

Introduction to Chemistry

In Chemistry the word *weight* is commonly used in place of the more proper term *mass*.

1.1 Atoms, Ions, and Isotopes

Protons and Neutrons

Protons and **neutrons** are subatomic particles which are approximately equal in mass. Their mass is roughly equivalent to one atomic mass unit, *amu*. The proton has a unit positive charge of +1. In SI units this charge is equal to 1.6 x 10^{-19} C. The neutron, as its name suggests, is neutral.

Electrons

Electrons are subatomic particles have a mass that is almost 2000 times less than that of a proton or neutron. *On the MCAT the mass of an electron is typically ignored when compared to the masses of protons and neutrons.* The charge of an electron is equal and opposite to that of a proton. Its unit charge is -1, which is equal to -1.6 x 10^{-19} C.

Atoms and Ions

Atoms are the basic building blocks of matter. At the core of an atom is a nucleus composed of protons and neutrons. Orbiting the nucleus are electrons. Atoms are neutrally charged, and therefore contain the same number of electrons as protons.

An **ion** results when an atom, or a group of atoms, gains or loses electrons. Chemical species that lose electrons become positively charged; positively charged ions and are called **cations**. Species that gain electrons take on a negative charge and are referred to as **anions**.

Elements and Isotopes

The number of protons in the nucleus of an atom is equal to the **atomic number,** *Z*. An **element** is a substance composed of atoms all bearing the same atomic number. All atoms of an element share the same chemical properties. Each element is given a name, and is represented by a unique one or two letter symbol, for example, the symbol for the element carbon is *C*.

Elements are substances which cannot broken down into simpler substances by ordinary chemical means. All common forms of matter are composed of elements.

Isotopes are atoms of the same element which have different numbers of neutrons. Most naturally occurring elements have more than one isotope. The **mass number,** *A*, of any particular isotope is equal to the number of protons plus the number of neutrons contained in that atom. For example, the most common isotope of carbon contains 6 protons and 6 neutrons. The mass number of this isotope, therfore, is 12.

The convention for expressing the mass number, *A*, and the atomic number, *Z*, is: ${}^{A}_{Z}\mathrm{X}$, where X is the symbol for the element. The isotope, carbon twelve, is expressed as: ${}^{12}_{6}\mathrm{C}$.

1.2 Atoms

The Atomic Mass Unit

The ${}^{12}_{6}C$ isotope is used as the standard to define the atomic mass unit (amu). One atom of this isotope is arbitrarily assigned a mass of exactly 12 amu. Since ${}^{12}_{6}C$ contains 6 protons and 6 neutrons, the mass of a proton, or a neutron, is approximately equal to one amu. The mass of one amu is approximately 1.7×10^{-24} g.

Avogadro's Number

In order to make working in grams more convenient, the concept of the mole (mol) was developed. The mole is a number defined by the number of ${}^{12}C$ atoms contained in a sample that weighs exactly 12 g. One mole of amu's, by definition, is exactly equal to 1 gram.

The number of objects in a mole is referred to as **Avogadro's number,** and is approximately equal to 6×10^{23}. Although a mole typically refers to atoms, groups of atoms, or subatomic particles, its use is not restricted to to these examples. A mole, like a dozen, may be used to refer to anything. The mass of 6×10^{23} objects is referred to as **molar mass.**

exercise ►

What is the approximate molar mass (in grams) of a proton?

solution:

${}^{12}C$ contains 6 protons, 6 neutrons and 6 electrons. The mass of the electrons is negligable, and the masses of the nucleons (the particles in the nucleus) are approximately equal.

Since one mole of ${}^{12}C$ atoms contains 12 neucleons and weighs exactly 12 grams, the six moles of protons in one mole of ${}^{12}C$ must weigh six grams. Each mole of protons, therefore, weighs 1 gram.

OR

$$\frac{1.7 \times 10^{-24}\text{ g}}{\text{proton}} \times \frac{6 \times 10^{23}\text{ protons}}{\text{mol}} = \frac{1\text{ g}}{\text{mol of protons}}$$ ◆

■ Note that a mole of protons or neutrons, are roughly equal to 1 gram.

Atomic Weight (a.k.a. average atomic mass)

Since carbon contains more than one isotope, one mole of naturally occurring carbon will have a weight which is not exactly 12 g. The mass of a naturally occuring sample of one mole of an element is called the **atomic weight**. The atomic weight is the average of all isotope weights based on the abundance of each isotope. *The abundance is simply the decimal equivalent of percent. Just as the sum of all percentages equals 100, the sum of all abundances equals one.*

For example, carbon has three isotopes: ${}^{12}C$ with a mass of 12.00000 amu, which comprises 98.892% of all carbon atoms, ${}^{13}C$ with a mass of 13.00335 amu, and ${}^{14}C$ which is present in only trace quantities.

In order to find the atomic weight of carbon we used the following formula:

$$\boxed{AW = \sum a \cdot w}$$

where *AW* is the atomic weight, *a* is the abundence, and *w* is the isotope weight. The symbol Σ, indicates that you are to take the sum of all the abundances times isotope weights.

Since ^{14}C is only present in trace amounts we may ignore it. The abundance of ^{12}C is 0.98892. The abundance of ^{13}C is found by subtracting the abundance of ^{12}C form one. *Just as we would subtract a percentage from 100, we subtract an abundance from one.*. This yeilds:

$$AW = (0.98892)(12.00000) + (0.01108)(13.00335) = 12.011 \text{ amu/atom} = 12.011 \text{ g/mol.}$$

The atomic weights of all the elements are given in periodic tables which are part of the MCAT test booklets.

1.3 Molecules

Molecular Formula and Molecular Weight

A **molecule** (molec) is composed of two or more atoms which are bonded together. If a molecule contains more than one element than it is called a **compound**. One molecule of benzene contains six carbon atoms and six hydrogen atoms. The **molecular formula** of benzene is C_6H_6. The **subscripts** after each atomic symbol may be thought of as the number of atoms per molecule, or as the number of moles of atoms per mole of molecules.

The **molecular weight**, *MW* (a.k.a. molecular mass), of a molecule is found by the adding up all the atomic weights of each atom. Take benzene for example, 6 times the AW of C + 6 times the AW of H gives:

$$6(12 \text{ g/mol}) + 6(1 \text{ g/mol}) = 78 \text{ g/mol}$$

The MW of benzene is 78 g/mol, or 78 amu/molec.

Empirical Formula and Empirical Weight

The **empirical formula** of a compound is the simplest whole number ratio of the types of atoms in that compound. For benzene, C_6H_6, the empirical formula is CH. For glucose, $C_6H_{12}O_6$, the empirical formula is CH_2O. The **empirical weight**, *EW* (a.k.a. empirical mass), of a compound is found by adding the atomic weights of all atoms in the empirical formula. The empirical weights of benzene and glucose respectively are, 13 g/mol and 30 g/mol.

If the empirical formula and the approximate molecular weight of a compound is known, the molecular formula can be determined by use of the following equation:

$$\boxed{\text{Molecular formula} = \left(\frac{\sim MW}{EW}\right) \text{Empirical formula}}$$

example ►

A compound with an approximate MW of 35 g/mol has an empirical formula of HO, what is its molecular formula?

solution:

When the EW of 17 is divided into the ~MW, the result, rounded off to the nearest whole number is 2. Therefore, each element in the empirical formula must be multiplied by 2, to give the molecular formula H_2O_2.◆

Mass Percent (a.k.a. weight percent)

The mass percent of an element in a compound is defined as mass of that element divided by the mass of the compound. To get percent this quantity is multiplied by 100:

$$\boxed{\text{mass }\%_A = \frac{\text{mass}_A}{\text{mass}_{\text{Total}}} \times 100}$$

example ►

What is the percent by mass of carbon in $C_6H_{12}O_6$?

solution:

Let the quantity of $C_6H_{12}O_6$ be one mole. The mass of C is 6(12) = 72 g. The total mass is equal to the molecular weight. MW = 6(12) + 12(1) + 6(16) = 180 g.

$$\text{mass }\%_C = \frac{72\text{g}_C}{180\text{g}_{\text{Total}}} \times 100 = 40\%$$◆

example ►

A compound containing only hydrogen and carbon (a hydrocarbon) has a mass percent of hydrogen equal to 9.10%. What is its empirical formula?

solution:

Let the mass of the hydrocarbon equal 100 g. Since the hydrogen is 9.10 g the remaining carbon must have a mass of 90.9 g.

1. **Convert grams into moles.**
   90.9 g C x mol/12 g = 7.58 mol C
   9.10 g H x mol/1 g = 9.10 mol H

2. **Divide each number of moles, by the smallest number above.**
   7.58 mol C/7.58 = 1.00 mol C
   9.10 mol H/7.58 = 1.20 mol H
   This yields $CH_{1.2}$

3. **Express the empirical formula as the simplest whole number ratio.**
   Since 1.2 is one and one–fifth, multiplication by five is required to convert the moles of hydrogen into a whole number. Both the carbon and hydrogen are multiplied by five below:
   $C_{(1 \times 5)}H_{(1.2 \times 5)} = C_5H_6$◆

1.4 Chemical Reactions

A **chemical reaction** occurs when one substance is converted into another substance For example, when iron is exposed to oxygen in the presence of water at 25°C, rust is formed. This process is represented by the **chemical equation** below:

$$4\,Fe(s) + 3\,O_2(g) \xrightarrow[25°C]{H_2O} 2\,Fe_2O_3(s)$$

In chemical equations, substances to the left of the arrow are called **reactants**, while those to the right are the **products**. The reaction conditions are indicated above and below the arrow. Solids, liquids and gasses are indicated by the symbols: *(s)*, *(l)*, and *(g)*, respectively. When substances are dissolved in water the mixture is called an **aqueous solution** and is indicated by *(aq)*. These different states of matter are referred to as **phases.**

The numbers which precede each substance are called **coefficients**. The *4* preceding Fe, the *3* preceding O_2 and the *2* preceding Fe_2O_3, indicate that if 4 moles of Fe react completely with O_2, 3 moles of O_2 will be consumed to form 2 moles of Fe_2O_3.

■ Note that chemical equations do not indicate the number of moles actually present.

In a balanced equation the number of moles of each element are the same on each side of the equation. The number of moles of each element may be obtained by multiplication of the coefficients with the subscripts associated with each element. Both the reactants and the products in the above equation contain four moles of iron atoms, and six moles of oxygen atoms.

exercise ►

How many moles of the element oxygen are present in four moles of barium nitrate, $Ba(NO_3)_2$?

solution:

$$4\,Ba(NO_3)_2,$$

Multiplying the coefficient 4 by the subscripts 3 and 2 gives 24 moles of oxygen.◆

exercise ►

Balance this equation: $SiCl_4 + Si \longrightarrow Si_2Cl_6$.

solution:

Since the reactant *Si* is not bonded to another element it is best to save it for last. Since the subscripts 4 and 6 on the chlorine are both common denominators of 12, the coefficient 3 and 2 can be added to the $SiCl_4$ and Si_2Cl_6 respectively to give:

$$3\,SiCl_4 + Si \longrightarrow 2\,Si_2Cl_6.$$

Since there are 4 moles of Si on both sides of this equation it is not necessary to add a coefficient in front of the Si, since the equation is already balanced.◆

exercise ►

Balance this equation (use whole numbers for all coefficients):

$$C_4H_{10} + O_2 \longrightarrow CO_2 + H_2O.$$

solution:

Since the reactant O_2 is not bonded to another element it is best to save it for last. The carbon can be balanced by adding a 4 in front of CO_2. The hydrogen can be balanced by adding a 5 in front of H_2O. This gives:

$$C_4H_{10} + O_2 \longrightarrow 4\ CO_2 + 5\ H_2O.$$

There are now a total of 13 moles of oxygen on the right side of the equation. Placing 7.5 or 13/2 in front of the O_2 balances the equation to yield:

$$C_4H_{10} + 13/2\ O_2 \longrightarrow 4\ CO_2 + 5\ H_2O.$$

Finally multiplying the entire equation through by 2 , gives

$$2\ C_4H_{10} + 13\ O_2 \longrightarrow 8\ CO_2 + 10\ H_2O.\blacklozenge$$

Limiting Reagents

$$4\ Fe(s) + 3\ O_2(g) \xrightarrow[25°C]{H_2O} 2\ Fe_2O_3(s)$$

If 4 moles of Fe and 4 moles of O_2 where initially present in the above reaction, it can be concluded that 1 mole of oxygen is in excess, since only 3 moles of oxygen are needed.

Since there is not enough Fe to react completely with the O_2, the Fe is called the **limiting reagent.** The limiting reagent is that reactant which, because it is present in insufficient amounts, limits the quantity of products formed. *Note that the substance with more moles, need* <u>*not*</u> *be the limiting reagent.*

example ►

For the above reaction, how many moles of O_2 are required to react completely with 5 moles of Fe?

solution:

$$5 \text{ mol Fe x } \frac{3 \text{ mol } O_2}{4 \text{ mol Fe}} = 3.75 \text{ mol } O_2 \blacklozenge$$

example ►

For the above equation, 279 g of Fe, and 128 g of O_2, are allowed to react. Which is the limiting reagent? By how many grams is the other reagent in excess? (The AW of Fe is 55.8 g/mol and that of O is 16 g/mol.)

solution:

It is first necessary to convert grams into moles. (Note that the MW of O_2 is 32 g/mol.)

$$279\text{ g Fe x }\frac{\text{mol Fe}}{55.8\text{ g}} = 5.00\text{ mol Fe}$$

$$128\text{ g }O_2\text{ x }\frac{\text{mol }O_2}{32.0\text{ g}} = 4.00\text{ mol }O_2$$

Note from the previous question that complete reaction of 5 moles of iron require 3.75 moles of O_2. Since 4 moles of O_2 are present, the O_2 is in excess, and the Fe is the limiting reagent.

We have 4 moles of O_2, but only 3.75 moles are required, therefore, we have 0.25 moles of O_2 in excess. Converting to grams gives:

$$0.25\text{ mol }O_2\text{ x }\frac{32\text{ g }O_2}{\text{mol}} = 8.0\text{ g }O_2\text{ in excess}$$ ◆

Theoretical Yield

When a reaction takes place it is common for the quantity of products produced to be less than that expected based on the chemical equation. The **theoretical yield** of a reaction is the quantity of products produced assuming that the reaction proceeds forward 100%.

example ►

Using the quantities given in the preceding example, what is the theoretical yield? (The MW of Fe_2O_3 is 160 g/mol.)

solution:

Since Fe was found to be the limiting reagent, it must be used to determine the theoretical yield. Note that 5 moles of Fe are present.

$$5\text{ mol Fe x }\frac{2\text{ mol }Fe_2O_3}{4\text{ mol Fe}}\text{ x }\frac{160\text{ g }Fe_2O_3}{\text{mol }Fe_2O_3} = 400\text{ g }Fe_2O_3$$ ◆

Percentage Yield

The **percentage yield** is a measure of how far a reaction proceeds in the forward direction. It is defined as the **actual yield** (the actual amount of products formed) divided by the theoretical yield, times 100:

$$\boxed{\frac{\textit{Actual}\text{ Y.}}{\text{Theor. Y.}}\text{ x }100 = \text{Percentage Yield}}$$

example ►

When 279 g of Fe and 128 g of O_2 are allowed to react, it is found that only 300 g of Fe_2O_3 are formed. Referring back to the last two examples, what is the percentage yield of this reaction?

solution:

The actual yield is given as 300 g. The theoretical was found to be 400 g.

$$\frac{300\ g}{400\ g} \times 100 = 75\% ◆$$

Types of Reactions

To be continued.