimental observations about the nature of matter.

By the early 1800s, the Law of Conservation of Matter (Section 1) and the Law of Definite Proportions (Section 1) were both accepted as general descriptions of how matter behaves. John Dalton, an English schoolteacher, tried to explain why matter behaves in such simple and systematic ways as those expressed above. In 1808, he published the first "modern" ideas about the existence and nature of atoms. He summarised and expanded the nebulous concepts of early philosophers and scientists; more importantly, his ideas were based on *reproducible experimental results* of measurements by many scientists. Taken together, these ideas form the core of **Dalton's Atomic Theory,** one of the highlights of scientific thought. In condensed form, Dalton's ideas may be stated as follows:

1. An element is composed of extremely small indivisible particles called atoms.
2. All atoms of a given element have identical properties, which differ from those of other elements.
3. Atoms cannot be created, destroyed, or transformed into atoms of another element. (However, later it was found that this statement is true only for chemical reactions and not for nuclear reactions. We will discus this when we discuss Nuclear chemistry in later units).
4. Compounds are formed when atoms of different elements combine with each other in small whole-number ratios.
5. The relative numbers and kinds of atoms are constant in a given compound.

Dalton believed that atoms were solid indivisible spheres, an idea we now reject. But he showed remarkable insight into the nature of matter and its interactions. Some of his ideas could not be verified (or refuted) experimentally at the time. They were based on the limited experimental observations of his day. Even with their shortcomings, Dalton's ideas provided a framework that could be modified and expanded by later scientists. Thus John Dalton is the father of modern atomic theory.

The smallest particle of an element that maintains its chemical identity through all chemical and physical changes is called an **atom** (figure 5). In unit on "atomic structure" we shall study the structure of the atom in detail; let us simply summarise here the main features of atomic composition. Atoms, and therefore *all* matter, consist principally of three fundamental particles: *electrons, protons,* and *neutrons*. These are the basic building blocks of all atoms. The masses and charges of the three fundamental particles are shown in Table 5. The masses of protons and neutrons are nearly equal, but the mass of an electron is much smaller. Neutrons carry no charge. The charge on a proton is equal in magnitude, but opposite in sign, to the charge on an electron. Because atoms are electrically neutral,

$$\boxed{\text{an atom contains equal numbers of electrons and protons.}}$$

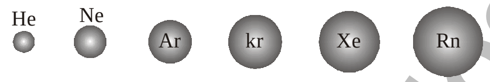


Figure 5. Relative sizes of monatomic molecules (single atoms) of the noble gases

Table 5: ***Fundamental Particles of Matter***

Particle (symbol)	Mass*	Charge (relative scale)
electron (e^-) proton (p or p^+) neutron (n or n^0)	0.00054858 amu 1.0073 amu 1.0087 amu	1– 1+ None
*1 amu = 1.6605×10^{-24} g (We will discuss "amu" as we proceed further in this section)		

The **atomic number** (symbol is **Z**) of an element is defined as the number of protons in the nucleus. In the periodic table, elements are arranged in order of increasing atomic numbers.

A **molecule** is the smallest particle of an element or compound that can have a stable independent existence. In nearly all molecules, two or more atoms are bonded together in very small, discrete units (particles) that are electrically neutral.

Individual oxygen atoms are not stable at room temperature and atmospheric pressure. Single atoms of oxygen mixed under these conditions quickly combine to form pairs. The oxygen with which we are all familiar is made up of two atoms of oxygen; it is a *diatomic* molecule, O_2. Hydrogen, nitrogen, fluorine, chlorine, bromine, and iodine are other examples of diatomic molecules (figure 6)

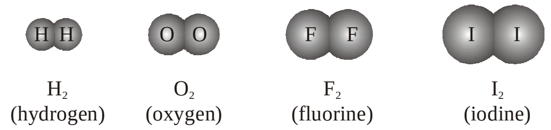
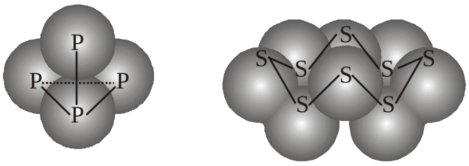


Figure 6: Models of diatomic molecules of some elements approximately to scale.

[**Note:** For Group 0 elements, the noble gases, a molecule contains only one atom and so an atom and a molecule are the same (figure 5).]

Some other elements exist as more complex molecules. Phosphorus molecules consist of four atoms, while sulphur exists as eight-atom molecules at ordinary temperatures and pressures. Molecules that contain two or more atoms are called *polyatomic* molecules. See Figure 7.



(a) (b)

Figure 7 (a) A model of the P_4 molecule of white phosphorus. (b) A model of the S_8 ring found in rhombic sulphur.

[**Note:** You should remember the common elements that occur as diatomic molecule: H_2, O_2, N_2, F_2, Cl_2, Br_2, I_2.]

In modern terminology, O_2 is named dioxygen, H_2 is dihydrogen, P_4 is tetraphosphorus, and so on. Even though such terminology is officially preferred, it has not yet gained wide acceptance. Most chemists still refer to O_2 as oxygen, H_2 as hydrogen, P_4 as phosphorus, and so on.

Molecules of compounds are composed of more than one kind of atom. A water molecule consists of two atoms of hydrogen and one atom of oxygen. A molecule of methane consists of one carbon atom and four hydrogen atoms.

Atoms are the components of molecules, and molecules are the components of many elements and most compounds. We are able to see samples of compounds and elements that consist of large numbers of atoms and molecules. It would take millions of atoms to make a row as long as the diameter of the "." at the end of this sentence.

Chemical Formulas

The **chemical formula** for a substance shows its chemical composition. This represents the elements present as well as the ratio in which the atoms of the elements occur. The formula for a single atom is the same as the symbol for the element. Thus, Na can represent a single sodium atom. It is unusual to find such isolated atoms in nature, with the exception of the noble gases (He, Ne, Ar, Kr, Xe, and Rn). A subscript following the symbol of an element indicates the number of atoms in a molecule. For instance, F_2 indicates a molecule containing two fluorine atoms and P_4 a molecule containing four phosphorus atoms.

Some elements exist in more than one form. Familiar examples include (1) oxygen, found as O_2 molecules, and ozone, found as O_3 molecules (fig. 8) and (2) two different crystalline forms of carbon—diamond and graphite. Different forms of the same element in the same physical state are called **allotropic modifications** or **allotropes**.

Compounds contain two or more elements in chemical combination in fixed proportions. Many compounds exist as molecules (Table 6). Hence, each molecule of hydrogen chloride, HCl, contains one atom of hydrogen and one atom of chlorine; each molecule of carbon tetrachloride, CCl_4, contains one carbon atom and four chlorine atoms.

Table 6 : Names and Formulas of some common molecular compounds.

Name	Formula	Name	Formula	Name	Formula
water	H_2O	sulphur dioxide	SO_2	butane	C_4H_{10}
hydrogen peroxide	H_2O_2	sulphur trioxide	SO_3	pentane	C_5H_{12}
hydrogen chloride*	HCl	carbon monoxide	CO	benzene	C_6H_6
sulphuric acid	H_2SO_4	carbon dioxide	CO_2	methanol (methyl alcohol)	CH_3OH
nitric acid	HNO_3	methane	CH_4	ethanol (ethyl alcohol)	CH_3CH_2OH
acetic acid	CH_3COOH	ethane	C_2H_6	acetone	CH_3COCH_3
ammonia	NH_3	propane	C_3H_8	diethyl ether (ether)	$CH_3CH_2–O–CH_2CH_3$
*Called hydrochloric acid if dissolved in water					

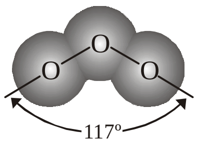


fig. 8. An O_3 molecule

We find many organic compounds in nature. **Organic compounds** contains C—C or C—H bonds or both. Eleven of the compounds listed in Table 6 are organic compounds (acetic acid and the last ten entries). All of the other compounds in the table are **Inorganic compounds**.

Compounds were first recognised as distinct substances because of their different physical properties and because they could be separated from one another by physical methods. Once the concept of atoms and molecules was established, the reason for these differences in properties could be understood: Two compounds differ from one another because their molecules are different. Conversely, if two molecules contain the same number of the same kinds of atoms, arranged the same way, then both are molecules of the same compound. Thus, the atomic theory explains the **Law of Definite Proportions** (Section 1).

This law, also known as the **Law of Constant Composition,** can now be extended to include its interpretation in terms of atoms. It is so important for performing the calculations in the chapter that we restate it here.

Different pure samples of a compound always contain the same elements in the same proportion by mass; this corresponds to atoms of these elements combined in fixed numerical ratios.

So we see that for a substance composed of molecules, the *chemical formula* gives the number of atoms of each type in the molecule. But this formula does not express the order in which the atoms in the molecules are bonded together. The **structural formula** shows the order in which atoms are connected. The lines connecting atomic symbols represent chemical bonds between atoms. The bonds are actually forces that tend to hold atoms at certain distances and angles from one another. For instance, the structural formula of propane shows that the three C atoms are linked in a chain, with three H atoms bonded to each of the end C atoms and two H atoms bonded to the center C (fig. 9)

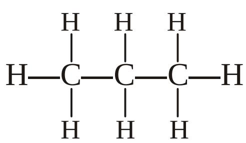


fig. 9. Structural formula of propane

Ions and Ionic Compounds

So far we have discussed only compounds that exist as discrete molecules. Some compounds, such as sodium chloride. NaCl, consist of collections of large numbers of ions. An **ion** is an atom or group of atoms that carries an electrical charge. Ions that possess a *positive* charge, such as the sodium ion, Na^+, are called **cations**. Those carrying a *negative* charge, such as the chloride ion, Cl^-, are called **anions**. The charge on an ion *must* be included as a superscript on the right side of the chemical symbol(s) when we write the formula for the individual ion.

As we shall see in the unit on atomic structure, an atom consists of a very small, very dense, positively charged *nucleus* surrounded by a diffuse distribution of negatively charged particles called *electrons*. The number of positive charges in the nucleus defines the identity of the element to which the atom corresponds. Electrically neutral atoms contain the same number of electrons outside the nucleus as positive charges (protons) within the nucleus. Ions are formed when neutral atoms lose or gain electrons. An Na^+ ion is formed when a sodium atom loses one electron, and a Cl^- ion is formed when a chlorine atom gains one electron.

The compound NaCl consists of an extended array of Na^+ and Cl^- ions. A two-dimensional view of it is shown here:

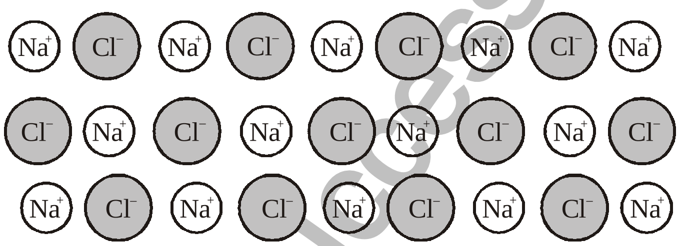


fig. 10. A two-dimensional view of NaCl crystal

Every Na^+ ion is surrounded by six Cl^- ions in 3–D space. In figure above we can see only four Cl^- ions surrounding each Na^+ ion because it is a 2–D figure. You can visualise a Cl^- ion above and below the Na^+ ion coming out of the plane of paper and going inside the plane of paper. And you will have the real structure of NaCl in your mind. (Can you also see that just like Na^+, each Cl^- ion is also surrounded by six Na^+ ions?)

An important point that should be noted in the discussion above is that in the NaCl crystal, the constituent substances are Na^+ and Cl^- ions and **NOT** NaCl molecules. There is no such thing as a NaCl molecule. This is because, as you can see in the figure 10 above, we cannot specify so as to which Cl^- ion belongs to which Na^+ ion as each Na^+ ion attracts all the six Cl^- ions surrounding it equally with electrostatic force of attraction. This is the case with all the ionic compounds.

Because there are no "molecules" of ionic substances, we should not refer to "a molecule of sodium chloride, NaCl," for example. Instead, we refer to a **formula unit** (FU) of NaCl, which consists of one Na^+ ion and two Cl^- ion. Likewise, one formula unit of $CaCl_2$ consists of one Ca^{2+} ion and two Cl^- ions. Similarly, we speak of the formula unit of all ionic compounds. It is also acceptable to refer to a formula unit of a molecular compound. One formula unit of propane, C_3H_8, is the same as one molecule of C_3H_8; it contains three C atoms and eight H atoms bonded together into a group.

For the present, we shall tell you which substances are ionic and which are molecular when it is important to know. Later you will learn to make the distinction yourself.

(Please note that the general term "formula unit" applies to molecular or ionic compounds whereas the more specific term "molecule" applies *only* to elements and compounds that exist as discrete molecules.)

Important points

(1) Compound, whether ionic or molecular, is electrically neutral; i.e., it has no net charge. In NaCl this means that the Na^+ and Cl^- ions are present in a 1 : 1 ratio and this is indicated by the formula NaCl.

(2) **Polyatomic ions** are groups of atoms that bear an electrical charge. Examples include the ammonium ion, NH_4^+, the sulphate ion, SO_4^{2-}, and the nitrate ion, NO_3^-. Table 7 shows the formulas, ionic charges, and names of some common ions.

Table 7: Formulas, Ionic Charges, and Names of Some Common Ions

Formula	Charge	Name	Formula	Charge	Name
Na^+	1+	sodium	F^-	1–	fluoride
K^+	1+	potassium	Cl^-	1–	chloride
NH_4^+	1+	ammonium	Br^-	1–	bromide
Ag^+	1+	silver	OH^-	1–	hydroxide
			CH_3COO^-	1–	acetate
Mg^{2+}	2+	magnesium	NO_3^-	1–	nitrate
Ca^{2+}	2+	calcium			
Zn^{2+}	2+	zinc	O^{2-}	2–	oxide
Cu^+	1+	copper(I) or cuprous	S^{2-}	2–	sulphide
Cu^{2+}	2+	copper(II) or cupric	SO_3^{2-}	2–	sulphite
Fe^{2+}	2+	iron(II) or ferrous	SO_4^{2-}	2–	sulphate
			CO_3^{2-}	2–	carbonate
Fe^{3+}	3+	iron(III) or ferric			
Al^{3+}	3+	aluminum	PO_4^{3-}	3–	phosphate

As we shall see in later units, some metals can from move than one kind of ion with positive charge. For such metals, we specify which ion we mean with a Roman numeral–e.g., iron (II) or iron (III). Because zinc forms no stable ions other than Zn^{2+}, we do not need to use Roman numerals in its name.

An Introduction to Naming Compounds

Throughout your study of chemistry you will have many occasions to refer to compounds by name. In this section, we shall see how to name a few compounds. More comprehensive rules for naming compounds will be presented at the appropriate place later in the unit on "general inorganic chemistry."

Table 6 includes examples of names for a few common molecular compounds. You should learn that short list. We shall name many more molecular compounds as we encounter them in later chapters.

The names of some common ions appear in Table 7. You should learn the names and formulas of these frequently encountered ions. They can be used to write the formulas and names of many ionic compounds. We write the formula of an ionic compound by adjusting the relative numbers of positive and negative ions so their total charges cancel (i.e., add to zero). The name of an ionic compound is formed by giving the names of the ions, with the positive ion named first.

Example – 12

Write the formulas for the following ionic compounds: (a) sodium fluoride, (b) calcium fluoride, (c) iron (II) sulphate, (d) zinc phosphate.

Critical thinking

In each case, we identify the chemical formulas of the ions from Table 7. These ions must be present in a ratio that gives the compound no net charge. The formulas and names of ionic compounds are written by giving the positively charged ion first.

Solution: **(a)** The formula for the sodium ion is Na^+ and the formula for the fluoride ion is F^- (Table 7). Because the charges on these two ions are equal in magnitude, the ions must be present in equal numbers, or in a 1 : 1 ratio. Thus, the formula for sodium fluoride is **NaF**.

(b) The formula for the calcium ion is Ca^{2+} and the formula for the fluoride ion is F^-. Now each positive ion (Ca^{2+}) provides twice as much charge as each negative ion (F^-). So there must be twice as many F^- ions as Ca^{2+} ions to equalise the charge. This means that the ratio of calcium to fluoride ions is 1 : 2. So the formula for calcium fluoride is CaF_2.

(c) The iron(II) ion is Fe^{2+} and the sulphate ion is SO_4^{2-} . As in part (a), the equal magnitudes of positive and negative charges tell us that the ions must be present in equal numbers, or in a 1 : 1 ratio. The formula for iron(II) sulphate is $FeSO_4$.

(d) The zinc ion is Zn^{2+} and the phosphate ion is PO_4^{3-} . Now it will take *there* Zn^{2+} ions to account for as much charge (6+ total) as would be present in *two* PO_4^{3-} ions (6– total). So the formula for zinc phosphate is $Zn_3(PO_4)_2$.

Example – 13

Name the following ionic compounds: (a) $(NH_4)_2S$, (b) $Cu(NO_3)_2$, (c) $ZnCl_2$, (d) $Fe_2(CO_3)_3$.

Critical thinking

In naming ionic compounds, it is helpful to inspect the formula for atoms or groups of atoms that we recognise as representing familiar ions.

Solution: **(a)** The presence of the polyatomic grouping NH_4 in the formula suggests to us the presence of the ammonium ion, NH_4^+. There are two of these, each accounting for 1 + in charge. To balance this, the single S must account for 2– in charge, or S^{2-}, which we recognise as the sulphide ion. Thus, the name of the compound is **ammonium sulphide**.

(b) The NO_3 grouping in the formula tells us that the nitrate ion, NO_3^-, is present. Two of these nitrate ions account for 2 × 1 – = 2 – in negative charge. To balance this, copper must account for 2+ charge and be the copper(II) ion. The name of the compound is **copper(II) nitrate** or, alternatively, **cupric nitrate**.

(c) The positive ion present is zinc ion, Zn^{2+}, and the negative ion is chloride, Cl^-. The name of the compound is **zinc chloride.**

(d) Each CO_3 grouping in the formula must represent the carbonate ion, CO_3^{2-}. The presence of *three* such ions accounts for a total of 6– in negative charge, so there must be a total of 6+ present in positive charge to balance this. It takes *two* iron ions to provide this 6+, so each ion must have a charge of 3+ and be Fe^{3+}, the iron(III) ion or ferric ion. The name of the compound is **iron(III) carbonate** or **ferric carbonate**.

(Please note that have we have used the information that the carbonate ion has a 2-charge to find the charge on the iron ions. The total charges must add to zero.)

ATOMIC WEIGHTS

As the chemists of the eighteenth and nineteenth centuries painstakingly sought information about the compositions of compounds and tried to systematise their knowledge, it became apparent that each element has a character mass relative to every other element. Although these early scientists did not have the experimental means to measure the mass of each kind of atom, they succeeded in defining a *relative* scale of atomic masses.

An early observation was that carbon and hydrogen have relative atomic masses, also traditionally called **atomic weights, AW,** of approximately 12 and 1, respectively. Thousands of experiments on the compositions of compounds have resulted in the establishment of a scale of relative atomic weights based on the **atomic mass unit** (**amu**), which is defined as *exactly* $\frac{1}{12}$ *of the mass of an atom of a particular kind of carbon atom, called carbon-12*.

On this scale, the atomic weight of hydrogen (H) is 1.00794 amu, that of sodium (Na) is 22.989768 amu, and that of magnesium (Mg) is 24.3050 amu. This tells us that Na atoms have nearly 23 times the mass of H atoms, while Mg atoms are about 24 times heavier than H atoms.

When you need values of atomic weights, consult the periodic table or the alphabetical listing of elements.

Note: (The term "atomic weight" is widely accepted because of its traditional use although it is properly a mass rather than a weight. "Atomic mass" is often used.)

The Mole

Even the smallest bit of matter that can be handled reliably contains an enormous number of atoms. So we must deal with large numbers of atoms in any real situation, and some unit for conveniently describing a large number of atoms is desirable. The idea of using a unit to describe a particular number (amount) of objects has been around for a long time. You are already familiar with the dozen (12 items) and the gross (144 items).

The SI unit for amount is the **mole,** abbreviated mol. It is *defined* as the amount of substance that contains as many entities (atoms, molecules, or other particles) as there are atoms in exactly 0.012 kg of pure carbon-12 atoms. Many experiments have refined the number and the currently accepted value is

$$1 \text{ mole} = 6.0221367 \times 10^{23} \text{ particles}$$

This number, often rounded off to 6.022×10^{23}, is called **Avogadro's number** in honour of Amedeo Avogadro (1776-1856), whose contributions to chemistry are manifold.

According to its definition, the mole unit refers to a fixed number of entities, whose identities must be specified. Just as we speak of a dozen eggs or a pair of aces, we refer to a mole of atoms or a mole of molecules (or a mole of ions, electrons, or other particles). We could even think about a mole of eggs, although the size of the required carton staggers the imagination! Helium exists as discrete He atoms, so one mole of helium consists of 6.022×10^{23} He *atoms*. Hydrogen commonly exists are diatomic (two-atom) molecules, so one mole of hydrogen is 6.022×10^{23} H_2 *molecules* and $2(6.022 \times 10^{23})$ H atoms.

Every kind of atom, molecule, or ion has a definite characteristic mass. It follows that one mole of a given pure substance also has a definite mass, regardless of the source of the sample. This idea is of central importance in many calculations throughout the study of chemistry and the related sciences.

Because the mole is defined as the number of atoms in 0.012 kg (or 12 grams) of carbon-12, and the atomic mass unit is defined as $\frac{1}{12}$ of the mass of a carbon-12 atom, the following convenient relationship is true.

> The mass of one mole of atoms of a pure element in grams is numerically equal to the atomic weight of that element in amu. This is also called the **molar mass** of the element: its units are grams/mole.

For instance, if you obtain a pure sample of the metallic element (Ti) whose atomic weight is 47.88 amu, and measure out 47.88 grams of it, you will have one mole, or 6.022×10^{23} titanium atoms.

The symbol for an element can (1) identify the element, (2) represent one atom of the element, or (3) represent one mole of atoms of the element. The last interpretation will be extremely useful in calculations in the next chapter.

A quantity of a substance may be expressed in a variety of ways. For example, consider a dozen eggs and 55.847 grams of iron, or one mole of iron. We can express the amount of eggs or iron present in any of several different units. We can then construct unit factors to relate an amount of the substance expressed in one kind of unit to the same amount expressed in another unit.

Unit Factors for Eggs	**Unit Factors for Iron**
$\dfrac{12 \text{ eggs}}{1 \text{ doz eggs}}$	$\dfrac{6.022\times10^{23} \text{ Fe atoms}}{1 \text{ mol Fe atoms}}$
$\dfrac{12 \text{ eggs}}{24 \text{ ounces of eggs}}$	$\dfrac{6.022\times10^{23} \text{ Fe atoms}}{55.847 \times \text{g Fe}}$
and so on	and so on

Table 8: Mass of One of Atoms of Some Common Elements

Element	A Sample with a Mass of	Contains
Carbon	12.011 g C	6.022×10^{23} C atoms or 1 mole of C atoms
Titanium	47.88 g Ti	6.022×10^{23} Ti atoms or 1 mole of Ti atoms
Gold	196.96654 g Au	6.022×10^{23} Au atoms or 1 mole of Au atoms
Hydrogen	1.00794 g H_2	6.022×10^{23} H atoms or 1 mole of H atoms (3.011×10^{23} H_2 molecules or 1/2 mole of H_2 molecules
Sulfur	32.066 g S_8	6.022×10^{23} S atoms or 1 mole of S atoms (0.7528×10^{23} S_8 molecules or 1/8 mole of S_8 molecules)

As Table 8 suggests, the concept of a mole as applied to atoms is especially useful. It provides a convenient basis for comparing the masses of equal numbers of atoms of different elements.

The relationship between the mass of a sample of an element and the number of moles of atoms in the sample is illustrated in the next example.

Example – 14

How many moles of atoms does 245.2 g of iron metal contain?

Critical thinking

The atomic weight of iron is 55.85 amu. This tells us that the molar mass of iron is 55.85 g/mol, or that one mole of iron atoms is 55.85 grams of iron. We can express this as either of two unit factors:

$$\frac{1 \text{ mol Fe atoms}}{55.85 \text{ g Fe}} \quad \text{or} \quad \frac{55.85 \text{ g Fe}}{1 \text{ mol Fe atoms}}$$

Because one mole of iron has a mass of 55.85 g, we expect that 245.2 g will be a fairly small number of moles (greater than one, but less than ten).

Solution: $\underline{?}$ mole Fe atoms = $245.2 \text{ g Fe} \times \dfrac{1 \text{ mol Fe atoms}}{55.85 \text{ g Fe}} = 4.390$ mol Fe atoms

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Once the number of moles of atoms of an element is known, the number of atoms in the sample can be calculated, as next example illustrates.

Example – 15

How many atoms are contained in 4.390 moles of iron atoms?

Critical thinking

One mole of atoms of an element contains Avogadro's number of atoms, or 6.022×10^{23} atoms. This lets us generate the two unit facts.

$$\frac{6.022 \times 10^{23} \text{ atoms}}{1 \text{ mol atoms}} \quad \text{and} \quad \frac{1 \text{ mol atoms}}{6.022 \times 10^{23} \text{ atoms}}$$

Solution: $\underline{?}$ mole Fe atoms = $4.390 \text{ mol Fe atoms} \times \dfrac{6.022 \times 10^{23} \text{ atoms}}{1 \text{ mol atoms}} = 2.644 \times 10^{24}$ Fe atoms.

If we know the atomic weight of an element on the carbon-12 scale, we can use the mole concept and Avogadro's number to calculate the *average* mass of one atom of that element in grams (or any other mass unit we choose).

Example – 16

Calculate the mass of one iron atom in grams.

Critical thinking

We expect that the mass of a single atom in grams would be a very small number. We know that one mole of Fe atoms has a mass of 55.85 g and contains 6.022×10^{23} Fe atoms. We use this information to generate unit factors to carry out the desired conversion.

Solution: $$\frac{\underline{?}\text{ g Fe}}{\text{Fe atom}} = \frac{55.85\text{ g Fe}}{1\text{ mol Fe atoms}} \times \frac{1\text{ mol Fe atoms}}{6.022\times10^{23}\text{ Fe atoms}} = 9.274 \times 10^{-23}\text{ g Fe/Fe atom}$$

Thus, we see that the mass of one Fe atom is only 9.274×10^{-23} g.

FORMULA WEIGHTS, MOLECULAR WEIGHTS, AND MOLES

The **formula weight (FW)** of a substance is the sum of the atomic weights (AW) of the elements in the formula, each taken the number of times the element occurs. Hence a formula weight gives the mass of one formula unit in amu.

Formula weights, like the atomic weight on which they are based, are relative masses. The formula weight for sodium hydroxide, NaOH, (rounded off to the nearest 0.01 amu) is found as follows.

No. of Atoms of Stated Kind			**× Mass of One Atom**	**= Mass Due to Element**
1 × Na	=	1	× 23.00 amu	= 23.00 amu of Na
1 × H	=	1	× 1.01 amu	= 1.01 amu of H
1 × O	=	1	× 16.00 amu	= 16.00 amu of O
			Formula weight of NaOH	= 40.01 amu

The term "formula weight" is correctly used for either ionic or molecular substances. When we refer specifically to molecular (nonionic) substances, i.e., substances that exist as discrete molecules, we often substitute the term **molecular weight (MW)**.

Example – 17

Calculate the formula weight (molecular weight) of 2,4,6-trinitrotoluene (TNT), $C_7H_5(NO_2)_3$, using the precisely known values for atomic weights given in the International Table of Atomic Weights in any standard periodic table.

Critical thinking

We add the atomic weights of the elements in the formula, each multiplied by the number of times the element occurs. Because the least precisely known atomic weight (12.011 amu for C) is known to three significant figures past the decimal point, the result is shown to only that number of significant figures.

Solution:	**No. of Atoms of Stated Kind**		**× Mass of One Atom**	**= Mass Due to Element**
	7 × C =	7	× 12.011 amu	= 84.077 amu of C
	5 × H =	5	× 1.00794 amu	= 5.03970 amu of H
	3 × N =	3	× 14.00674 amu	= 42.02022 amu of N
	6 × O =	6	× 15.9994 amu	= 95.9964 amu of O
	Formula weight (molecular weight) of 2,4, 6-trinitrotoluene (TNT)			= 227.133 amu

The amount of substance that contains the mass in grams numerically equal to its formula weight in amu contains 6.022×10^{23} formula unit, or *one mole* of the substance. This is sometimes called the **molar mass** of the substance. Molar mass is *numerically equal* to the formula weight of the substance (the atomic weight for atoms of elements), and has the units grams/mole.

One mole of sodium hydroxide is 40.01 g of NaOH, and one mole TNT is 227.133 g of $C_7H_5(NO_2)_3$. One mole of any molecular substance contains 6.022×10^{23} molecules of the substance, as Table-9 illustrates.

Table 9: One Mole of Some Common Molecular Substances

Substance	**Molecular Weight**	**A Sample with a Mass of**	**Contains**
hydrogen	2.016	2.016g H_2	6.022×10^{23} H_2 molecules or 1 mol of H_2 molecules (contains $2 \times 6.022 \times 10^{23}$ H atoms or 2 mol of H atoms)
oxygen	32.00	32.00g O_2	6.022×10^{23} O_2 molecules or 1 mol of O_2 molecules (contains $2 \times 6.022 \times 10^{23}$ O atoms or 2 mol of O atoms)
methane	16.04	16.04 g CH_4	6.022×10^{23} CH_4 molecules or 1 mol of CH_4 molecules (contains 6.022×10^{23} C atoms $4 \times 6.022 \times 10^{23}$ H atoms)
2,4,6-trinitroto-luene (TNT)	227.13	227.13 g $C_7H_5(NO_2)_3$	6.022×10^{23} $C_7H_5(NO_2)_3$ molecules or 1 mol of $C_7H_5(NO_2)_3$ molecules

The formula unit (molecule) of oxalic acid is $(COOH)_2$ (FW = 90.04 amu; molar mass = 90.04 g/mol). However, when oxalic acid is obtained by crystallization from a water solution, two molecules of water are present for each molecule of oxalic acid, even though it appears dry. The formula of this **hydrate** is $(COOH)_2 \cdot 2H_2O$ (FW = 126.06 amu; molar mass = 126.06 g/mol). The dot shows that the crystals contain two H_2O molecules per

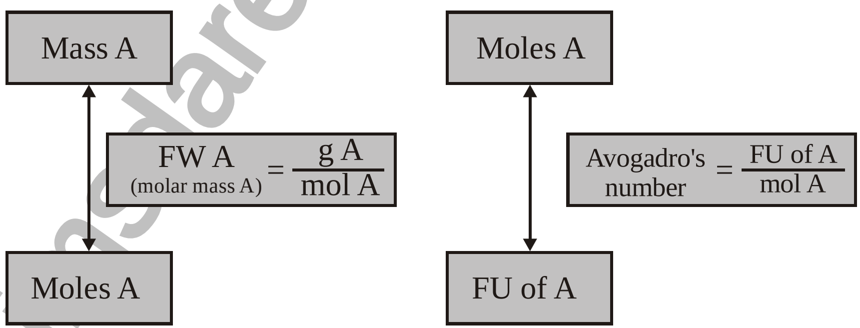
$(COOH)_2$ molecule. The water can be driven out of the crystals by heating to leave **anhydrous** oxalic acid, $(COOH)_2$. Anhydrous means “without water.” Copper(II) sulphate, an *ionic* compound, shows similar behaviour. Anhydrous copper(II) sulphate ($CuSO_4$; FW = 159.60 amu; molar mass = 159.60 g/mol) is almost white. Hydrated copper(II) sulphate ($CuSO_4 . 5H_2O$: FW = 249.68 amu; molar mass = 249.68 g/mol) is deep blue.

Because there are no simple NaCl molecules at ordinary temperature, it is inappropriate to refer to the “molecular weight” of NaCl or any ionic compound. One mole of an ionic compound contains 6.022×10^{23} formula units (FU) of the substance. Recall that one formula unit of sodium chloride consists of one sodium ion, Na^+, and one chloride ion, Cl^-. One mole, or 58.44 grams, of NaCl contains 6.022×10^{23} Na^+ ions and 6.022×10^{23} Cl^- ions. See Table 10.

Table 10: One Mole of Some Ionic Compounds

Compound	Formula Weight	A Sample with a Mass of 1 mole	Contains
sodium chloride	58.44	58.44 g NaCl	6.022×10^{23} Na^+ ions or 1 mole of Na^+ ions 6.022×10^{23} Cl^- ions or 1 mole of Cl^- ions
calcium chloride	111.0	111.0 g $CaCl_2$	6.022×10^{23} Ca^{2+} ions or 1 mole of Ca^{2+} ions $2(6.022 \times 10^{23})$ Cl^- ions or 2 moles of Cl^- ions
aluminum sulphate	342.1	342.1 g $Al_2(SO_4)_3$	$2(6.022 \times 10^{23})$ Al^{3+} ions or 2 moles of Al^{3+} ions $3(6.022 \times 10^{23})$ SO_4^{2-} ions or 3 moles of SO_4^{2-} ions

The mole concept, together with Avogadro’s number provides important connections among the extensive properties mass of substance, number of moles of substance, and number of molecules or ions. These are sammarised as follows.



The following examples show the relations between numbers of molecules, atoms, or formula units and their masses.

Example – 18

What is the mass in grams of 10.0 billion SO_2 molecules?

Critical thinking

One mole of SO_2 contains 6.02×10^{23} SO_2 molecules and has a mass of 64.1 grams.

Solution: $$\underline{?}\text{ g } SO_2 = 10.0 \times 10^9 \text{ } SO_2 \text{ molecules} \times \frac{64.1 \text{ g } SO_2}{6.02 \times 10^{23} \text{ } SO_2 \text{ molecules}}$$

$$= 1.06 \times 10^{-12} \text{ g } SO_2$$

i.e. Ten billion SO_2 molecules have a mass of only 0. 0000000000106 gram.

Example – 19

How many (a) moles of O_2, (b) O_2 molecules, and (c) O atoms are contained in 40.0 grams of oxygen gas (dioxygen) at 25ºC?

Critical thinking

We construct the needed unit factors from the following equalities: (a) the mass of one mole of O_2 is 32.0 g (molar mass O_2 = 32.0 g/mol); (b) one mole of O_2 contains 6.02×10^{23} O_2 molecules: (c) one O_2 molecule contains two O atoms.

Solution: One mole of O_2 contains 6.02×10^{23} O_2 molecules, and its mass is 32.0 g.

(a) $$\underline{?}\text{ mol } O_2 = 40.0 \text{ g } O_2 \times \frac{1 \text{ mol } O_2}{32.0\text{g } O_2} = 1.25 \text{ mol } O_2$$

(b) $$\underline{?}\text{ } O_2 \text{ molcules} = 40.0 \text{ g } O_2 \times \frac{6.02 \times 10^{23} \text{ } O_2 \text{ molecules}}{32.0 \text{ g } O_2} = 7.52 \times 10^{23} \text{ molecules}$$

or, we can use the number of moles of O_2 calculated in part (a) to find the number of O_2 molecules.

$$O_2 \text{ molecules} = 1.2 \text{ mol } O_2 \times \frac{6.02 \times 10^{23} \text{ } O_2 \text{ molecules}}{1 \text{ mol } O_2} \times 7.52 \times 10^{23} O_2 \text{ molecules}$$

(c) $$\underline{?}\text{ O atoms} = 40.0 \text{ g } O_2 \times \frac{6.02 \times 10^{23} \text{ } O_2 \text{ molecules}}{32.0 \text{ g } O_2} \times \frac{2 \text{ O atoms}}{1 \text{ } O_2 \text{ molecule}}$$

$$= 1.50 \times 10^{24} \text{ O atoms.}$$

Example – 20

Calculate the number of hydrogen atoms in 39.6 grams of ammonium sulphate. $(NH_4)_2SO_4$.

Critical thinking

One mole of $(NH_4)_2SO_4$ *is 6.02 × 10*^{23} *formula units (FU) and has a mass of 32 g.*

$$\underset{(NH_4)_2\ SO_4}{\text{g of}} \longrightarrow \underset{(NH_4)_2\ SO_4}{\text{mol of}} \longrightarrow \underset{(NH_4)_2\ SO_4}{\text{Fu of}} \longrightarrow \text{H atoms}$$

Solution: $\underline{?}\,\text{H atoms} = 39.6\text{ g }(NH_4)_2SO_4 \times \dfrac{1\text{ mol }(NH_4)_2SO_4}{132\text{ g }(NH_4)_2SO_4}$

$$\times \frac{6.02\times10^{23}\text{ FU }(NH_4)_2SO_4}{1\text{ mol }(NH_4)_2SO_4} \times \frac{8\text{ H atoms}}{1\text{ FU }(NH_4)_2SO_4}$$

$= 1.44 \times 10^{24}$ H atoms

The term "millimole" (mmol) is useful in laboratory work. As the prefix indicates, one **mmol** is 1/1000 of a mole. Small masses are frequently expressed in milligrams (mg) rather than grams. The relation between millimoles and milligrams is the same as that between moles and grams (Table 11).

Table - 11: ***comparison of moles and millimoles***

Compound	1 Mole	1 Millimole
NaOH	40.0 g	40.0 mg or 0.0400 g
H_3PO_4	98.1 g	98.4 g mg or 0.0981 g
SO_2	64.1 g	64.1 mg or 0.0641 g
C_3H_8	44.1 g	44.1 mg or 0.0441 g

Example – 21

Calculate the number of millimoles of sulphuric acid in 0.147 gram of H_2SO_4.

Critical thinking

1 mol H_2SO_4 *= 98.1 g* H_2SO_4*; 1 mmol* H_2SO_4 *= 98.1 mg* H_2SO_4*. We can use these equalities to solve this problem by either of two methods. Method 1: Express formula weight in g/mmol, then convert g* H_2SO_4 *to mmol* H_2SO_4*. Method 2: Convert g* H_2SO_4 *to mg* H_2SO_4*, then use the unit factor mg/mmol to convert to mmol* H_2SO_4*.*

Solution: Method 1: $\underline{?}\text{ mmol }H_2SO_4 = 0.147\text{ g }H_2SO_4 \times \dfrac{1\text{ mmol }H_2SO_4}{0.0981\text{ g }H_2SO_4} = 1.50\text{ mmol }H_2SO_4$

Method 2: Using 0.147 g H_2SO_4 = 147 mg H_2SO_4, we have,

$$\underline{?}\text{ mmol }H_2SO_4 = 147\text{ mg }H_2SO_4 \times \frac{1\text{ mmol }H_2SO_4}{98.1\text{ mg }H_2SO_4} = 1.50\text{ mmol }H_2SO_4$$

PERCENT COMPOSITION AND FORMULAS OF COMPOUNDS

If the formula of a compound is known, its chemical composition can be expressed as the mass percent of each element in the compound. For example, one carbon dioxide molecule, CO_2, contains one C atom and two O atoms. Percentage is the part divided by the whole times 100 percent (or simply parts per 100), so we can represent the percent composition of carbon dioxide as follows.

$$\% \text{ C} = \frac{\text{mass of C}}{\text{mass of } CO_2} \times 100\% = \frac{\text{AW of C}}{\text{MW of } CO_2} \times 100\% = \frac{12.0 \text{ amu}}{44.0 \text{ amu}} \times 100\% = 27.3\%$$

$$\% \text{ O} = \frac{\text{mass of O}}{\text{mass of } CO_2} \times 100\% = \frac{2 \times \text{AW of O}}{\text{MW of } CO_2} \times 100\% = \frac{2(16.0 \text{ amu})}{44.0 \text{ amu}} \times 100\% = 72.7\ \%\text{O}$$

One *mole* of CO_2(44.0 g) contains one *mole* of C atoms (12.0 g) and two *moles* of O atoms (32.0 g). Therefore, we could have used these masses in the preceding calculation. These numbers are the same as the ones used—only the units are different. In Example 22 we shall base our calculation on one *mole* rather than one *molecule*.

Example – 22

Calculate the percent composition of HNO_3 by mass.

Critical thinking

We first calculate the mass of one mole. Then we express the mass of each element as a percent of the total.

Solution: The molar mass of HNO_3 is calculated first.

No. of Mol of Atoms		× Mass of One Mol of Atoms	= Mass Due to Element
1 × H =	1	× 1.0 g	= 1.0 g of H
1 × N =	1	× 14.0 g	= 14.0 g of N
3 × O =	3	× 16.0 g	= 48.0 g or O
		Mass of 1 mol of HNO_3	= 63.0 g

Now, its percent composition is

$$\% \text{ H} = \frac{\text{mass of H}}{\text{mass of } HNO_3} \times 100\% = \frac{1.0 \text{ g}}{63.0 \text{ g}} \times 100\% = 1.6\% \text{ H}$$

$$\% \text{ N} = \frac{\text{mass of N}}{\text{mass of } HNO_3} \times 100\% = \frac{14.0 \text{ g}}{63.0 \text{ g}} \times 100\% = 22.2\% \text{ N}$$

$$\% \text{ O} = \frac{\text{mass of O}}{\text{mass of } HNO_3} \times 100\% = \frac{48.0 \text{ g}}{63.0 \text{ g}} \times 100\% = 76.2\% \text{ O}$$

Nitric acid is 1.6% H, 22.2% N, and 76.2% O by mass. All samples of pure HNO_3 have this composition, according to the Law of Definite Proportions.

DERIVATION OF FORMULAS FROM ELEMENTAL COMPOSITION

Each year thousands of new compounds are made in laboratories or discovered in nature. One of the first steps in characterising a new compound is the determination of its present composition. A *qualitative* analysis is performed to determine *which* elements are present in the compound. then a *quantitative* analysis is performed to determine the *amount* of each element.

Once the percent composition of a compound (or its elemental composition by mass) is known, the simplest formula can be determined. The **simplest** or **empirical formula** for a compound is the smallest whole-number ratio of atoms present. For molecular compounds the **molecular formula** indicates the *actual* number of atoms present in a molecule of the compound. It may be the same as the simplest formula or else some whole-number multiple of it. For example, the simplest and molecular formulas for water are both H_2O. However, for hydrogen peroxide, they are HO and H_2O_2, respectively.

Example – 23

Compounds containing sulphur and oxygen are serious air pollutants; they represent the major cause of acid rain. Analysis of a sample of a pure compound reveals that it contains 50.1% sulphur and 49.9% oxygen by mass. What is the simplest formula of the compound?

Critical thinking:

One mole of atoms of any element is 6.022×10^{23} atoms, so the ratio of moles of atoms in any sample of a compound is the same as the ratio of atoms in that compound. This calculation is carried out in two steps. Step 1: Let's consider 100.0 grams of compound which contains 50.1 grams of S and 49.9 grams of O. We calculate the number of moles of atoms of each. Step 2: We then obtain a whole-number ratio between these numbers that gives the ration of atoms in the sample, and hence in the simplest formula for the compound.

Solution: **Step 1:**

$$\underline{?}\text{ mol S atoms} = 50.1\text{ g S} \times \frac{1\text{ mol S atoms}}{32.1\text{ g S}} = 1.56\text{ mol S atoms}$$

$$\underline{?}\text{ mol O atoms} = 49.9\text{ g O} \times \frac{1\text{ mol O atoms}}{16.0\text{ g O}} = 3.12\text{ mol O atoms}$$

Step 2: Now we know that 100.0 grams of the compound contains 1.56 moles of S atoms and 3.12 moles of O atoms. We obtain a whole-number ratio between these numbers that gives the ratio of atoms in the simplest formula.

$$\frac{1.56}{1.56} = 1.00\text{ S}$$

$$\frac{3.12}{1.56} = 2.00\text{ O}$$

SO_2

A simple and useful way to obtain whole number ratios among several number follows. (a) Divide each number by the smallest, and then, (b) if necessary multiply all of the resulting numbers by the smallest whole number that will eliminate fractions.

The solution for last example 23 can be set up in tabular form.

Element	Relative Mass of Element	Relative Number of Atom (divide mass by AW)	Divide by Smaller Number	Smallest Whole Number Ratio of Atoms
S	50.1	$\frac{50.1}{32.1} = 1.56$	$\frac{1.56}{1.56} = 1.00$ S	SO_2
O	49.9	$\frac{49.9}{16.0} = 3.12$	$\frac{3.12}{1.56} = 2.00$ O	

This tabular format provides a convenient way to solve simplest-formula problems as the next example illustrates.

Example – 24

A 20.882-gram sample of an ionic compound is found to contain 6.072 grams of Na, 8.474 grams of S, and 6.336 grams of O. What is its simplest formula?

Critical thinking

We reason as in last example 23 calculating the number of moles of each element and the ratio among them. Here we use the tabular format that was introduced above.

Solution:

Element	Relative Mass of Element	Relative Number of Atom (divide mass by AW)	Divide by Smaller Number	Convert Fractions to Whole Number (multiply by integer)	Smallest Whole Number Ratio of Atoms
Na	6.072	$\frac{6.072}{230} = 0.264$	$\frac{0.264}{0.264} = 1.00$	1.00 × = 2 Na	
S	8.474	$\frac{8.474}{32.1} = 0.264$	$\frac{0.264}{0.264} = 1.00$	1.00 × = 2 S	$Na_2S_2O_3$
O	6.336	$\frac{6.336}{16.0} = 0.396$	$\frac{0.396}{0.264} = 1.50$	1.50 × = 3 O	

The ratio of atoms in the simplest formula must *be a whole-number ratio* (by definition). To convert the ratio 1 : 1 : 1.5 to a whole-number ratio, each number in the ratio was multiplied by 2, which gave the simplest formula $Na_2S_2O_3$.

PROBLEM-SOLVING TIP ***Know Common Fractions in Decimal Form***

As Example 24 illustrates, sometimes we must convert a fraction to a whole number by multiplication by the correct integer. But we must first recognise which fraction is represented by a nonzero part of a number. The decimal equivalents of the following fractions may be useful.

Fraction	Decimal Equivalent (to 2 places)	To Convert to Integer, Multiply by
1/2	0.50	2
1/3	0.33	3
2/3	0.67	3
1/4	0.25	4
3/4	0.75	4
1/5	0.20	5

The fractions 2/5, 3/5/, 4/5 are equal to 0.40, 0.60, and 0.80, respectively; these should be multiplied by 5.

When we use the procedure given in this section, we often obtain numbers such as 0.99 and 1.52. Because there is usually some error in results obtained by analysis of samples (as well as roundoff errors), we would interpret 0.99 as 1.0 and 1.52 as 1.5.

Determination of molecular formulas

Percent composition data yield only simplest formulas. To determine the molecular formula for a molecular compound, *both* its simplest formula and its molecular weight must be known.

Millions of compounds are composed of carbon, hydrogen, and oxygen. Analyses for C and H can be performed in a C—H combustion system. An accurately known mass of a compound is burned in a furnace in a stream of oxygen. The carbon and hydrogen in the sample are converted to carbon dioxide and water vapour, respectively. The resulting increases in masses of the CO_2 and H_2O absorbers can then be related to the masses and percentages of carbon and hydrogen in the original sample.

Example – 25

Hydrocarbons are organic compounds composed entirely of hydrogen and carbon. A 0.1647-gram sample of a pure hydrocarbon was burned in a C—H combustion train to produce 0.4931 gram of CO_2 and 0.2691 gram of H_2O. Determine the masses of C and H in the sample and the percentages of these elements in this hydrocarbon.

Critical thinking

Step 1:
We use the observed mass of CO_2, 0.4931 grams, to determine the mass of carbon in the original sample. There is one mole of carbon atoms, 12.01 grams, in each mole of CO_2. 44.01 grams; we use this information to construct the unit factor

$$\frac{12.01\text{ g C}}{44.01\text{ g CO}_2}$$

Step 2:
Likewise, we can use the observed mass of H_2O, 002691 grams, to calculate the amount of hydrogen in the original sample. We use the fact that there are two moles of hydrogen atom, 2.016 grams, in each mole of H_2O, 18.02 grams, to construct the unit factor

$$\frac{2.016\text{ g H}}{18.02\text{ g H}_2\text{O}}$$

Step 3:
Then we calculate the percentages by mass of each element in turn, using the relationship

$$\%\text{ element} = \frac{\text{g element}}{\text{g sample}} \times 100\%$$

Solution: Step 1: $$\underline{?}\text{ g C} = 0.4931\text{ g CO}_2 \times \frac{12.01\text{ g C}}{44.01\text{ g CO}_2} = 0.1346\text{ g C}$$

Step 2: $$\underline{?}\text{ g H} = 0.2691\text{ g H}_2\text{O} \times \frac{2.016\text{ g H}}{18.02\text{ g H}_2\text{O}} = 0.03010\text{ g H}$$

Step 3: $$\%\text{ C} = \frac{0.1346\text{ g C}}{0.1647\text{ g sample}} \times 100\% = 81.72\%\text{ C}$$

$$\%\text{ H} = \frac{0.03010\text{ g H}}{0.1647\text{ g sample}} \times 100\% = 18.28\%\text{ H}$$

Total = 100.00%

When the compound to be analysed contains oxygen, the calculation of the amount or percentage of oxygen in the sample is somewhat different. Part of the oxygen that goes to form CO_2 and H_2O comes from the sample and part comes from the oxygen stream supplied. Therefore we cannot directly determine the amount of oxygen already in the sample. The approach is to analyse as we did in Example 25 for all elements *except* oxygen. Then we subtract the sum of their masses from the mass of the original sample to obtain the mass of oxygen. The next example illustrates such a calculation.

Example – 26

A 0.1014-gram sample of purified glucose was burned in a C—H combustion train to produce 0.1486 gram of CO_2 and 0.0609 gram of H_2O. An elemental analysis showed that glucose contains only carbon, hydrogen, and oxygen. Determine the masses of C, H, and O in the sample and the percentages of these elements in glucose.

Critical thinking

Steps 1 and 2: *We first calculate the masses of carbon and hydrogen as we did in Example 25.*

Step 3: *The rest of the sample must be oxygen because glucose has been shown to contain only C, H, and O. So we subtract the masses of C and H from the total mass of sample.*

Step 4: *Then we calculate the percentage by mass for each element.*

Solution: **Step 1:** $\underline{?}\text{ g C} = 0.1486\text{ g }CO_2 \times \dfrac{12.01\text{ g C}}{44.01\text{ g }CO_2} = 0.04055\text{ g C}$

Step 2: $\underline{?}\text{ g H} = 0.0609\text{ g }H_2O \times \dfrac{2.016\text{ g H}}{18.02\text{ g }H_2O} = 0.00681\text{ g H}$

Step 3: $\underline{?}\text{ g O} = 0.1014\text{ g sample} - [0.04055\text{ g C} + 0.00681\text{ g H}] = 0.0540\text{ g O}$

Step 4: Now we can calculate the percentages by mass for each element:

$$\%\text{ C} = \frac{0.04055\text{ g C}}{0.1014\text{ g}} \times 100\% = 39.99\%\text{ C}$$

$$\%\text{ H} = \frac{0.00681\text{ g H}}{0.1014\text{ g}} \times 100\% = 6.72\%\text{ H}$$

$$\%\text{ O} = \frac{0.0540\text{ g O}}{0.1014\text{ g}} \times 100\% = 53.2\%\text{ O}$$

Total = 99.9%

For many compounds the molecular formula is a multiple of the simplest formula. Consider butane, C_4H_{10}. The simplest formula for butane is C_2H_5, but the molecular formula contains twice as many atoms; i.e., $2 \times (C_2H_5) = C_4H_{10}$. Benzene, C_6H_6, is another example. The simplest formula for benzene is CH, but the molecular formula contains six times as many atoms; i.e., $6 \times (CH) = C_6H_6$.

The molecular formula for a compound is either the same as , or an integer multiple of, the simplest formula.
molecular formula = n × simplest formula

So we can write

$$\text{molecular weight} = n \times \text{simple formula weight}$$

$$\Rightarrow \quad n = \frac{\text{molecular weight}}{\text{simple t-formula weight}}$$

The molecular formula is then obtained by multiplying the simplest formula by the integer, n.

Example – 27

In Example 26 we found the elemental composition of glucose. Other experiments show that its molecular weight is approximately 180 amu. Determine the simplest formula and the molecular formula of glucose.

Critical thinking

Step 1: *We first use the masses of C, H, and O found in Example 26 to determine the simplest formula.*

Step 2: *We can use the simplest formula to calculate the simplest-formula weight. Because the molecular weight of glucose is known (approximately 180 amu), we can determine the molecular formula by dividing the molecular weight by the simplest-formula weight.*

$$n = \frac{\text{molecular weight}}{\text{simpest-formula weight}}$$

The molecular weight is n times the simplest-formula weight, so the molecular formula of glucose is n times the simplest formula.

Solution: **Step 1:**

Element	Mass of Element	Moles of Element (divide mass by AW)	Divide by Smallest	Smallest Whole Number Ratioof Atoms
C	0.04055 g	$\frac{0.04055}{12.01} = 0.003376$ mol	$\frac{0.003376}{0.003376} = 1.00$ C	
H	0.00681 g	$\frac{0.00681}{1.008} = 0.00676$ mol	$\frac{0.00676}{0.003376} = 2.00$ H	CH_2O
O	0.0540 g	$\frac{0.0540}{16.00} = 0.00338$ mol	$\frac{0.00338}{0.003376} = 1.00$ O	

Step 2:

The simplest formula is CH_2O, which has a formula weight of 30.02 amu. Because the molecular weight of glucose is approximately 180 amu, we can determine the molecular formula by dividing the molecular weight by the simplest-formula weight.

$$n = \frac{180 \text{ amu}}{30.02} = 6.00$$

The molecular weight is six times the simplest-formula weight, $6 \times (CH_2O) = C_6H_{12}O_6$, so the molecular formula of glucose is $C_6H_{12}O_6$.

As we shall see when we discuss the composition of compounds in some detail, two (and sometimes more) elements may form more than one compound. The **Law of Multiple Proportions** summarises many experiments on such compounds. It is usually stated: When two elements, A and B, form more than one compound, the ratio of the masses of element B that combine with a given mole element A in each of the compounds can be expressed by small whole numbers. Water, H_2O, and hydrogen peroxide. H_2O_2, provide an example. The ratio of masses of oxygen that combine with a given mass of hydrogen is 1 : 2 in H_2O and H_2O_2. Many similar examples, such as CO and CO_2 (1 : 2 ratio) and SO_2 and SO_3 (2 : 3 ratio), are known. The Law of Multiple Proportions had been recognised from studies of elemental composition before the time of Dalton. It provided additional support for his atomic theory.

Example – 28

What is the ratio of the masses of oxygen that are combined with 1.00 gram of nitrogen in the compounds N_2O_3 and NO?

Critical thinking

First we calculate the mass of O that combines with one gram of N in each compound. Then we determine the ratio of the values $\frac{\text{g O}}{\text{g N}}$ *for the two compounds.*

Solution: In N_2O_3: $\frac{\text{? g O}}{\text{g N}} = \frac{48.0 \text{ g O}}{28.0 \text{ g N}} = 1.71 \text{ g O/g N}$

In NO: $\frac{\text{? g O}}{\text{g N}} = \frac{16.0 \text{ g O}}{14.0 \text{ g N}} = 1.14 \text{ g O/g N}$

$$\text{The ratio is } \left\{ \begin{array}{l} \frac{\text{g O}}{\text{g N}} \text{ (in } N_2O_3) \\ \frac{\text{g O}}{\text{g N}} \text{ (in NO)} \end{array} \right. \rightarrow \frac{1.71 \text{ g O/g N}}{1.14 \text{ g O/g N}} = \frac{1.5}{1.0} = \frac{3}{2}$$

We see that the ratio is 3 mass units of O (in N_2O_3) to 2 mass units of O (in NO).

SOME OTHER INTERPRETATIONS OF CHEMICAL FORMULAS

Once we master the mole concept and the meaning of chemical formulas, we can use them in many, other ways. The examples in this section illustrate a few additional kinds of information we can get from a chemical formula and the mole concept.

Example – 29

What mass of chromium is contained in 35.8 grams of $(NH_4)_2Cr_2O_7$?

Critical thinking

Let us first solve the problem in several steps.

Step 1: *The formula tells us that each mole of $(NH_4)_2Cr_2O_7$ contains two moles of Cr atoms, so we first find the number of moles of $(NH_4)_2Cr_2O_7$, using the unit factor,*

$$\frac{1 \text{ mol } (NH_4)_3Cr_2O_7}{252.0 \text{ g } (NH_4)_2Cr_2O_7}$$

Step 2: *Then we convert the number of moles of $(NH_4)_2Cr_2O_7$ iron the number of moles of Cr atoms it contains, using the unit factor,*

$$\frac{2 \text{ mol Cr atoms}}{1 \text{ mol } (NH_4)_2Cr_2O_7}$$

Step 3: *We then use the atomic weight of Cr to convert the number of moles of chromium atoms to mass of chromium.*

$$\text{Mass } (NH_4)_2Cr_2O_7 \rightarrow \text{mol}(NH_4)_2Cr_2O_7 \rightarrow \text{mol Cr} \rightarrow \text{Mass Cr}$$

Solution: Step 1: $\underline{?}$ mol $(NH_4)_2Cr_2O_7$ = 35.8 g $(NH_4)_2Cr_2O_7 \times \dfrac{1 \text{ mol } (NH_4)_2Cr_2O_7}{252.0 \text{ g } (NH_4)_2Cr_2O_7}$

$= 0.142$ mol $(NH_4)_2Cr_2O_7$

Step 2: $\underline{?}$ mol Cr atoms = 0.142 mol $(NH_4)_2Cr_2O_7 \times \dfrac{2 \text{ mol Cr atoms}}{1 \text{ mol } (NH_4)_2Cr_2O_7}$

$= 0.284$ mol Cr atoms

Step 3: $\underline{?}$ g Cr = 0.284 mol Cr atoms $\times \dfrac{52.0 \text{ g Cr}}{1 \text{ mol Cr atoms}} = 14.8$ g Cr

If you understand the reasoning in these conversion. you should be able to solve this problem in a single setup: (try it yourself first and then compare to the solution given below.)

$$\underline{?} \text{ g Cr} = 35.8 \text{ g } (NH_4)_2Cr_2O_7 \times \frac{1 \text{ mol } (NH_4)_2Cr_2O_7}{252.0 \text{ g } (NH_4)_2Cr_2O_7} \times \frac{2 \text{ mol Cr atoms}}{1 \text{ mol } (NH_4)_2Cr_2O_7} \times \frac{52.0 \text{ g Cr}}{1 \text{ mol Cr}} = 14.8 \text{ g Cr}$$

Example – 30

What mass of potassium chlorate, $KClO_3$, would contain 40.0 grams of oxygen?

Critical thinking

The formula $KClO_3$ tells us that each mole of $KClO_3$ contains three moles of oxygen atoms. Each mole of oxygen atoms weighs 16.0 grams. So we can set up the solution to convert:

Mass O $\rightarrow$ mol O $\rightarrow$ mol $KClO_3$ $\rightarrow$ Mass $KClO_3$

Solution: $\underline{?}$ g $KClO_3$ = 40.0 g O $\times \dfrac{1 \text{ mol O atoms}}{16.0 \text{ g O atoms}} \times \dfrac{1 \text{ mol } KClO_3}{3 \text{ mol O atoms}} \times \dfrac{122.6 \text{ g } KClO_3}{1 \text{ mol } KClO_3}$

$= 102$ g $KClO_3$

PROBLEM-SOLVING TIP ***How Do We Know When...***

How do we know when to represent oxygen as O and when as O_2.? A *compound* that contains oxygen *does not* contain O_2 molecules. So we solve problems such as Examples 30 and 31 in terms of moles of O atoms. Thus, we must use the formula weight for O, which is $\dfrac{16.0\, g \text{ O atoms}}{1 \text{ mol O atoms}}$. Similar reasoning applies to compounds containing other elements that are polyatomic molecules in *pure elemental form*, such as H_2, Cl_2, or P_4.

Example – 31

(a) What mass of sulphur dioxide, SO_2, would contain the same mass of oxygen as is contained in 33.7 g of arsenic pentoxide, As_2O_5 ?

(b) What mass of calcium chloride $CaCl_2$, would contain the same number of chloride ions as are contained in 48.6 g of sodium chloride, NaCl?

Critical thinking

(a) *We would find explicitly the number of grams of O in 33.7 g of As_2O_5, and then find the mass of SO_2 that contains that same number of grams of O. But this method includes some unnecessary calculation. We need only convert to moles of O (because this is the same amount of O regardless of its environment) and then to SO_2.*

$$\text{Mass } As_2O_5 \rightarrow \text{mol } As_2O_5 \rightarrow \text{mol O atoms} \rightarrow \text{mol } SO_2 \rightarrow \text{Mass } SO_2$$

(b) *Because one mole always consists of the same number (Avogadro's number) of items, we can reason in terms of moles of Cl^- ions:*

$$\text{Mass NaCl} \rightarrow \text{mol NaCl} \rightarrow \text{mol } Cl^- \text{ ions} \rightarrow \text{mol } CaCl_2 \rightarrow \text{Mass } CaCl_2$$

Solution: (a)

$$\underline{?}\ g SO_2 = 33.7 g\ As_2O_5 \times \frac{1 mol As_2O_5}{229.8 g\ As_2O_5} \times \frac{5 mol\ O\ atoms}{1\ mol\ As_2O_5}$$

$$\times \frac{1\ mol\ SO_2}{2 mol\ O\ atoms} \times \frac{64.1 g\ SO_2}{1 mol\ SO_2} = 23.5 g\, SO_2$$

(b)

$$\underline{?}\ g CaCl_2 = 48.6\ NaCl \times \frac{1 mol\ NaCl}{58.4 g\ NaCl} \times \frac{1 mol\ Cl^-}{1 mol\ NaCl}$$

$$\times \frac{1 mol\ CaCl_2}{2\ mol\ Cl^-} \times \frac{111.0 g\ CaCl_2}{1 mol\ CaCl_2} = 46.2 g\ CaCl_2$$

We have already mentioned the existence of hydrates (for example, $(COOH)_2 \cdot 2H_2O$ and $CuSO_4 \cdot 5H_2O$. In such hydrates, two components, water and another compound are present in a definite integer ratio by moles. The following example illustrates how we might find use the formula of a hydrate.

Example – 32

A reaction requires pure anhydrous calcium sulphate, $CaSO_4$. Only an unidentified hydrate of calcium sulphate, $CaSO_4 \cdot xH_2O$, is available.

(a) We heat 67.5 g of the unknown hydrate until all the water has been driven off. The resulting mass of pure $CaSO_4$ is 53.4 g. What is the formula of the hydrate, and what is its formula weight?

(b) Suppose we wish to obtain enough of this hydrate to supply 95.5 grams of $CaSO_4$. How many grams should we weigh out?

Critical thinking

(a) *To find the formula of the hydrate, we must find the value of x in the formula $CaSO_4 \cdot xH_2O$. The mass of water removed from the sample is equal to the difference in the two masses given The value of x is the number of moles of H_2O per mole of $CaSO_4$ in the hydrate.*

(b) *The formula weights of $CaSO_4$, 136.2 g/mol, and of $CaSO_4 \cdot xH_2O$, (136.2 + x18.0 g/mol, allow us to write the conversion factor required for the calculation.*

Solution: (a) $\underline{?}$ g water driven off = 67.5 g $CaSO_4{\bullet}xH_2O$ – 53.4 g $CaSO_4$ = 14.1 g H_2O

$$x = \frac{\underline{?}\,\text{mol } H_2O}{\text{mol } CaSO_4} = \frac{14.1\,\text{g } H_2O}{53.4\,\text{g } CaSO_4} \times \frac{1\,\text{mol } H_2O}{18.0\,\text{g } H2O} \times \frac{136.2\,\text{g } CaSO_4}{1\,\text{mol } CaSO_4} = \frac{2.00\,\text{mol } H_2O}{\text{mol } CaSO_4}$$

Thus, the formula of the hydrate is $CaSO_4 \cdot 2H_2O$. Its formula weight is

$$FW = 1 \times (\text{formula weight } CaSO_4) + 2 \times (\text{formula weight } H_2O)$$

$$= 136.2 \text{ g/mol} + 2(18.0\,\text{g/mol}) = 172.2 \text{ g/mol}$$

(b) The formula weights of $CaSO_4$ (136.2 g/mol) and of $CaSO_4{\bullet}2H_2O$(172.2 g/mol) allow us to write the unit factor

$$\frac{172.2 \text{ g } CaSO_4 \cdot 2H_2O}{136.2\,\text{g } CaSO_4}$$

We use this factor to perform the required conversion:

$$\underline{?}\,\text{g}\,CaSO_4 \cdot 2H_2O = 95.5 \text{ g } CaSO_4 \text{ desired} \times \frac{172.2\,\text{g } CaSO_4 \cdot 2H_2O}{136.2\,\text{g } CaSO_4}$$

$$= 121\,\text{g } CaSO_4 \cdot 2H_2O$$

Many important relationships have been introduced in this chapter. Some of the most important transformations you have seen in sections 1 and 2 are summarised in the following figure:

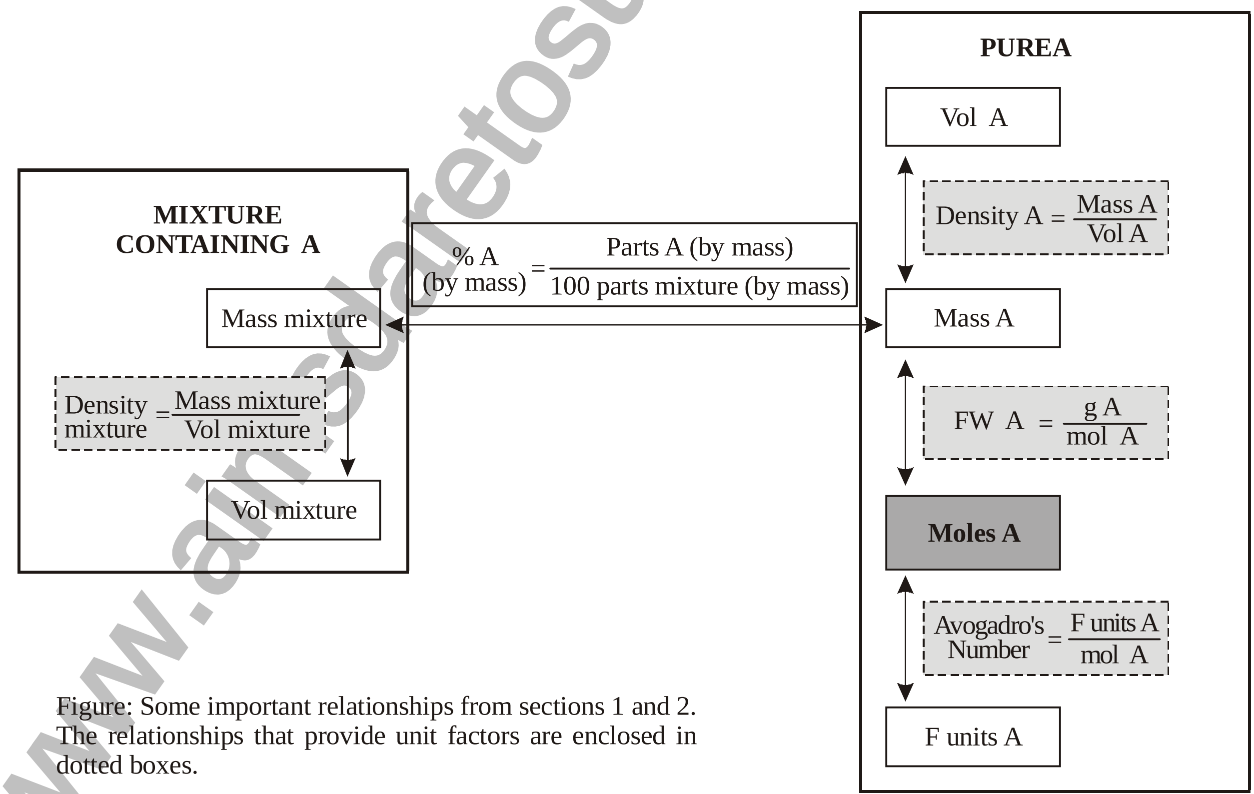


Figure: Some important relationships from sections 1 and 2. The relationships that provide unit factors are enclosed in dotted boxes.

Having finished this section, we are hereby providing you with some good scientific articles from various sources. The idea behind providing you with these articles is to help you develop a broad scientific aptitude in general. We have always believed that proper education is not just about cramming of some results and facts and then reproducing them in the examination hall. It is about going deep into every concept that comes your way and focus on understanding it completely. If this is taken care of, the examination generally takes care of it self. This is the approach that has been maintained throughout this study material. But we expect more for our students than just understanding and applying concepts in solving problems. We want you to develop a broad scientific aptitude on the concepts that you learn and appreciate its various nuances and its application and relation to our daily lives. To help you in developing this broad aptitude, we hereby are providing you with some general scientific articles generally using whatever you have learnt in this chapter so far. We will be providing you with more such articles as you proceed along with your studies. We hope that this should inspire you enough to look for such articles on your own from various sources like the internet and scientific journals. In the process, we hope that you come to appreciate science on your own and develop a flair for the subject you have decided to make your career in!

Article 1: *Science and the Scientific Method*

All live and nonliving entities, from single cells and complex animals to stars and galaxies, undergo change. They emerge, age, die, and continue to change even after death. Science is the careful inquiry into the many changes that occur within us, around us, and throughout our universe.

Many people are interested in and seek explanations for the changes they observe. Some become so interested and fascinated with transformations that they devote their lives to the study of changes. In so doing, they gain a deeper understanding of reality. Science is a vehicle used to study changes; those who pursue this study are called *scientists*.

Some of the observations that we make about the properties and behaviour of the matter around us are *qualitative* (one substance is a shiny solid that conducts electricity; another substance reacts violently with water) and others are *quantitative* (the pressure of a gas is measured to be 1 atmosphere; water is 11.1% hydrogen and 88.9% oxygen by mass). The result of an accurate observation is a *fact*. Often we find patterns in the facts we observe. A *law* is a concise verbal or mathematical statement of a generally observed pattern of behaviour.

After many facts have been observed, it is human nature to try to discover or develop an explanation for the observations (facts). Scientists use their imaginations and reasoning powers to develop speculative explanations called *hypotheses* for their observations. They then perform additional experiments, make further observations to weed out false conjecture, and try to develop better explanations called *theories*. This approach used by scientists to develop, test, and refine explanations for observed changes is called the *scientific method*.

We can understand the scientific method by applying it to an everyday situation such as trying to determine why an air conditioner in a car isn't working properly. Let's assume that one day, while driving to school, you notice that the inside of your car is warmer and more humid than usual. You have observed a change. You remember that while you were stopped at a traffic light yesterday, another student drove into the side of your car. You speculate that the damage from the impact causes your car door to close improperly. If the door does not close properly, the cool air produced by the air conditioner can escape, and the air conditioner becomes less effective. This is your educated guess or hypothesis for the change in temperature and humidity that you observe.

You check the car door and find that it is damaged and doesn't close properly. Your educated guess has passed this additional observation. However, after having your door properly aligned at a body shop, you notice that the heat and humidity within the car are still uncomfortable. Your educated guess was wrong, or at least incomplete. Your tentative conjecture to explain the change in humidity and heat is not supported by reproducible observations or experiments. So you look further. You consider the possibility that the refrigerant level inside the air conditioner may be low, so you take your car to an airconditioning service center. The mechanic corrects the problem, and your air conditioner then works properly. You now feel confident that your "refrigerant explanation" was correct. Your last educated guess (hypothesis) has now been elevated to the level of a theory.

Unknown to you, however, the mechanic had noticed that a piece of paper had blown against the cooling coils of your air conditioner. This paper prevented the air conditioner from working properly. The mechanic simply removed the paper and charged you $95 to "service your air conditioner," and you go through life mistakenly confident about your "refrigerant theory."

A good hypothesis or theory does more than correctly explain reality. *A good hypothesis or theory must also be able to predict reality*. When prehistoric men observed changes, they sometimes created superstitions to explain their observations. For example, when thunderstorms suddenly developed, the angry Thunder God explanation was invoked. Although this explanation provided our early ancestors with an explanation for sudden changes in weather, it did not give them the ability to make correct predictions about thunderstorms, a serious shortcoming. A hypothesis or theory gains status as its predictions are tested and verified.

Sometimes, both scientists and the public misuse the word "theory" to mean either an educated guess or a fact. A theory is neither a fact nor a simple educated guess. Facts are things that we can observe. Correctly made measurements are facts. A simple educated guess is a hypothesis. A theory may have begun as an educated guess, but the educated guess achieves the status of theory only after it survives repeated testing.

For any given phenomenon, it is possible to create many explanations. Only with time, continued observations, and confirmation by others can one explanation dominate all others and become accepted by the majority of scientists. Even so, the accepted *law* or model may still require modifications as more observations are made.

Article 2: *The "size" of Avogadro's Number*

If you think that the value of Avogadro's number, 6×10^{23}, is too large to be useful to anyone but chemists, look up into the sky on a clear night. You may be able to see about 3,000 stars with the naked eye, but the total number of stars swirling around you in the known universe is approximately equal to Avogadro's number. Just think, the known universe contains approximately one mole of stars! You don't have to leave Earth to encounter such large numbers. The water in the Pacific Ocean has a volume of about 6×10^{23} milliliters and a mass of about 6×10^{23} grams.

Avogadro's number is almost incomprehensibly large. For example, if one mole of rupees were given away at the rate of a million rupees per second beginning when the Earth first formed some 4.5 billion years ago, would any remain today? Surprisingly, about three fourths of the original mole of rupees would be left today; it would take about fourteen billion, five hundred million more years to give away the remaining money at one million rupees per second.

Computers can be used to provide another illustration of the magnitude of Avogadro's number. If a computer can count up to one billion in one second, it would take that computer about 20 million years to count up to 6×10^{23}. In contrast, recorded human history goes back only a few thousand years.

The impressively large size of Avogadro's number can give us very important insights into the very small sizes of individual molecules. Suppose one drop of water evaporates in one hour. There are about 20 drops in one milliliter of water, which weighs one gram. So one drop of water is about 0.05 gram of water. How many H_2O molecules evaporate per second?

$$\frac{\underline{?}\ H_2O\ \text{molecules}}{\text{second}} = \frac{0.05\ \text{g}\ H_2O}{1\,\text{hr}} \times \frac{1\,\text{mol}\ H_2O}{18\ \text{g}\ H_2O} \times$$

$$\frac{6\times10^{23}\ H_2O\ \text{molecules}}{1\,\text{mol}\ H_2O} \times \frac{1\,\text{hr}}{60\,\text{min}} \times \frac{1\,\text{min}}{60\,\text{s}}$$

$$= 5\times10^{17}\ H_2O\ \text{molecules/second}$$

$5\times10^{17}\ H_2O$ molecules evaporating per second is five hundred million billion H_2O molecules evaporating per second–a number that is beyond our comprehension! This calculation helps us to recognise that water molecules are incredibly small. There are approximately 1.7×10^{21} water molecules in a single drop of water.

By gaining some appreciation of the vastness of Avogadro's number, we gain a greater appreciation of the extremely tiny volumes occupied by individual atoms, molecules, and ions.

Article 3: *Names of the Elements*

If you were to discover a new element, how would you name it? Throughout history, scientists have answered this question in different ways. Most have chosen to honour a person or place or to describe the new substance.

Until the Middle Ages only nine elements were known: gold, silver, tin, mercury, copper, lead, iron, sulphur, and carbon. The metals' chemical symbols are taken from descriptive Latin names: *aurum* ("yellow"). *Argentum*("shining"), *Stannum* ("dripping" or "easily melted"), *hydrargyrum* ("silvery water"), *cuprum* (Cyprus, where many copper mines were located, *plumbum* (exact meaning unknown–possibly "heavy")_, and *ferrum* (also unknown). Mercury is named after the planet, one reminder that the ancients associated metals with gods and celestial bodies. In turn, both the planet, which moves rapidly across the sky, and the element, which is the only metal that is liquid at room temperature and thus flows rapidly, are named for the fleet god of messengers in Roman mythology. In English, mercury is nicknamed "quicksilver."

Prior to the reforms of Antoine Lavoisier, chemistry was a largely nonquantitative, unsystematic science in which experimenters had little contact with each other. In 1787 Lavoisier published his *Methode de Nomenclature chimique*, which proposed, among other changes, that all new elements be named descriptively. For the next 125 years, most elements were given names that corresponded to their properties. Greek roots were one popular source, as evidenced by hydrogen (*hydros-gen*, "water-producing"), Oxygen (*oksys-gen*, "acid-producing") nitrogen (*nitron-gen* "soda-producing"), bromine (*bromos*, "stink"), and argon (*a-er-gon*, "No reaction"). The discoverers of argon, Ramsay and Rayleigh, originally proposed the name *aeron* (from *aer* or air) but critics thought it was too close to the biblical name Aaron! Latin roots such as radius ("ray") were also used (radium and radon are both naturally radioactive elements that emit "rays"). Colour was often the determining property, especially after the invention of the spectroscope in 1859, because different elements (or the light that they emit) have prominent characteristic colours. Cesium, indium, iodine, rubidium, and thallium were all named in this manner. Their respective Greek and Latin roots denote blue-gray, indigo, violet, red, and green (*thallus* means "tree sprout"). Because of the great variety of colours of its compounds, iridium takes its name from the Latin *iris*, meaning rainbow. Alternatively, an element name might suggest a mineral or the ore that contained it. One example is wolfram or tungsten (W), which was isolated from wolframite. Two other "inconsistent" elemental symbols, K and Na, arose from occurrence as well. *Kalium* was first obtained from the saltwort plant, *Salsola kali*, and *natrium* from niter. Their English names, potassium and sodium, are derived from the ores potash and soda.

Other elements, contrary to Lavoisier's suggestion, were named after planets, mythological figures, places, or superstitions. "Celestial elements" include helium (sun), tellurium (earth), selenium (moon–the element was discovered in close proximity to tellurium), cerium (the asteroid Ceres, which was discovered only two years before the element), and uranium (the planet Uranus, discovered a few years earlier). The first two transuranium elements (those *beyond* uranium) to be produced were named neptunium and plutonium for the next two planets, Neptune and Pluto. The names promethium (Prometheus, who stole fire from heaven), vanadium (Scandinavian goddess, Vanadis), titanium (Titans, the first sons of the earth), tantalum (Tantalos, father of Niobe), and thorium (Thor, Scandinavian god of war) all arise from Greek or Norse mythology.

"Geographical elements," sometimes honoured the discoverer's native country or workplace. The Latin names for Russia (*ruthenium*), France (*Gallium*), Paris (*lutetium*), and Germany (*germanium*) were among those used. Marie Sklodowska Curie named one of the elements that she discovered polonium, after her native Poland. Often the locale of discovery lends its name to the element; the record holder is certainly the Swedish village Ytterby, the site of ores from which the four elements terbium, erbium, ytterbium, and yttrium were isolated. Elements honouring important scientists include curium, einsteinium, nobelium, fermium, and lawrencium.

Most of the elements now known were given titles peacefully, but a few were not. Niobium, isolated in 1803 by Ekeberg from an ore that also contained tantalum, and named after the Greek goddess Niobe (daughter of Tantalus), was later found to be indentical to an 1802 discovery of Hatchett, columbium. (Interestingly, Hatchett first found the element in an ore sample that had been sent to England more than a century earlier by John Winthrop, the first governor of Connecticut.) While "niobium" became the accepted designation in Europe, the Americans, not surprisingly, chose "columbium." It was not until 1949–when the International Union of Pure and Applied Chemistry (IUPAC) ended more than a century of controversy by ruling in favor of mythology, that element 41 received a unique name.

In 1978, the IUPAC recommended that, at leas for now, **the elements beyond element 103 be known by systematic names based on numerical roots; element 104 is unnilquadium (un for 1, nil for 0, quad for 4, plus the -ium ending), followed by unnilpentium, unnilhexium, and so on**. Arguments over the proper names of elements 104 and 105 prompted the IUPAC to begin hearing claims of priority to numbers 104 through 109, but names for those elements have not yet been officially approved by IUPAC. In November 1993, the Nomenclature Committee of the American Chemical Society (ACS) endorsed the use in the United States of the names rutherfordium (Rf) and hahnium (Ha) for elements 104 and 105; and nielsbohrium (Ns), hassium (Hs) and meitnerium (Mt) for 107, 108, and 109, respectively. The name seaborgium (Sg), proposed by ACS for element 106, has been rejected by IUPAC. You will recognise some of these as derived from the names of scientists prominent in the development of atomic theory; others are named for scientists who were involved in the discovery of these heavy elements.

TRY YOURSELF - II

Q. 1 Convert each of the following into a correct formula represented with correct notation.
(a) $AlOH_3$; (b) $Mg(CO_3)$; (c) $Zn(CO_3)_2$; (d) $(NH_4)^2SO_4$; (e) $Mg_2(SO_4)_2$.

Q. 2 Two (2,000) moles of Ni atoms have the same mass as 1.223 moles of atoms of another element. What is the atomic weight of the other element? What is it? (Refer to any standard list of Atomic weights).

Q. 3 How many moles of substance are contained in each of the following samples?
(a) 16.8 g of NH_3; (b) 3.25 kg of ammonium bromide; (c) 5.6 g of PCl_5; (d) 126.5 g of Sn.

Q. 4 A 0.1647 gram sample of a pure hydrocarbon was burned in a C-H combustion train to produce 0.5694 gram of CO_2 and 0.0826 gram of H_2O. Determine the masses of C and H in the sample and the percentages of these elements in this hydrocarbon.

Q. 5 Nitric oxide, NO, is produced in internal combustion engines. When NO comes in contact with air, it is quickly converted into nitrogen dioxide, NO_2, a very poisonous, corrosive gas. What mass of O is combined with 3.00 g of N in (a) NO and (b) NO_2? Show that NO and NO_2 obey the Law of Multiple Proportions.

Q. 6 What mass of NaCl would contain the same *total* number of ions as 245 g of $MgCl_2$?

Q. 7 (a) What is the percent by mass of oxalic acid, $(COOH)_2$, in a sample of pure oxalic acid dihydrate, $(COOH)_2{\bullet}2H_2O$? (b) What is the percent by mass of $(COOH)_2$ in a sample that is 72.4% $(COOH)_2{\bullet}2H_2O$ by mass?

Q. 8 We can drive off the water from copper sulphate pentahydrate, $CuSO_4{\bullet}5H_2O$, by heating to obtain anhydrous (meaning "without water") $CuSO_4$. An experiment calls for 10.0 g of $CuSO_4$. How many grams of $CuSO_4{\bullet}5H_2O$ should we use to supply 10.0 g of $CuSO_4$

Q. 9 One method of analysing for the amount of Cr_2O_3 in a sample involves converting the chromium to $BaCrO_4$, and then weighing the amount of $BaCrO_4$ formed. Suppose that this process could be carried out with no loss of chromium. How many grams of Cr_2O_3 are present in the original sample for every gram of $BaCrO_4$ that could be isolated and weighed?

Q. 10 Near room temperature the density of water is 1.00 g/mL and the density of ethanol (grain alcohol) is 0.789 g/mL. What volume of ethanol contains the same number of molecules are present in 125 mL of H_2O?.

TRY YOURSELF - II

ANSWERS

1. $Al(OH)_3$; $MgCO_3$; $ZnCO_3$; $(NH_4)_2SO_4$; $MgSO_4$

2. ≈ 95.97 ; Molybdenum

3. (a) 0.988 mol NH_3

(b) 33.2 mol NH_4Br

(c) 0.027 mol PCl_5

(d) 1.066 mol Sn

4. 94.35% carbon; 5.61% hydrogen

5. (a) 3.43 grams of O

(b) 6.86 grams of O

6. 225 grams of NaCl

7. (a) 71.40%

(b) 51.7%

8. 15.6 grams of $CuSO_4.5H_2O$

9. 0.3 grams of Cr_2O_3

10. 406 mL ethanol