# Stoichiometry
# Unit 1

## Learning Outcomes

*At the end of this unit students will be expected to:*

### The Mole and Molar Mass

- define molar mass and perform mole-mass inter-conversions for pure substances
- explain how a major scientific milestone, the mole, changed chemistry

### Calculations and Chemical Equations

- identify mole ratios of reactants and products from balanced chemical equations
- identify practical problems that involve technology where equations were used
- state a prediction and a hypothesis based on available evidence and background information
- perform stoichiometric calculations related to chemical equations

### Stoichiometric Experimentation

- design stoichiometric experiments identifying and controlling major variables
- use instruments effectively and accurately for collecting data
- identify and explain sources of error and uncertainty in measurement using precision and accuracy
- communicate questions, ideas, and intentions, and receive, interpret, understand, support, and respond to the ideas of others
- identify various constraints that result in trade-offs during the development and improvement of technologies

### Applications of Stoichiometry

- identify various stoichiometric applications
- predict how the yield of a particular chemical process can be maximized
- explain how data support or refute the hypotheses or prediction of chemical reactions
- compare processes used in science with those used in technology
- analyse society's influence on science and technology

# An Introduction to Stoichiometry
## Lesson 1

How does a chemist count out $3.25 \times 10^{23}$ atoms of the element sodium, Na? In industry, when making extremely large amounts of chemicals through chemical reactions, how do chemists know how much of the reactants to react or how much product will be formed? This unit will answer these questions and other questions related to amount of matter.

The word *stoichiometry* comes from the Greek words, *stoicheion* (meaning any first thing or principle) and *metron* (meaning measure). Stoichiometry deals with the *mass-mass* or *mole-mole* relationship among reactants and products in a balanced chemical equation. It answers questions like how much of one reactant will react with a given amount of another reactant and how much product will be formed.

Whether a compound is being prepared by the chemical industry or in a high school laboratory, it deals with extremely large numbers of atoms, molecules, or ions. For example, if we were to prepare the flavouring agent, ethyl butyrate, we might use 10.0 g of butyric acid. This relatively small mass of butyric acid contains $6.8 \times 10^{22}$ molecules, a huge number. If we had an imaginary device that could count molecules at the rate of one million per second, it would take over two billion years for the device to count that many molecules.

To overcome the problem of dealing with extremely large numbers of particles, chemists use the mole concept. A *mole* is an amount (quantity) of matter that contains $6.02 \times 10^{23}$ particles of that matter. Depending on the make-up of matter, the particles could be atoms, molecules, ions, or formula units. The symbol for the mole unit is *mol*. The mole is one of the seven base units of the standard SI system of measurement. German chemist, Wilhelm Ostwald, introduced the mole unit around 1900. The mole is used by chemists to communicate the amount of matter they are using. The number $6.02 \times 10^{23}$ is called *Avogadro's number*, after the Italian scientist, Amedeo Avogadro, (1776-1856).

Another definition of the mole is: a mole of a substance is the quantity of that substance that contains the same number of chemical units (atoms, molecules, formula units, or ions) as there are atoms in exactly 12 g of the isotope carbon-12. The mole is thus a unit of measurement. A certain number of moles of a substance contain a certain number of grams and a certain number of particles.

National Mole Day is celebrated yearly across North American high schools on October 23$^{rd}$, starting at 6:02 a.m. The mole was a major milestone in chemistry; chemists could quantify their observations and use it to make quantitative predictions. The mole concept makes it extremely simple to convert from numbers of atoms or molecules or ions (in moles) to mass in grams, and vice versa. Avogadro's number provides a way of converting between the number of particles on the micro scale (which we cannot see) and the number of moles on the macro scale (which we can see).

## TEXT READING

*4.3 Balancing Chemical Equations* and *Translating Balanced Chemical Equations* on page 138.

### LESSON EXERCISES

1. Explain what is meant by stoichiometry in your own words.
2. Explain what is meant by the mole unit of measurement in your own words.
3. Why is it important to be able to use the mole unit of measurement?
4. What is the value of Avogadro's number?
5. How many molecules are there in two moles of water?

### JOURNAL ENTRY

# Calculating Average Atomic Masses
## Lesson 2

The *mass* of an atom is the sum of all of the electrons, protons, and neutrons it contains. The standard for atomic mass is the isotope carbon-12. It has a mass of exactly 12.0000 u. *Isotopes* are atoms of the same element that have different masses because of different numbers of neutrons. One atomic mass unit, 1 u, is 1/12 the mass of a carbon-12 atom. One atomic mass unit, 1 u, is equivalent to $1.66 \times 10^{-24}$ g. It is easier to use atomic mass units because the mass of an atom in grams, g, is extremely small. The mass of all element isotopes is compared to the carbon-12 atom.

### Sample Problem 1

What is the mass of an atom with a mass 7/6 times greater than the carbon-12 atom?

### Solution

The mass of an atom 7/6 times bigger than a carbon-12 atom is 7/6 × 12.0000 u = 14.0000 u.

### Sample Problem 2

What is the mass of an atom 4/13 times the mass of a chromium atom, if a chromium atom is 13/3 times the mass of a carbon-12 atom.

### Solution

The mass of the atom is:
4/13 × 13/3 × 12.0000 u = 16.0000 u.

If you looked up the atomic mass of an element (the mass of a single atom), you would find that atomic masses are not whole number integers. The atomic masses used in this course are located in the *Chemistry Data Booklet*.

Where do these masses come from? Examine hydrogen; this element has three different isotopes, ${}^{1}_{1}H$, ${}^{2}_{1}H$, and ${}^{3}_{1}H$ with masses of approximately 1.0 u, 2.0 u, and 3.0 u, respectively. If we took an average of these three isotopes, we would get an average of 2.0 u for the hydrogen atom. Looking up the atomic mass of hydrogen, we see that it is 1.01 u. Why the big difference? The atomic mass of an element is the weighted average of all of the naturally occurring isotopes of an element. The weighted average takes into consideration just how much there is of each isotope in nature. In the case of hydrogen, 99.985 per cent of all of the hydrogen atoms found in nature are the lightest isotope of hydrogen, ${}^{1}_{1}H$, with an approximate mass of 1.0 u. So the weighted average is close to 1.0 u.

### Sample Problem 3

Natural neon consists of 90.92 per cent ${}^{20}_{10}Ne$, actual mass 19.992 44 u; 0.257 per cent ${}^{21}_{10}Ne$, actual mass 20.993 95 u; and 8.82 per cent ${}^{22}_{10}Ne$, actual mass 21.991 38 u. Calculate the average atomic mass of natural neon.

### Solution

Assume that we have a sample of 100 000 neon atoms. Calculate the average by finding the sum of the masses of all of the 100 000 atoms and then dividing by 100 000. If we wanted to find the average weight of a student in a classroom, we would add up the weights of all of the students in the room and then divide by the number of students in the classroom. The average atomic mass of natural neon is found in a similar manner:

First calculate the amount of each isotope in 100 000 atoms.

$100\,000 \times 90.92\% = 90\,920\ ^{20}_{10}Ne$

$100\,000 \times 0.257\% = 257\ ^{21}_{10}Ne$

$100\,000 \times 8.82\% = 8820\ ^{22}_{10}Ne$

Then calculate the total atomic mass for each isotope, add the masses together, and divide by 100 000.

$$\frac{(90920 \times 19.992\,44\text{ u} + 257 \times 20.993\,95\text{ u} + 8820 \times 21.991\,38\text{ u})}{100\,000}$$

$= 20.2$ u.

## SAMPLE PROBLEM 4

Natural nitrogen consists of 99.63 per cent $^{14}_{7}N$, actual mass 14.003 07 u; and 0.37 per cent $^{15}_{7}N$, actual mass 15.000 11 u. Calculate the average atomic mass of natural nitrogen for a sample size of 10 000.

## SOLUTION

First calculate the amount of each isotope in 10 000 atoms.

$10\,000 \times 99.63\% = 9963\ ^{14}_{7}N$

$10\,000 \times 0.37\% = 37\ ^{15}_{7}N$

Then calculate the total atomic mass for each isotope, add the masses together, and divide by 10 000.The average atomic mass of natural nitrogen is:

$$\frac{9963 \times 14.003\,07\text{ u} + 37 \times 15.000\,11\text{ u}}{10\,000} = 14.01\text{ u}$$

## LESSON EXERCISES

1. Natural magnesium consists of 78.70 per cent $^{24}_{12}Mg$, actual mass 23.985 04 u; 10.13 per cent $^{25}_{12}Mg$, actual mass 24.985 84 u; and 11.17 per cent $^{26}_{12}Mg$, actual mass 25.982 59 u. Calculate the average atomic mass of natural magnesium.
2. Natural silicon consists of 92.21 per cent $^{28}_{14}Si$, actual mass 27.976 93 u; 4.70 per cent $^{29}_{14}Si$, actual mass 28.976 49 u; and 3.09 per cent $^{30}_{14}Si$, actual mass 29.973 76 u. Calculate the average atomic mass of natural silicon.
3. Natural argon consists of 0.337 per cent $^{36}_{18}Ar$, actual mass 35.967 55 u; 0.063 per cent $^{38}_{18}Ar$, actual mass 37.962 72 u; and 99.60 per cent $^{40}_{18}Ar$, actual mass 39.962 38 u. Calculate the average atomic mass of natural argon.

## JOURNAL ENTRY

# Determining Atomic Mass, Molecular Mass, Formula Mass, and Molar Mass
## Lesson 3

*Atomic mass* is the mass of an atom of an element in atomic mass units, u. This value is given in the periodic table, or international table of atomic masses, or in the *Chemistry Data Booklet*. For example, the atomic mass of an atom of the element sodium, Na, is 22.99 u. The atomic mass of an atom of the element xenon, Xe, is 131.29 u.

*Molecular mass* is the mass of a molecule of a compound or an element, in atomic mass units, u. It is the sum of all of the atomic masses of all of the atoms in the molecule. For example, the molecular mass of a molecule of water, $H_2O$, is the sum of:

2H = 2 × 1.01 u = 2.02 u and
1O = 1 × 15.99 u = 15.99 u.

The molecular mass of water is 18.01 u.

The molecular mass of $F_2$ is the sum of the atomic masses of F and F:

18.99 u + 18.99 u = 37.98 u.

Not all compounds are made up of molecules; which result from covalent bonding. Some compounds form by ionic bonding and are made up of ions. There are no molecules present so the compounds cannot be defined using molecular mass. These compounds consist of formula units instead of molecules. A *formula unit* is the smallest unit of an ionic compound that represents the properties of the ionic compound. It is the simplest whole number ratio of positive and negative ions in the ionic compound. A formula unit will have the same formula as the formula for the ionic compound.

*Formula mass* is the mass of a formula unit. A formula unit of table salt, NaCl (sodium chloride) is NaCl. The formula mass of the ionic compound NaCl is the sum of the atomic masses of the ions in the formula unit. This is:

$1Na^+$ = 22.99 u + $1Cl^-$ = 35.45 u = 58.44 u.

*Molar mass* is the mass of one mole of atoms, molecules, ions, or formula units in grams. It is the atomic mass, molecular mass, or formula mass expressed in *grams* instead of *atomic mass units*. The molar mass is numerically equal to the atomic mass, molecular mass, or formula mass, but in grams.

Looking at the chemicals we have referred to in this lesson, we know the molar mass of Na is 22.99 g, and the molar mass of Xe is 131.29 g, the molar mass of $H_2O$ is 18.01 g, the molar mass of $F_2$ is 37.98 g and the molar mass of NaCl is 58.44 g.

To determine the molar mass of a substance, look up the atomic mass, or molecular mass, or formula mass, and drop the u unit and replace it with the g unit.

Does this mean that the atomic mass of Na, 22.99 u is the same as the molar mass of Na, 22.99 g? No, only the numbers are the same, the amounts are completely different because of different units. For example, the atomic mass of Na is the mass of a single Na atom. One atomic mass unit = $1.66 \times 10^{-24}$ g. The amount in grams is:

$22.99 \times 1.66 \times 10^{-24}$ g = $3.82 \times 10^{-23}$ g.

The *molar mass* of Na is the mass of a mole of Na atoms, $6.02 \times 10^{23}$ Na atoms. This would be 22.99 g. As you can see, $3.82 \times 10^{-23}$ g is much smaller than 22.99 g. The table below summarizes the relationship between molar mass and atomic mass, molecular mass, and formula mass.

| substance | particles | atomic mass | molecular mass | formula mass | molar mass |
|---|---|---|---|---|---|
| **Na** | atoms | 22.99 u | | | 22.99 g |
| **Xe** | atoms | 131.29 u | | | 131.29 g |
| **$H_2O$** | molecules | | 18.01 u | | 18.01 g |
| **$F_2$** | molecules | | 37.98 u | | 37.98 g |
| **NaCl** | formula units | | | 58.44 u | 58.44 g |

## SAMPLE PROBLEM 1

What are the atomic masses of Mn and K?

## SOLUTION

The atomic mass of Mn is 54.94 u and the atomic mass of K is 39.10 u.

## SAMPLE PROBLEM 2

What are the molecular masses of $NH_3$ and $CH_3OH$?

## SOLUTION

The molecular mass of $NH_3$ is:

1N = 1 × 14.01 u = 14.01 u

3H = 3 × 1.01 u = 3.03 u

14.01 u + 3.03 u = 17.04 u

The molecular mass of $CH_3OH$ is:

1C = 1 × 12.01 u = 12.01 u

4H = 4 × 1.01 U = 4.04 u

1O = 1 × 15.99 u = 15.99 u

12.01 u + 4.04 u + 15.99 u = 32.04 u

## SAMPLE PROBLEM 3

What are the formula masses of $K_2SO_4$ and $(NH_4)_3PO_4$?

## SOLUTION

The formula mass of $K_2SO_4$ is:

2K = 2 × 39.10 u = 78.20 u

1S = 1 × 32.07 u = 32.07 u

4O = 4 × 15.99 u = 63.96 u

Formula mass = 174.23 u

The formula mass of $(NH_4)_3PO_4$ is:

3N = 3 × 14.01 u = 42.03 u

12H = 12 × 1.01 u = 12.12 u

1P = 1 × 30.97 u = 30.97 u

4O = 4 × 15.99 u = 63.96 u

Formula mass = 149.08 u

## SAMPLE PROBLEM 4

What is the molar mass of Al, $C_2H_6$, and $Mg(OH)_2$?

## Solution

The atomic mass of Al is 26.98 u. The molar mass of Al is therefore 26.98 g.

The molecular mass of $C_2H_6$ is 30.08 u. The molar mass of $C_2H_6$ is therefore 30.08 g.

The formula mass of $Mg(OH)_2$ is 58.31 u. The molar mass of $Mg(OH)_2$ is therefore 58.31 g.

## Text Reading

*Molar Mass* on page 143.

### Lesson Exercises

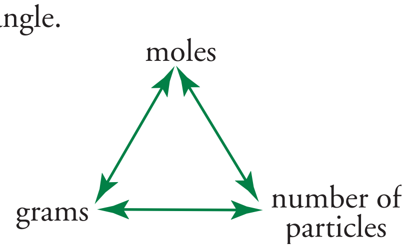
1. What is the atomic mass of Li, Ag, and Kr?
2. What is the molecular mass of $CCl_2F_2$, $C_2H_5OH$, and $NBr_3$?
3. What is the formula mass of $Na_2CO_3$, $(NH_4)_2SO_4$, and $Ca_3(PO_4)_2$?
4. What is the molar mass of Ne, $SiCl_4$, $K_3PO_4$, Mg, $Sr(NO_3)_2$, and $CH_3Br$?

### Journal Entry

# MOLE CONVERSIONS
## LESSON 4

A mole is an amount or quantity of matter that contains $6.02 \times 10^{23}$ particles of that matter. Depending upon the make-up of matter, the particles could be atoms, molecules, ions, or formula units.

There are different ways of expressing the amount of matter you have, such as grams or moles. By knowing the number of moles or grams you have, you can calculate how many particles you have. Otherwise, even if you could count individual particles, you could not count the extremely large number of particles in a sample of matter. For a given sample of matter, moles, grams, and number of particles are all interrelated. If you know one of the three (moles, grams, or number of particles) you can calculate how much you have of the other two. These conversions are known as *mole conversions*. The three quantities are found in the mole triangle.




Use the *factor label method* or *dimensional analysis* when doing mole conversions. In this method, you write down and keep track of the units or dimensions. You multiply a given quantity by a ratio (conversion factor), which converts the given quantity into the quantity you require. There is a relationship between the two quantities in the ratio. You treat the same units as factors that you would cancel in the same way that numbers are cancelled in algebra. For example, $\cancel{2} \times \frac{3}{\cancel{2}} = 3$. The 2s are identical numbers that we can cancel out. Likewise, we can cancel similar units by treating them as similar numbers. For example, $\cancel{g} \times mol/\cancel{g} = mol$.

### SAMPLE PROBLEM 1

How many mol and atoms are there in a 22.3 g sample of copper, Cu?

### SOLUTION

Multiply 22.3 g of Cu by a conversion factor (ratio) that will change it into mol.

$$\frac{22.3\ g \times mol\ Cu}{g\ Cu}$$

What is the relationship between # mol of Cu and # g of Cu? Choose a relationship that is simple to figure out. Assume there is 1 mol of Cu. Now determine how many g of Cu there are in 1 mol of Cu. The molar mass of Cu is 63.55 g. Therefore, the ratio is:

$$\frac{1\ mol\ Cu}{63.55\ g\ Cu}$$

We can now complete the calculation

$$\frac{22.3\ \cancel{g\ Cu} \times 1\ mol\ Cu}{63.55\ \cancel{g\ Cu}} = 0.351\ mol\ Cu.$$

The number of atoms in 22.3 g of Cu is 22.3 g Cu × atoms Cu/g Cu.

Now determine the relationship between # atoms of Cu and # g of Cu. Use Avogadro's number, $6.02 \times 10^{23}$, for the number of Cu atoms. This is also the number of Cu atoms in 1 mol of Cu. Put the molar mass of Cu on the bottom of the ratio to give:

$$\frac{6.02 \times 10^{23} \text{ atoms Cu}}{63.55 \text{ g Cu}}$$

Now complete the calculation:

$$\frac{22.3 \text{ g } \cancel{\text{Cu}} \times 6.02 \times 10^{23} \text{ atoms Cu}}{63.55 \text{ g } \cancel{\text{Cu}}}$$

$= 2.11 \times 10^{23}$ atoms Cu.

## Sample Problem 2

How many grams and atoms are there in 0.750 mol of potassium, K?

## Solution

$$\frac{0.750 \cancel{\text{ mol K}} \times 39.10 \text{ g K}}{1 \cancel{\text{ mol K}}} = 29.3 \text{ g K}$$

$$\frac{0.750 \cancel{\text{ mol K}} \times 6.02 \times 10^{23} \text{ atoms K}}{1 \cancel{\text{ mol K}}}$$

$= 4.52 \times 10^{23}$ atoms K

## Sample Problem 3

How many grams and mol are there in $2.56 \times 10^{23}$ atoms Mg?

## Solution

$$\frac{2.56 \times 10^{23} \cancel{\text{ atoms Mg}} \times 24.31 \text{ g Mg}}{6.02 \times 10^{23} \cancel{\text{ atoms Mg}}}$$

$= 10.3$ g Mg.

$$\frac{2.56 \times 10^{23} \cancel{\text{ atoms Mg}} \times 1 \text{ mol Mg}}{6.02 \times 10^{23} \cancel{\text{ atoms Mg}}}$$

$= 0.425$ mol Mg

So far, we have restricted our calculations to those involving elements. Now look at mole conversions involving compounds with molecules and formula units. The method of doing these mole conversions is similar to the method used when working with elements and atoms.

## Sample Problem 4

How many mol and molecules are there in 25.0 g of water, $H_2O$?

## Solution

$$\frac{25.0 \text{ g } \cancel{H_2O} \times 1 \text{ mol } H_2O}{18.01 \text{ g } \cancel{H_2O}}$$

$= 1.39$ mol $H_2O$

$$\frac{25.0 \text{ g } \cancel{H_2O} \times 6.02 \times 10^{23} \text{ molecules } H_2O}{18.01 \text{ g } \cancel{H_2O}}$$

$= 8.36 \times 10^{23}$ molecules $H_2O$

## Sample Problem 5

How many g and formula units are there in 1.15 mol of $AgNO_3$?

## Solution

$$\frac{1.15 \cancel{\text{ mol } AgNO_3} \times 169.85 \text{ g } AgNO_3}{1 \cancel{\text{ mol } AgNO_3}}$$

$= 195$ g $AgNO_3$

$$\frac{1.15 \cancel{\text{ mol } AgNO_3} \times 6.02 \times 10^{23} \text{ formula units } AgNO_3}{1 \cancel{\text{ mol } AgNO_3}}$$

$= 6.92 \times 10^{23}$ formula units $AgNO_3$

## Sample Problem 6

How many mol and g are there in $4.50 \times 10^{22}$ molecules of $NH_3$?

### SOLUTION

$$\frac{4.50 \times 10^{22} \cancel{\text{ molecules NH}_3} \times 1 \text{ mol NH}_3}{6.02 \times 10^{23} \cancel{\text{ molecules NH}_3}}$$

$= 0.0748 \text{ mol } NH_3$

$$\frac{4.50 \times 10^{22} \cancel{\text{ molecules NH}_3} \times 17.04 \text{ g NH}_3}{6.02 \times 10^{23} \cancel{\text{ molecules NH}_3}}$$

$= 1.27 \text{ g } NH_3$

### SAMPLE PROBLEM 7

How many atoms are there in 42.5 g of $CF_4$?

### SOLUTION

$$42.5 \cancel{\text{ g CF}_4} \times \frac{6.02 \times 10^{23} \text{ molecules CF}_4}{87.97 \cancel{\text{ g CF}_4}}$$

$= 2.91 \times 10^{23} \cancel{\text{ molecules CF}_4} \times 5 \text{ atoms/}1\cancel{\text{molecule}}$

$= 1.45 \times 10^{24}$ atoms

### TEXT READING

*Mass-Amount Conversions* on pages 143-144.

### LESSON EXERCISES

1. How many g and atoms are present in 0.450 mol of barium, Ba?
2. How many mol and atoms are there in 35.06 g of iron, Fe?
3. How many mol and g are there in $5.25 \times 10^{24}$ atoms of rubidium, Rb?
4. How many g and molecules are present in 2.76 mol of propane, $C_3H_8$?
5. How many mol and formula units are present in 85.62 g of $CaSO_4$?
6. How many mol and g are present in $3.75 \times 10^{23}$ formula units of KCl?
7. How many atoms are present in 56.25 g of $C_2H_6$?
8. How many formula units are there in 75.00 g of LiF?
9. Which is the largest mass in g, 25.2 g of $CaBr_2$, 1.80 mol of $CH_4$, or $1.75 \times 10^{24}$ molecules of $NCl_3$?
10. Why isn't the atomic mass of potassium, K, the same as the molar mass of potassium?

### JOURNAL ENTRY

# Calculating the Percentage Composition of a Compound
## Lesson 5

Quite often if a compound is being identified, its *percentage composition* is calculated. This involves calculating the *percentage by mass* of each element in the compound. To calculate the percentage composition for an element in a compound we need to know the total mass of the element in the compound and the total mass of the compound. The percentage of the element in the compound is the *total mass of the element* in the compound divided by the *total mass of the compound*. This ratio is then multiplied by 100.

### Sample Problem 1

What is the percentage composition of propane gas, $C_3H_8$, used in barbecues?

### Solution

The molecular mass of $C_3H_8$ is 44.11 u.

$$\%C = \frac{\text{total mass of C in compound}}{\text{mass of compound}} \times 100$$

There are three C atoms in $C_3H_8$ so the total mass of C is 3 × 12.01 u = 36.03 u.

$$\%C = \frac{36.03 \text{ u}}{44.11 \text{ u}} \times 100 = 81.68\%$$

$$\%H = \frac{\text{total mass of H in compound}}{\text{mass of compound}} \times 100$$

There are eight H atoms in $C_3H_8$ so the total mass of H is 8 × 1.01 u = 8.08 u.

$$\%H = \frac{8.08 \text{ u}}{44.11 \text{ u}} \times 100 = 18.3\%$$

### Sample Problem 2

What is the percentage composition of $Ca_3(PO_4)_2$? The formula mass of $Ca_3(PO_4)_2$ is 310.10 u.

### Solution

$$\%Ca = \frac{\text{total mass of Ca in compound}}{\text{mass of compound}} \times 100$$

There are three Ca atoms in $Ca_3(PO_4)_2$ so the total mass of Ca is 3 × 40.08 u = 120.24 u.

$$\%Ca = \frac{120.24 \text{ u}}{310.10 \text{ u}} \times 100 = 38.774\%$$

$$\%P = \frac{\text{total mass of P in compound}}{\text{mass of compound}} \times 100$$

There are two P atoms in $Ca_3(PO_4)_2$ so the total mass of P is 2 × 30.97 u = 61.94 u.

$$\%P = \frac{61.94 \text{ u}}{310.10 \text{ u}} \times 100 = 19.97\%$$

$$\%O = \frac{\text{total mass of O in compound}}{\text{mass of compound}} \times 100$$

There are eight O atoms in $Ca_3(PO_4)_2$ so the total mass of O is 8 × 15.99 u = 127.92 u.

$$\%O = \frac{127.92 \text{ u}}{310.10 \text{ u}} \times 100 = 41.251\%$$

### Sample Problem 3

Calculate the percentage composition of $CH_3COOH$. The molecular mass of $CH_3COOH$ is 60.04 u.

### Solution

$$\%C = \frac{\text{total mass of C in compound}}{\text{mass of compound}} \times 100$$

There are two C atoms in $CH_3COOH$ so the total mass of C is 2 × 12.01 u = 24.04 u.

$$\%C = \frac{24.02\ u}{60.04\ u} \times 100 = 40.01\%$$

$$\%H = \frac{\text{total mass of H in compound}}{\text{mass of compound}} \times 100$$

There are four H atoms in $CH_3COOH$ so the total mass of H is 4 × 1.01 u = 4.04 u.

$$\%H = \frac{4.04\ u}{60.04\ u} \times 100 = 6.73\%$$

$$\%O = \frac{\text{total mass of O in compound}}{\text{mass of compound}} \times 100$$

There are two O atoms in $CH_3COOH$ so the total mass of O is 2 × 15.99 u = 31.98 u.

$$\%O = \frac{31.98\ u}{60.04\ u} \times 100 = 53.26\%$$

## SAMPLE PROBLEM 4

Calculate the percentage of O in $CaCO_3$. In this question we are only calculating the percentage by mass of a single element in the compound, oxygen.

## SOLUTION

The formula mass of $CaCO_3$ is 100.06 u.

$$\%O = \frac{\text{total mass of O in compound}}{\text{mass of compound}} \times 100$$

There are three O atoms in $CaCO_3$ so the total mass of O is 3 × 15.99 u = 47.97 u.

$$\%O = \frac{47.97\ u}{100.06\ u} \times 100 = 47.94\%$$

In percentage composition problems, the sum of all of the percentages of all of the elements should be 100 per cent or very close to 100 per cent. If it is not, then you know there is a mistake in the solution to the problem. Also, percentage composition problems ask you to calculate the percentage by mass of each element in the compound. So, in sample problem 2, we did not calculate the percentage of $PO_4$ because $PO_4$ is not an element. We needed to calculate the percentage of P and O.

## LESSON EXERCISES

1. Calculate the percentage composition of one of the components of gasoline, octane, $C_8H_{18}$.
2. Calculate the percentage composition of the antacid, milk of magnesia, $Mg(OH)_2$.
3. Calculate the percentage composition of baking soda, sodium hydrogen carbonate, $NaHCO_3$.
4. Calculate the percentage composition of the acid found in car batteries, sulphuric acid, $H_2SO_4$.
5. Calculate the percentage composition of the explosive, TNT, trinitrotoluene, $C_7H_5(NO_2)_3$.
6. What is the percentage of N in smelling salts, $(NH_4)_2CO_3$?

## JOURNAL ENTRY

# Empirical Formula
## Lesson 6

*Empirical formula* is the simplest formula for writing a compound. It is the simplest whole number ratio of atoms or ions in the formula for the compound. It does not tell us the actual number of atoms of an element in a molecule of the compound. Determining the empirical formula for a compound can be used to identify the compound. To calculate the empirical formula of a compound we must know either the masses of all of the elements in the compound or the percentage composition by mass of the compound.

### Sample Problem 1

What is the empirical formula for a compound which on analysis shows 2.2% H, 26.7% C, and 71.1% O?

### Solution

We are trying to find the subscripts X, Y, and Z in the empirical formula for the compound, $H_xC_yO_z$. Use the following steps:

- Assume you have 100 g of the compound. Since per cent means out of 100, we will know how many g we have of each element.
- Change % to g
- Change g to mol
- Divide the smallest number of moles into itself and all of the other number of moles.
- If the results of the above step are whole numbers for each element or very close to whole numbers for each element, then these numbers represent the subscripts of the elements in the empirical formula and we have calculated the empirical formula.
- If a mole ratio is not close to a whole number, then multiply the ratio by the smallest whole number to make the ratio a whole number. Then multiply all of the mole ratios by this whole number to give the subscripts for the empirical formula.

These steps are summarized in the table below.

| **H** | **C** | **O** |
|---|---|---|
| 2.2 g | 26.7 g | 71.1 g |
| $\frac{2.2\text{ g H} \times 1\text{ mol H}}{1.01\text{ g H}}$ = 2.2 mol H | $\frac{26.7\text{ g C} \times 1\text{ mol C}}{12.01\text{ g C}}$ = 2.22 mol C | $\frac{71.1\text{ g O} \times 1\text{ mol O}}{15.99\text{ g O}}$ = 4.45 mol O |
| $\frac{2.2\text{ mol}}{2.2\text{ mol}}$ = 1.0 | $\frac{2.22\text{ mol}}{2.2\text{ mol}}$ = 1.0 | $\frac{4.45\text{ mol}}{2.2\text{ mol}}$ = 2.0 |
| **1** | **1** | **2** |

The empirical formula is $H_1C_1O_2$ or $HCO_2$.

### Sample Problem 2

What is the empirical formula of a compound which on analysis shows 69.9% Fe and 30.1% O?

| **Fe** | **O** |
|---|---|
| 69.9 g | 30.1 g |
| $\frac{69.9\text{ g Fe} \times 1\text{ mol Fe}}{55.8\text{ g Fe}}$ = 1.25 mol Fe | $\frac{30.1\text{ g O} \times 1\text{ mol O}}{15.99\text{ g O}}$ = 1.88 mol O |
| $\frac{1.25\text{ mol}}{1.25\text{ mol}}$ = 1.00 | $\frac{1.88\text{ mol}}{1.25\text{ mol}}$ = 1.50 |
| **2 × 1.00 = 2.00** | **2 × 1.50 = 3.00** |

Since we did not have all whole numbers or numbers very close to whole numbers, multiply 1.50 by the smallest whole number that will make 1.50 a whole number. In this case, multiply by two. Then multiply all other ratios by two. Therefore, 1.00 becomes 2.00. The empirical formula of the compound is $Fe_2O_3$.

## SAMPLE PROBLEM 3

When 2.435 g of antimony, Sb, is heated with an excessive amount of sulphur, S, a chemical reaction takes place. The antimony joins with some of the sulphur and the excess sulphur escapes, leaving only the compound. If 3.397 g of the compound was made, what is the empirical formula of the compound?

## SOLUTION

This problem is worded differently than the two previous problems. We are not given percentages of each element in the compound. However, if we know the mass of each element in the compound, we can determine the empirical formula by changing g to mol and following the rest of the steps. We know the mass of the antimony, Sb. It is 2.435 g. The compound contains only Sb and S. We know the mass of the compound is 3.397 g.

| Sb | S |
|---|---|
| 2.435 g | 0.962 g |
| $\frac{2.435 \text{ g Sb} \times 1 \text{ mol Sb}}{121.76 \text{ g Sb}}$ = 0.02000 mol Sb | $\frac{0.962 \text{ g S} \times 1 \text{ mol S}}{32.07 \text{ g S}}$ = 0.03000 mol S |
| $\frac{0.02000 \text{ mol}}{0.02000 \text{ mol}}$ = 1.000 | $\frac{0.03000 \text{ mol}}{0.02000 \text{ mol}}$ = 1.500 |
| **2 × 1.000 = 2.000** | **2 × 1.500 = 3.000** |

The mass of S present in the compound is 3.397 g - 2.435 g = 0.962 g S. The empirical formula is $Sb_2S_3$.

## TEXT READING

*Molecular Compounds* and *Empirical Molecular Formulas* on pages 269-271.

## LESSON EXERCISES

1. Determine the empirical formula of a compound which on analysis shows: 38.67% C; 16.23% H; and 45.10% N.
2. Determine the empirical formula of a compound which on analysis shows: 71.65% Cl; 24.27% C; and 4.07% H.
3. A 61.94 g sample of phosphorous was reacted with enough oxygen to form a white compound made up of only P and O and having a mass after reaction of 141.94 g. Determine the empirical formula of this compound.
4. Determine the empirical formula of a compound with the following percentages: Fe = 63.53% and S = 36.47%.
5. Determine the empirical formula of a compound which on analysis shows: 26.52% Cr; 24.52% S; and 48.96% O.
6. Determine the empirical formula of a compound with the following percentages: 63.1% C; 11.92% H; and 24.97% F.

## JOURNAL ENTRY

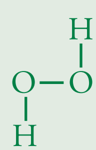
# Determining Molecular Formula
## Lesson 7

The empirical formula gives us the minimum information about a compound. It just shows the simplest whole-number ratio of atoms of the elements in a compound. A formula that provides more information about a compound is the molecular formula. The *molecular formula* shows the actual number of atoms of each element joined together to form a single molecule of the compound. More information is provided by a structural formula for a compound. A *structural formula* shows exactly which atoms are joined together in a molecule. The diagram below summarizes information provided by an empirical formula, a molecular formula, and a structural formula.

**Hydrogen Peroxide**

empirical formula $HO$

molecular formula $H_2O_2$

structural formula 

To determine the molecular formula for a compound we need to know its empirical formula and its molecular mass. The molecular mass is some whole number multiple of the mass of the empirical formula. We can determine the molecular formula by using the following equation:

(mass of empirical formula)$x$ = molecular mass

(the total mass of the empirical formula that was given or determined)(some whole number) = (total mass of the molecule)

We then multiply each subscript in the empirical formula by the whole number $x$ to give the molecular formula.

### Sample Problem 1

A certain compound has an empirical formula of $CH_2O$ and a molecular mass of 180 u. What is the molecular formula of the compound?

### Solution

(mass of empirical formula)$x$ = molecular mass

$(1 \times 12.01\text{ u} + 2 \times 1.01\text{ u} + 1 \times 15.99\text{ u})x = 180\text{ u}$

$(30.02\text{ u})x = 180\text{ u}$

$x = 6.00$

Since the empirical formula is $CH_2O$, multiply each subscript in the empirical formula by 6.00 to give the molecular formula $C_6H_{12}O_6$.

### Sample Problem 2

The empirical formula for a compound is $P_2O_5$ and the molecular mass of the compound is 283.88 u. What is the molecular formula of the compound?

### Solution

(mass of empirical formula)$x$ = molecular mass

$(2 \times 30.97\text{ u} + 5 \times 15.99\text{ u})x = 283.88\text{ u}$

$(141.89\text{ u})x = 283.88\text{ u}$

$x = 2.0007$

Since the empirical formula is $P_2O_5$, we multiply each subscript in the empirical formula by 2 to give the molecular formula $P_4O_{10}$.

## SAMPLE PROBLEM 3

A certain compound was found to have 21.9% Na; 45.7% C; 1.9% H; and 30.5% O. What is the molecular formula for this compound if its molecular mass is 210 u?

## SOLUTION

First determine the empirical formula.

| **Na** | **C** | **H** | **O** |
|---|---|---|---|
| 21.9 g | 45.7 g | 1.9 g | 30.5 g |
| 0.952 mol | 3.80 mol | 1.9 mol | 1.91 mol |
| $\frac{0.952 \text{ mol}}{0.952 \text{ mol}}$ = 1.00 | $\frac{3.80 \text{ mol}}{0.952 \text{ mol}}$ = 3.99 | $\frac{1.9 \text{ mol}}{0.952 \text{ mol}}$ = 2.0 | $\frac{1.91 \text{ mol}}{0.952 \text{ mol}}$ = 2.01 |
| **1** | **4** | **2** | **2** |

The empirical formula is $NaC_4H_2O_2$.

(mass of empirical formula)$x$ = molecular mass

$1 \times 22.99 + 4 \times 12.01 + 2 \times 1.01 + 2 \times 15.99 =$ 105.03 u

$(105.03 \text{ u})x = 210 \text{ u}$

$x = 2.00$

The molecular formula is $Na_2C_8H_4O_4$.

## LESSON EXERCISES

1. The empirical formula of a certain compound is CH. It has a molecular mass of 52 u. What is its molecular formula?
2. A certain compound was analysed and found to have 32.0% C; 6.7% H; and 18.7% N; with the rest of the compound made up of oxygen, O. The molecular mass is 75.0 u. What is the molecular formula of the compound?
3. What is the molecular formula for a compound whose molecular mass is 116 u and whose empirical formula is CHO?
4. What is the molecular formula for a compound whose molecular mass is 90 u and whose empirical formula is $CH_2O$?
5. A certain compound was found to contain 48.48% C; 5.05% H; 14.14% N; and 32.32% O. The molecular mass is 198 u. What is its molecular formula?
6. A certain compound was found to have the following percentage composition: 49.38% C; 3.55% H; 9.40% O; and 37.67% S. The molecular mass is 170.2 u. What is its molecular formula?

## JOURNAL ENTRY

# STOICHIOMETRY CALCULATIONS
## LESSON 8

A balanced chemical equation provides a lot of information. It tells what the reactants are in a chemical reaction and what the products are. A balanced chemical equation obeys the *law of conservation of mass*. The *total mass of all reactants* before a chemical reaction equals the *total mass of products* after the chemical reaction takes place. This concept leads us to the *mole-mole* or *mass-mass* relationship in a balanced chemical equation, known as stoichiometry. Examine this relationship through the following balanced chemical equation:

$$C_{(s)} + O_{2(g)} \rightarrow CO_{2(g)}$$

From this equation we can infer that one atom of C reacts with one molecule of $O_2$ to produce one molecule of $CO_2$. From a mass relationship, 12 u of C would react with 32 u of $O_2$ to produce 44 u of $CO_2$. The law of conservation of mass is preserved. We could also say that 10 atoms of C reacts with 10 molecules of $O_2$ to make 10 molecules of $CO_2$. Again, from a mass relationship, 120 u of C reacts with 320 u of $O_2$ to produce 440 u of $CO_2$. We could also say that $6.02 \times 10^{23}$ atoms of C reacts with $6.02 \times 10^{23}$ molecules of $O_2$ to produce $6.02 \times 10^{23}$ molecules of $CO_2$.

This next relationship is very important. Based on the last statement and since there are $6.02 \times 10^{23}$ particles in one mol of a substance, we can say that one mol of C reacts with one mol of $O_2$ to make one mol of $CO_2$. That is, 12 g of C reacts with 32 g of $O_2$ to produce 44 g of $CO_2$. This is the mole-mole and mass-mass relationship in a balanced chemical equation. It is the *stoichiometry* of a chemical reaction. This information is summarized in the table below:

| $C_{(s)}$ + | $O_{2(g)}$ → | $CO_{2(g)}$ |
|---|---|---|
| 1 atom | 1 molecule | 1 molecule |
| 12 u | 32 u | 44 u |
| 10 atoms | 10 molecules | 10 molecules |
| 120 u | 320 u | 440 u |
| $6.02 \times 10^{23}$ atoms | $6.02 \times 10^{23}$ molecules | $6.02 \times 10^{23}$ molecules |
| 1 mol | 1 mol | 1 mol |
| 12 g | 32 g | 44 g |

Examine some calculations involving balanced chemical equations using the *factor-label* method or *dimensional analysis* to solve the problems. This is the same method used for mole conversions.

### SAMPLE PROBLEM 1

Given the reaction:

$C_{(s)} + O_{2(g)} \rightarrow CO_{2(g)}$

a) How many g of $O_2$ reacts with 9.8 g of C?

b) How many g of $CO_2$ are produced when 19.6 g of C is completely reacted with enough $O_2$?

c) How many mol of $CO_2$ are produced from reacting 0.85 mol of $O_2$?

d) How many g of C are needed to react with 1.45 mol of $O_2$?

e) How many molecules of $O_2$ are needed to react with 5.26 g of C?

f) How many molecules of $CO_2$ are made by reacting $2.41 \times 10^{23}$ atoms of C?

## SOLUTION

a) $C_{(s)} + O_{2(g)} \rightarrow CO_{2(g)}$

| $C_{(s)}$ | $O_{2(g)}$ | $CO_{2(g)}$ |
|---|---|---|
| 1 mol | 1 mol | |
| 12.01 g | 31.98 g | |
| 9.8 g | ? g | |

$$9.80 \text{ g C} \times \frac{31.98 \text{ g } O_2}{12.01 \text{ g C}} = 26.1 \text{ g } O_2$$

b) $C_{(s)} + O_{2(g)} \rightarrow CO_{2(g)}$

| $C_{(s)}$ | $O_{2(g)}$ | $CO_{2(g)}$ |
|---|---|---|
| 1 mol | | 1 mol |
| 12.01 g | | 43.99 g |
| 19.6 g | | ? g |

$$19.6 \text{ g C} \times \frac{43.99 \text{ g } CO_2}{12.01 \text{ g C}} = 71.8 \text{ g } CO_2$$

c) $C_{(s)} + O_{2(g)} \rightarrow CO_{2(g)}$

| $C_{(s)}$ | $O_{2(g)}$ | $CO_{2(g)}$ |
|---|---|---|
| | 1 mol | 1 mol |
| | 0.85 mol | ? mol |

$$0.85 \text{ mol } O_2 \times \frac{1 \text{mol } CO_2}{1 \text{mol } O_2} = 0.85 \text{ mol } CO_2$$

d) $C_{(s)} + O_{2(g)} \rightarrow CO_{2(g)}$

| $C_{(s)}$ | $O_{2(g)}$ | $CO_{2(g)}$ |
|---|---|---|
| 1 mol | 1 mol | |
| 12.01 g | 31.98 g | |
| ? g | 1.45 mol | |

$$1.45 \text{ mol } O_2 \times \frac{12.01 \text{ g C}}{1 \text{ mol } O_2} = 17.4 \text{ g C}$$

e) $C_{(s)} + O_{2(g)} \rightarrow CO_{2(g)}$

| $C_{(s)}$ | $O_{2(g)}$ | $CO_{2(g)}$ |
|---|---|---|
| 1 mol | 1 mol | |
| 12.01 g | $6.02 \times 10^{23}$ molecules | |
| 5.26 g | ? molecules | |

$$5.26 \text{ g C} \times \frac{6.02 \times 10^{23} \text{ molecules } O_2}{12.01 \text{ g C}} = 2.64 \times 10^{23} \text{ molecules } O_2$$

f) $C_{(s)} + O_{2(g)} \rightarrow CO_{2(g)}$

| $C_{(s)}$ | $O_{2(g)}$ | $CO_{2(g)}$ |
|---|---|---|
| 1 mol | | 1 mol |
| $6.02 \times 10^{23}$ atoms | | $6.02 \times 10^{23}$ molecules |
| $2.41 \times 10^{23}$ atoms | | ? molecules |

$$2.41 \times 10^{23} \text{ atoms C} \times \frac{6.02 \times 10^{23} \text{ molecules } CO_2}{6.02 \times 10^{23} \text{ atoms C}}$$

$$= 2.41 \times 10^{23} \text{ molecules } CO_2$$

The sample problem we just solved involved a stoichiometry in which one mol reacted with one mol to produce one mol. Not all chemical reactions have this simple one to one stoichiometry. Examine some problems where the stoichiometry is not as simple. These problems are solved using a similar method to sample 1 problems.

## SAMPLE PROBLEM 2

Given the reaction:

$Fe_2O_3 + 3CO \rightarrow 2Fe + 3CO_2$

a) How many g of $Fe_2O_3$ reacts with 26.75 g of CO?

b) How many g of $CO_2$ are produced from reacting 45.7 g of $Fe_2O_3$ with sufficient CO?

c) How many mol of Fe are produced from reacting 1.80 mol of CO?

d) How many g of $Fe_2O_3$ are needed to react with 6.50 mol of CO?

e) How many molecules of CO are needed to react with 74.85 g of $Fe_2O_3$?

f) How many atoms of Fe can be made by reacting $4.75 \times 10^{24}$ formula units of $Fe_2O_3$?

## Solution

a) $Fe_2O_3 + 3CO \rightarrow 2Fe + 3CO_2$

1 mol    3 mol

159.67 g    84.00 g

? g    26.75 g

$$26.75 \text{ g CO} \times \frac{159.67 \text{ g } Fe_2O_3}{84.00 \text{ g CO}} = 50.85 \text{ g } Fe_2O_3$$

b) $Fe_2O_3 + 3CO \rightarrow 2Fe + 3CO_2$

1 mol    3 mol

159.67 g    131.97 g

45.7 g    ? g

$$45.7 \text{ g } Fe_2O_3 \times \frac{131.97 \text{ g } CO_2}{159.67 \text{ g } Fe_2O_3} = 37.8 \text{ g } CO_2$$

c) $Fe_2O_3 + 3CO \rightarrow 2Fe + 3CO_2$

3 mol    2 mol

1.80 mol    ? mol

$$1.80 \text{ mol CO} \times \frac{2 \text{ mol Fe}}{3 \text{ mol CO}} = 1.20 \text{ mol Fe}$$

d) $Fe_2O_3 + 3CO \rightarrow 2Fe + 3CO_2$

1 mol    3 mol

159.67 g

? g    6.50 mol

$$6.50 \text{ mol CO} \times \frac{159.67 \text{ g } Fe_2O_3}{3 \text{ mol CO}} = 346 \text{ g } Fe_2O_3$$

e) $Fe_2O_3 + 3CO \rightarrow 2Fe + 3CO_2$

1 mol    3 mol

159.67 g    $1.81 \times 10^{24}$ molecules

74.85 g    ? molecules

$$74.85 \text{ g } Fe_2O_3 \times \frac{1.81 \times 10^{24} \text{ molecules CO}}{159.67 \text{ g } Fe_2O_3} = 8.48 \times 10^{23} \text{ molecules CO}$$

f) $Fe_2O_3 + 3CO \rightarrow 2Fe + 3CO_2$

1 mol    2 mol

$6.02 \times 10^{23}$ formula units    $1.20 \times 10^{24}$ atoms

$4.75 \times 10^{24}$ formula units    ? atoms

$$4.75 \times 10^{24} \text{ formula units } Fe_2O_3 \times \frac{1.20 \times 10^{24} \text{ atoms Fe}}{6.02 \times 10^{23} \text{ formula units } Fe_2O_3}$$

$$= 9.47 \times 10^{24} \text{ atoms Fe}$$

## Lesson Exercises

1. Given the following reaction:
   $N_2O_4 + 2N_2H_4 \rightarrow 3N_2 + 4H_2O$
   a) How many g of $N_2H_4$ reacts with 75.62 g of $N_2O_4$?
   b) How many g of $N_2$ are formed from reacting 48.25 g of $N_2H_4$?
   c) How many mol of $H_2O$ are produced by reacting 7.5 mol of $N_2H_4$?
   d) How many g of $N_2O_4$ are needed to react with 3.65 mol of $N_2H_4$?
   e) How many molecules of $N_2O_4$ are needed to react with 58.65 g of $N_2H_4$?
   f) How many molecules of $N_2$ can be made by reacting $3.65 \times 10^{21}$ molecules of $N_2O_4$?

2. Given the following reaction:
   $MnO_2 + 4HCl \rightarrow MnCl_2 + Cl_2 + 2H_2O$
   a) How many g of $MnO_2$ reacts with 45.76 g of HCl?
   b) How many g of $H_2O$ are formed from reacting 65.25 g of $MnO_2$?
   c) How many mol of $Cl_2$ are made by reacting 14.5 mol of HCl?
   d) How many g of HCl are needed to react with 4.25 mol of $MnO_2$?
   e) How many molecules of HCl are needed to react with 36.92 g of $MnO_2$?
   f) How many molecules of $H_2O$ can be made by reacting $6.00 \times 10^{24}$ molecules of HCl?
3. a) How many g of $Al_2(SO_4)_3$ can be made by reacting 15.86 g of Al with enough $H_2SO_4$ according to the following equation:
   $2Al_{(s)} + 3H_2SO_{4(aq)} \rightarrow Al_2(SO_4)_{3(aq)} + 3H_{2(g)}$
   b) How many g of Al reacts with 42.95 g of $H_2SO_4$?
   c) How many mol of $Al_2(SO_4)_3$ can be made from 6.00 mol of $H_2SO_4$?
4. Given the following reaction:
   $4FeS_{(s)} + 7O_{2(g)} \rightarrow 2Fe_2O_{3(s)} + 4SO_{2(g)}$
   a) How many g of $O_2$ reacts with 25.16 g of FeS?
   b) How many g of $Fe_2O_3$ can be made from 75.00 g of $O_2$?
   c) How many mol of $Fe_2O_3$ can be made from 12.00 mol of FeS?
5. How many g of $CaCl_2$ can be made from 14.85 g of HCl according to the following equation
   $CaO_{(s)} + 2HCl_{(aq)} \rightarrow CaCl_{2(aq)} + H_2O_{(l)}$
6. Given the following balanced chemical equation:
   $2VO_{(s)} + 3Fe_2O_3(s) \rightarrow 6FeO_{(s)} + V_2O_{5(s)}$
   a) State the mol-mol relationship in this reaction.
   b) State the mass-mass relationship in this reaction.
   c) Show that this balanced chemical equation obeys the law of conservation of mass.

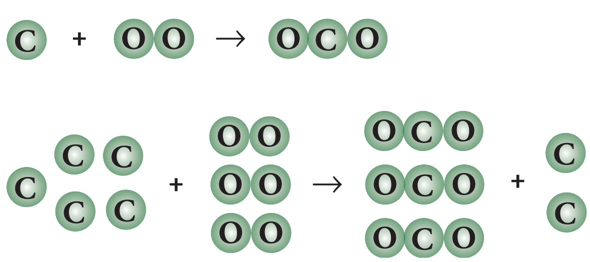
### JOURNAL ENTRY

# LIMITING REAGENT
## LESSON 9

In the last lesson, we calculated how much of one reactant was needed to react with a certain amount of another reactant. However, sometimes reactants are just mixed together and we try to predict how much product will be made. Under these conditions there is usually too much of one reactant present (a reactant in excess). In this case one of the reactants is used up completely. Some of the other reactant is used with some left over (in excess). The reactant that is completely used up is called the *limiting reagent* (reagent is a chemical). It is called the limiting reagent because it is completely used up and you cannot make any more product. It limits the amount of product that can be made.

Examine the following reactions and diagram to help understand the concept of limiting reagent.

$$C_{(s)} + O_{2(g)} \rightarrow CO_{2(g)}$$




We can see that one atom of carbon, C, reacts with one molecule of oxygen, OO, to make one molecule of carbon dioxide, OCO. If we reacted five carbon atoms, C, with three oxygen molecules, OO, it is only possible for three molecules of carbon dioxide to form.

Oxygen, $O_2$, OO, is the limiting reagent because it is completely used up and limits the amount of carbon dioxide product made. We cannot make any more $CO_2$ because there is no OO to react with the two carbon atoms in excess. It is easier to understand the concept of limiting reagent when we relate it to numbers of particles reacting. It is more difficult when talking about grams of reactants that are reacting. However, we have learned in this unit that grams of a chemical and numbers of particles are directly related to each other. In the same way so many particles of one reactant reacts with so many particles of another reactant, we know that a certain mass of one reactant reacts with a certain mass of another reactant. Examine some limiting reagent problems.

### SAMPLE PROBLEM 1

Given the reaction $C_{(s)} + O_{2(g)} \rightarrow CO_{2(g)}$

How many g of $CO_2$ can be made by reacting 16.45 g of C with 26.25 g of $O_2$?

| $C_{(s)}$ + | $O_{2(g)}$ → | $CO_{2(g)}$ |
|---|---|---|
| 1 mol | 1 mol | 1 mol |
| 12.01 g | 31.98 g | 43.99 g |
| 16.45 g | 26.25 g | ? g |

### SOLUTION

One approach to solving a limiting reagent problem is to assume that one of the reactants is the limiting reagent. It doesn't matter which reactant you choose. Then calculate to see if there is enough of the other reactant to completely use up the reactant assumed to be the limiting reagent. In this question, assume that carbon is the limiting reagent. Next, calculate to see if the assumption is correct.

$$16.45 \text{ g C} \times \frac{31.98 \text{ g } O_2}{12.01 \text{ g C}} = 43.80 \text{ g } O_2$$

Our calculation shows that if all of the C is to be used up, making C the limiting reagent, it will need to react with 43.80 g $O_2$. However, we only have 26.25 g of $O_2$ available. Therefore, C cannot be the limiting reagent. The other reactant, $O_2$, is the limiting reagent. If you can prove that one of the reactants is not the limiting reagent, the other reactant is the limiting reagent. We do not have to show that $O_2$ is the limiting reagent now that we have shown that C is not the limiting reagent.

However, for demonstration purposes, look at the calculations as if we had assumed in the beginning that $O_2$ was the limiting reagent. Calculate how much C we would need to use up the $O_2$ completely.

$$26.25 \text{ g } O_2 \times \frac{12.01 \text{ g C}}{31.98 \text{ g } O_2} = 9.858 \text{ g C}$$

We need 9.858 g of C to use up all of the $O_2$ and make $O_2$ the limiting reagent. We have 16.45 g of C to react, more than enough. This calculation proves that $O_2$ is the limiting reagent. Now that we know which reactant is used up completely, we can calculate how much product can be made. If we used 16.45 g of C (the reactant in excess) to determine how much product is made, the amount calculated would be incorrect and more than the actual amount we can make in this reaction.

$$26.25 \text{ g } O_2 \times \frac{43.99 \text{ g } CO_2}{31.98 \text{ g } O_2} = 36.11 \text{ g } CO_2$$

Therefore, 36.11 g $CO_2$ will be made in this reaction.

## SAMPLE PROBLEM 2

What is the maximum amount of $Fe_2O_3$ that can be made by reacting 50.00 g of FeS with 50.00 g of $O_2$ according to the following reaction?

| $4FeS_{(s)}$ + | $7O_{2(g)}$ → | $2Fe_2O_{3(s)}$ + | $4SO_{2(s)}$ |
|---|---|---|---|
| 4 mol | 7 mol | 2 mol | |
| 351.68 g | 223.86 g | 319.34 g | |
| 50.00 g | 50.00 g | ? g | |

## SOLUTION

Assume that FeS is the limiting reagent, (LR).

$$50.00 \text{ g FeS} \times \frac{223.86 \text{ g } O_2}{351.68 \text{ g FeS}} = 31.83 \text{ g } O_2$$

If FeS is the limiting reagent we would need 31.83 g of $O_2$ to use up the FeS completely. There is 50.00 g of $O_2$ available to react, more than enough. Therefore, FeS is the limiting reagent.

$$50.00 \text{ g FeS} \times \frac{319.34 \text{ g } Fe_2O_3}{351.68 \text{ g FeS}} = 45.40 \text{ g } Fe_2O_3$$

The maximum amount of $Fe_2O_3$ that can be made by reacting 50.00 g of FeS with 50.00 g of $O_2$ is 45.40 g of $Fe_2O_3$.

## SAMPLE PROBLEM 3

Given the reaction:

$Li_3N_{(s)} + 3H_2O_{(l)} \rightarrow 3LiOH_{(s)} + NH_{3(g)}$

How much LiOH can be made by reacting 28.00 g of $Li_3N$ with 32.00 g $H_2O$?

| $Li_3N_{(s)}$ + | $3H_2O_{(l)}$ → | $3LiOH_{(s)}$ + | $NH_{3(g)}$ |
|---|---|---|---|
| 1 mol | 3 mol | 3 mol | |
| 34.83 g | 54.03 g | 71.82 g | |
| 28.00 g | 32.00 g | ? g | |

## SOLUTION

Assume that $Li_3N$ is the limiting reagent.

$$28.00 \text{ g } Li_3N \times \frac{54.03 \text{ g } H_2O}{34.83 \text{ g } Li_3N} = 43.43 \text{ g } H_2O$$

If $Li_3N$ is the limiting reagent, we need 43.43 g of $H_2O$ to use up the $Li_3N$ completely. However, we only have 32.00 g of $H_2O$ available to react. Therefore, $Li_3N$ is not the limiting reagent. The $H_2O$ is the limiting reagent.

$$32.00 \text{ g } H_2O \times \frac{71.82 \text{ g LiOH}}{54.03 \text{ g } H_2O} = 42.54 \text{ g LiOH}$$

Therefore, 42.54 g of LiOH can be made from reacting 28.00 g of $Li_3N$ with 32.00 g $H_2O$.

## SAMPLE PROBLEM 4

Given the reaction:

$$CH_{4(g)} + 2O_{2(g)} \rightarrow CO_{2(g)} + 2H_2O_{(g)}$$

a) How much $CO_2$ can be made by reacting 1.68 mol of $CH_4$ with 45.27 g of $O_2$?

b) How many g of the reactant in excess will be left over unreacted?

| $CH_{4(g)}$ + | $2O_{2(g)}$ → | $CO_{2(g)}$ + | $2H_2O_{(g)}$ |
|---|---|---|---|
| 1 mol | 2 mol | 1 mol | |
| 16.05 g | 63.96 g | 43.99 g | |
| 1.67 mol | 45.27 g | ? g | |

## SOLUTION

a) Assume that $CH_4$ is the limiting reagent.

$$1.68 \text{ mol } CH_4 \times \frac{63.96 \text{ g } O_2}{1 \text{mol } CH_4} = 107 \text{ g } O_2$$

We need 107 g of $O_2$ to use up all of the $CH_4$ if $CH_4$ is the limiting reagent. However, we have only 45.27 g $O_2$ available to react. The $CH_4$ cannot be the limiting reagent. The $O_2$ is the limiting reagent.

$$45.27 \text{ g } O_2 \times \frac{43.99 \text{ g } CO_2}{63.96 \text{ g } O_2} = 31.14 \text{ g } CO_2$$

Therefore, 31.14 g $CO_2$ can be made by reacting 1.68 mol of $CH_4$ with 45.27 g $O_2$

b) $$45.27 \text{ g } O_2 \times \frac{16.05 \text{ g } CH_4}{63.96 \text{ g } O_2} = 11.36 \text{ g } CH_4$$

11.36 g $CH_4$ needed

$$1.68 \text{ mol } CH_4 \times \frac{16.05 \text{ g } CH_4}{1 \text{ mol } CH_4} = 27.0 \text{ g } CH_4$$

27.0 g $CH_4$ available

27.0 g $CH_4$ - 11.36 g $CH_4$ = 15.6 g $CH_4$ left over, unreacted.

## LESSON EXERCISES

1. Given the reaction:

   $$Mg_{(s)} + 2HCl_{(aq)} \rightarrow MgCl_{2(aq)} + H_{2(g)}$$

   How many g of $MgCl_2$ can be made by reacting 18.00 g of Mg with 18.00 g of HCl?

2. Given the following reaction:

   $$(NH_4)_2SO_4 + Pb(NO_3)_2 \rightarrow 2NH_4NO_3 + PbSO_4$$

   How much $NH_4NO_3$ can be made by reacting 135.0 g of $(NH_4)_2SO_4$ with 160.0 g of $Pb(NO_3)_2$?

3. Given the reaction:

   $2Na_3PO_4 + 3BaCl_2 \rightarrow Ba_3(PO_4)_2 + 6NaCl$

   What is the maximum amount of NaCl that can be produced by reacting 36.50 g of $Na_3PO_4$ with 82.75 g of $BaCl_2$?

4. Given the reaction:

   $P_4 + 5O_2 \rightarrow P_4O_{10}$

   a) How much product can be made by reacting 47.35 g of $P_4$ with 123.45 g of $O_2$?

   b) How many g of the reactant that is in excess will be left over unreacted?

5. Given the reaction:

   $SF_4 + 2H_2O \rightarrow SO_2 + 4HF$

   How much HF can be produced by reacting 35.14 g of $SF_4$ with 0.75 mol of $H_2O$?

## JOURNAL ENTRY

# PERCENT YIELD OF A CHEMICAL REACTION
## LESSON 10

In this unit we calculated how much product could be made if we reacted a certain amount of reactants, or if the limiting reagent was completely consumed. This yield of products we calculated is the predicted or theoretical yield under ideal conditions. *Theoretical yield* is the amount of products predicted in a chemical reaction based on stoichiometry calculations and assuming a complete reaction of all of the limiting reagent. The theoretical yield, however, is not usually equal to the actual yield for a chemical reaction. *Actual yield* is the measured amount of products made. What would make the actual yield less than the theoretical yield for a reaction? There are a number of possible reasons.

1. One of the major reasons is the "competing reaction" or "side reaction" that is taking place. Examine the following reaction:

   *main reaction* $2X + 3Y_2 \rightarrow 2XY_3$

   We are trying to produce $XY_3$. We calculate the theoretical yield of $XY_3$ based on how much of the reactants react. If there is a competing or side reaction

   *competing reaction* $XY_3 + Y_2 \rightarrow XY_5$

   in which some of the $XY_3$ produced is further reacted to make $XY_5$, then the actual yield will be lowered. There will be less $XY_3$ to be measured in the container after the chemical reactions.

2. Another possible reason for the actual yield to be less than the theoretical yield is that one of the reactants or products may be very volatile and evaporate.
3. Sometimes some of a product might stick to the inside walls of a glass container used.
4. If an insoluble product (precipitate) is produced and filtered to separate it from the solvent, a small amount of this product actually dissolves in the solvent and represents lost product in our measurements.

When chemicals are made in huge quantities in the chemical industry it is useful to know the per cent yield for a chemical reaction.

$$\text{per cent yield} = \frac{\text{actual yield}}{\text{theoretical yield}} \times 100$$

This is important because it shows exactly how much chemicals have to be reacted to make a certain amount of products. The per cent yield tells us how close the actual amount of products made is to the predicted maximum theoretical amount. The higher the per cent yield, the better.

### SAMPLE PROBLEM 1

Given the following reaction:

$Fe_{(s)} + S_{(s)} \rightarrow FeS_{(s)}$

Calculate the per cent yield for this reaction if 33.75 g of FeS was made when 23.95 g of Fe was completely used up by an excess of S.

| $Fe_{(s)}$ | + $S_{(s)}$ | $\rightarrow FeS_{(s)}$ |
|---|---|---|
| 1 mol | 1 mol | 1 mol |
| 55.85 g | 32.07 g | 87.92 g |

### SOLUTION

$$\text{percent yield} = \frac{\text{actual yield}}{\text{theoretical yield}} \times 100$$

We know the actual yield, but we must calculate the theoretical yield.

$$23.95 \text{ g Fe} \times \frac{87.92 \text{ g FeS}}{55.85 \text{ g Fe}} = 37.70 \text{ g FeS}$$

The theoretical yield is 37.70 g FeS. This is the maximum amount of product we could hope to make.

$$\text{percent yield} = \frac{33.75 \text{ g FeS}}{37.70 \text{ g FeS}} \times 100 = 89.52\%$$

### SAMPLE PROBLEM 2

Given the following reaction

$$3Fe + 4H_2O \rightarrow Fe_3O_4 + 4H_2$$

Calculate the percent yield if 198.0 g of $Fe_3O_4$ was produced when 75.00 g of $H_2O$ was reacted with an excess of Fe.

### SOLUTION

| $3Fe$ + | $4H_2O$ → | $Fe_3O_4$ + | $4H_2$ |
|---|---|---|---|
| | 4 mol | 1 mol | |
| | 72.04 g | 231.51 g | |

The theoretical yield is:

$$75.00 \text{ g } H_2O \times \frac{231.51 \text{ g } Fe_3O_4}{72.04 \text{ g } H_2O} = 241.0 \text{ g } Fe_3O_4$$

$$\text{percent yield} = \frac{198.0 \text{ g } Fe_3O_4}{241.0 \text{ g } Fe_3O_4} \times 100 = 82.16\%$$

### SAMPLE PROBLEM 3

Given the following reaction:

$$H_2S_{(g)} + 2NaOH_{(aq)} \rightarrow Na_2S_{(aq)} + 2H_2O_{(l)}$$

A mass of 6.75 g of NaOH was reacted with 4.50 g of $H_2S$. One of the products, $Na_2S$, had a mass of 5.95 g. Calculate the percent yield.

### SOLUTION

| $H_2S_{(g)}$ + | $2NaOH_{(aq)}$ → | $Na_2S_{(aq)}$ + | $2H_2O_{(l)}$ |
|---|---|---|---|
| 1 mol | 2 mol | 1 mol | |
| 34.09 g | 79.98 g | 78.05 g | |
| 6.75 g | 4.50 g | ? g | |

Before we can calculate the percent yield we must determine which reactant is the limiting reagent and then calculate the theoretical yield. Assume that NaOH is the limiting reagent.

$$6.75 \text{ g NaOH} \times \frac{34.09 \text{ g } H_2S}{79.98 \text{ g NaOH}} = 2.88 \text{ g } H_2S$$

If NaOH is the limiting reagent we need 2.88 g of $H_2S$ to use it up completely. We have 4.50 g of $H_2S$ available, more than enough. Therefore, NaOH is the limiting reagent. The theoretical yield is:

$$6.75 \text{ g NaOH} \times \frac{78.05 \text{ g } Na_2S}{79.98 \text{ g NaOH}} = 6.59 \text{ g } Na_2S$$

$$\text{percent yield} = \frac{5.95 \text{ g } Na_2S}{6.59 \text{ g } Na_2S} \times 100 = 90.3\%$$

### LESSON EXERCISES

1. Given the following reaction

$$AgNO_{3(aq)} + KBr_{(aq)} \rightarrow AgBr_{(s)} + KNO_{3(aq)}$$

   Calculate the per cent yield if 7.15 g of AgBr was produced when 7.00 g of $AgNO_3$ was reacted with an excess of KBr.

2. Aspirin, $C_9H_8O_4$, is produced by reacting salicylic acid, $C_7H_6O_3$, with acetic

anhydride, $C_4H_6O_3$, according to the following chemical equation:

$$C_7H_6O_3 + C_4H_6O_3 \rightarrow C_9H_8O_4 + C_2H_4O_2$$

Calculate the per cent yield if 8.00 g of $C_7H_6O_3$ is reacted with 16.00 g of $C_4H_6O_3$ and the actual yield of aspirin is 8.72 g.

3. Given the reaction:

$$CaO_{(s)} + 2HCl_{(aq)} \rightarrow CaCl_{2(aq)} + H_2O_{(l)}$$

Calculate the per cent yield if 42.50 g of CaO was reacted completely with an excess of HCl and 78.68 g of $CaCl_2$ was made.

4. Given the reaction:

$$Li_3N_{(s)} + 3H_2O_{(l)} \rightarrow 3LiOH_{(s)} + NH_{3(g)}$$

Calculate the per cent yield if 48.25 g of $Li_3N$ was reacted completely with an excess of $H_2O$ to make 91.36 g of LiOH.

JOURNAL ENTRY

# APPLICATIONS OF STOICHIOMETRY
## LESSON 11

In industry, chemicals are produced in large-scale amounts. Instead of producing 30 g or 275 g of products in a high school laboratory, products in large chemical plants are produced by the millions and billions of kilograms. Large amounts of money are involved. Close attention is paid to the stoichiometry of chemical reactions; how much has to be reacted to make a certain quantity of product. Maximizing the yield of chemical reactions increases profit. This lesson examines stoichiometry examples of chemical reactions used in industry.

The concept of the mole and stoichiometry has allowed us to do exact calculations on how much chemicals to react and how much product to expect. Without the mole and stoichiometry we would not know how much chemicals to react and how much product to expect. We would be just mixing random amounts of reactants and at times having large excesses of a reactant left over, and problems purifying the reaction system. This would be both costly and inefficient. Also, the mole and stoichiometry are important concepts when dealing with reactions we use in our daily lives, like the operation of an automobile safety airbag, neutralizing excess stomach acid, removing $CO_2$ from a spacecraft, operation of a Breathalyser, and reducing acid rain.

One of the chemical reactions that takes place when an automobile air safety bag inflates is:

$$2NaN_{3(s)} \rightarrow 2Na_{(s)} + 3N_{2(g)}$$

The total time for the airbag to fully inflate is only about 30 milliseconds or 0.030 seconds. The hot $N_{2(g)}$ produced fills the airbag. Scientists need to know what mass of $NaN_3$ is needed to produce the amount of $N_{2(g)}$ required to fully inflate the airbag.

The most commonly produced chemical in the world is sulphuric acid. This chemical has widespread uses, including the cleaning of metals, drugs, paints, storage batteries, explosives, fertilizers, etc. The production of sulphuric acid, $H_2SO_{4(l)}$, takes place in a number of steps. The first step involves burning nearly pure sulphur in air to produce sulphur dioxide:

$$S_{8(s)} + 8O_{2(g)} \rightarrow 8SO_{2(g)}$$

The sulphur dioxide is then reacted with oxygen at a temperature of 400°C, using a catalyst like $V_2O_5$ to produce sulphur trioxide:

$$2SO_{2(g)} + O_{2(g)} \rightarrow 2SO_{3(g)}$$

The sulphur trioxide is cooled to 100°C and then passed into a tower so it can be absorbed into 97 per cent sulphuric acid, producing pyrosulphuric acid:

$$H_2SO_{4(l)} + SO_{3(g)} \rightarrow \underset{\text{(pyrosulfuric acid)}}{H_2S_2O_{7(s)}}$$

When water is added to the pyrosulphuric acid, we obtain sulfuric acid:

$$H_2S_2O_{7(s)} + H_2O_{(l)} \rightarrow 2H_2SO_{4(l)}$$

There is another process used for the production of sulfuric acid. It is called the lead-chamber process and only accounts for 20 per cent or less of all sulfuric acid produced. This illustrates that sometimes there is more than one way to produce a chemical. There are a number of factors that will dictate which process is followed. These factors include:

- Availability of reactants
- Cost of reactants
- Cost of any catalysts used
- Per cent yield
- Rates of reactions
- Effect on the health of the workers
- Cost to deal with waste products

Once again, it is important for scientists to know the stoichiometry of these reactions and per cent yields if they are to produce the desired quantities.

When we want to neutralize excess stomach acidity, we take an antacid. Antacids usually contain one of the following: hydroxides, carbonates, or bicarbonates. Some typical reactions of antacids with excess stomach acidity are:

$$Mg(OH)_{2(s)} + 2HCl_{(aq)} \rightarrow MgCl_{2(aq)} + 2H_2O_{(l)}$$

$$CaCO_{3(s)} + 2HCl_{(aq)} \rightarrow CaCl_{2(aq)} + H_2O_{(l)} + CO_{2(g)}$$

$$NaHCO_{3(aq)} + HCl_{(aq)} \rightarrow NaCl_{(aq)} + H_2O_{(l)} + CO_{2(g)}$$

$$Al(OH)_{3(s)} + 3HCl_{(aq)} \rightarrow AlCl_{3(aq)} + 3H_2O_{(l)}$$

We can see that the stoichiometry can be quite different for the reactions. Thus, the number of moles of HCl acid that can be neutralized by one mol of antacid varies. Also, in taking antacids, we want to take enough antacid to do the job but not too much antacid, which may be harmful.

Aspirin, (acetylsalicylic acid) $C_9H_8O_4$, can be made according to the reaction:

$$\underset{\text{salicylic acid}}{C_7H_6O_3} + \underset{\text{acetyl chloride}}{C_2H_3OCl} \rightarrow \underset{\text{acetylsalicylic acid}}{C_9H_8O_4} + \underset{\text{hydrochloric acid}}{HCl}$$

However, commercial production of aspirin involves the reaction:

$$\underset{\text{acetic anhydride}}{C_4H_6O_3} + \underset{\text{salicylic acid}}{C_7H_6O_3} \rightarrow \underset{\text{aspirin}}{C_9H_8O_4} + \underset{\text{acetic acid}}{C_2H_4O_2}$$

Once again, a number of factors have to be considered when deciding which process to use.

The breathalyser uses the reaction:

$$8H^+ + \underset{\text{dichromate ion (orange)}}{Cr_2O_7^{2-}} + \underset{\text{ethyl alcohol}}{3C_2H_5OH} \rightarrow \underset{\text{chromium ion (green)}}{2Cr^{3+}} + 3C_2H_4O + 7H_2O$$

The ethyl alcohol, $C_2H_5OH$, in the breath reacts with the orange dichromate ions $Cr_2O_7^{2-}$ to produce green, $Cr^{3+}$ ions. The more alcohol in your breath, the more green $Cr^{3+}$ ions are produced. The colour intensity of the $Cr^{3+}$ is related to how much alcohol is in your breath and blood.

One way of reducing acid rain is to reduce the amount of $SO_2$ produced when burning coal or oil that contains sulphur. One method involves using a "flue scrubber". In this technique, powdered limestone, $CaCO_3$ is blown into the combustion chamber, where the heat decomposes it into calcium oxide and carbon dioxide:

$$CaCO_{3(s)} \rightarrow CaO_{(s)} + CO_{2(g)}$$

The CaO then combines with the $SO_{2(g)}$ to produce calcium sulphite:

$$CaO_{(s)} + SO_{2(g)} \rightarrow CaSO_{3(s)}$$

Stoichiometric calculations are quite important here to know how much limestone should be used each time to reduce the $SO_2$ emissions to a minimum.

## LESSON EXERCISES

1. Aspirin, $C_9H_8O_4$, is made commercially from the reaction of salicylic acid, $C_7H_6O_3$, and acetic anhydride, $C_4H_6O_3$, according to the following reaction:

$$C_7H_6O_3 + C_4H_6O_3 \rightarrow C_9H_8O_4 + C_2H_4O_2$$

   a) If the salicylic acid is used up completely when converted to aspirin, how much salicylic acid is necessary to produce 150 kg (1 kg = 1000 g) of aspirin?

   b) If the reaction had a 76 per cent yield, how much salicylic acid would be required?

   c) If the hypothetical cost for salicylic acid was \$15.00/kg and acetic anhydride was \$18.00/kg, which compound would you choose as the limiting reagent in order to keep the costs of production down to a minimum?

   d) What is the theoretical yield of aspirin if 210 kg of salicylic acid are allowed to react with 145 kg of acetic anhydride?

   e) If the actual yield in part d) was 207 kg, what is the per cent yield?

   f) What would you have to charge per kilogram of aspirin to cover just the costs of the chemicals?

   g) What mass of salicylic acid would be needed to produce 10 aspirin tablets, each with a mass of 0.350 g?

## JOURNAL ENTRY

# Measurement
## Lesson 12

A measurement consists of a number and a unit. The unit indicates what we are measuring. Scientists use the *SI system* (Le Systeme International d'Unites) of measurement. Canada has adopted this system.

Whenever a measurement is made there is uncertainty associated with it. If we measure the mass of an object to two decimal places, we would wonder what the third decimal place was. If we measure the mass of an object to five decimal places, we would be uncertain what the sixth decimal place was. A radar gun used to catch speeding drivers has an associated uncertainty. If the uncertainty was ± 2 km/h for the radar gun and you were told you were measured at 56 km/h, then you could have been travelling at a speed of anywhere from -2 (54 km/h) to +2 (58 km/h).

The ± notation gives us information about the uncertainty in the measurement and the possible range of values for the measurement. For instance, a balance that measures ±0.001 g is more certain than a balance that only measures ±0.1 g. Each measuring instrument, whether it be a graduated cylinder, thermometer, or balance, has its own measurement uncertainty. *Precision* is a measure of how well similar measurements agree with one another. In other words, how reproducible are the measurements?

The smaller the uncertainty in the measuring device, the more precise its measurements will be.

Examine five mass measurements made on the same object. They are listed below:

2.04 g
2.02 g
2.06 g
2.06 g
2.02 g

The average is 2.04 g. The range is from 2.02 g to 2.06 g. Therefore, we can express the measurement as 2.04 g ± 0.02.

Another term associated with measurement is *accuracy*. Accuracy is how close the measurement or average measurement is to the true accepted value. For instance, if the mass of an object was actually 95.34 g, then a measurement of 95.13 g is more accurate than a measurement of 94.05 g. One way of expressing how close your measurement is to the accepted value (the accuracy), is to calculate the per cent error.

Per cent error = ((absolute value of the difference in accepted value and measured value) divided by the accepted value) multiplied by 100.

$$\text{Per cent error} = \frac{\text{accepted value - measured value}}{\text{accepted value}} \times 100$$

You want a small per cent error to reflect how well you did the experiment.

### Sample Problem 1

Calculate the percent error if the accepted value is 39.87 kJ and the measured value is 37.45 kJ.

$$\text{per cent error} = \frac{39.87\text{ kJ} - 37.45\text{ kJ}}{39.87\text{ kJ}} \times 100$$

$$= 6.07\%$$

## LESSON EXERCISES

1. The following measurements were made on the same sample:

    37.28°C

    37.28°C

    37.24°C

    37.32°C

    37.24°C

    37.32°C

    a) Calculate the average of the measurements.

    b) Express the precision of the measurements using ± notation.

    c) What is the range of the measurements?

    d) If the accepted value is 37.31°C, calculate the percent error and comment on the accuracy of the measurements.

2. Is it possible to have good precision, but poor accuracy? Explain.

3. Is it possible to have poor precision, but good accuracy? Explain.

## JOURNAL ENTRY

# Do and Send

## *Stoichiometry Assignment Questions*

1. Natural sulphur consists of 95.0 per cent $^{32}_{16}S$, actual mass 31.972 07 u; 0.76% $^{33}_{16}S$, actual mass 32.971 46 u; 4.22 per cent $^{34}_{16}S$, actual mass 33.967 86 u; and 0.014 per cent $^{36}_{16}S$, actual mass 35.967 09 u. Calculate the average atomic mass of natural sulphur.

2. a) What is the atomic mass of Pb and I?

   b) What is the molecular mass of $PF_3$ and $C_3H_6FCl$?

   c) What is the molar mass of Zn, $Ba_3(PO_4)_2$, and $SiBr_2F_2$?

3. a) How many g and atoms are there in 0.820 mol of Cs?

   b) How many mol and molecules are there in 24.62 g of $C_4H_{10}$?

   c) How many g and formula units are there in 0.480 mol of $CaBr_2$?

   d) How many mol and g are present in $3.75 \times 10^{24}$ molecules of $C_5H_{12}$?

4. Calculate the percentage composition of $(NH_4)_2SO_4$.

5. Calculate the percentage composition of $KNO_3$.

6. Determine the empirical formula of a compound which on analysis shows 48.38 per cent C; 8.12 per cent H; and 43.50 per cent O by mass.

7. Determine the empirical formula of a compound that contains 33 per cent Na; 13 per cent Al; and 54 per cent F by mass.

8. A certain compound has an empirical formula of $CH_2F$ and a molecular mass of 66.04 u. What is the molecular formula?

9. A compound by analysis was found to contain 52.25 per cent C; 13.0 per cent H; and 34.8 per cent O. The molecular mass is 91.6 u. Determine the molecular formula.

10. Given the following reaction $2Fe_{(s)} + O_{2(g)} \rightarrow 2FeO_{(s)}$

    a) How many g of $O_2$ reacts with 137.42 g of Fe?

    b) How many g of FeO are formed from reacting 65.75 g of $O_2$?

    c) How many mol of FeO are produced from reacting 2.50 mol of $O_2$?

    d) How many g of $O_2$ are needed to react with 7.56 mol of Fe?

    e) How many molecules of $O_2$ are needed to react with 42.65 g of Fe?

    f) How many formula units of FeO can be made by reacting $5.75 \times 10^{22}$ atoms of Fe?

11. Given the following reaction $Al_2(SO_4)_3 + 3Ca(OH)_2 \rightarrow 3CaSO_4 + 2Al(OH)_3$
    a) How many g of $Al_2(SO_4)_3$ reacts with 76.45 g of $Ca(OH)_2$?
    b) How many g of $CaSO_4$ are formed from reacting 32.56 g of $Ca(OH)_2$?
    c) How many mol of $Al(OH)_3$ are produced by reacting 5.0 mol of $Ca(OH)_2$?
    d) How many g of $Ca(OH)_2$ are needed to react with 1.75 mol of $Al_2(SO_4)_3$?
    e) How many formula units of $Al_2(SO_4)_3$ are needed to react with 18.75 g of $Ca(OH)_2$?
    f) How many formula units of $Al(OH)_3$ are produced by reacting $2.76 \times 10^{24}$ formula units of $Ca(OH)_2$?
12. Given the reaction $N_2 + 3H_2 \rightarrow 2NH_3$. How many g of $NH_3$ can be made by reacting 25.65 g of $N_2$ with 13.46 g of $H_2$?
13. Given the reaction $2NaHCO_3 + H_2SO_4 \rightarrow 2CO_2 + Na_2SO_4 + 2H_2O$
    a) What is the maximum amount of $Na_2SO_4$ that can be made by reacting 145.0 g of $NaHCO_3$ with 165.0 g of $H_2SO_4$?
    b) How many g of the reactant that is in excess will be left over unreacted?
14. Given the reaction $2Fe_{(s)} + O_{2(g)} \rightarrow 2FeO_{(s)}$. When 42.65 g of Fe is fully reacted, the actual yield was 45.00 g of FeO. Calculate the percent yield for the reaction.
15. Given the reaction $CO_2 + 2LiOH \rightarrow Li_2CO_3 + H_2O$. When 65.0 g of $CO_2$ was fully reacted, the actual yield of $Li_2CO_3$ was 65.92 g. Calculate the percent yield for the reaction.
16. Phosphoric acid, $H_3PO_{4(l)}$, is made by reacting concentrated sulfuric acid with pulverized phosphate rock, $Ca_3(PO_4)_{2(s)}$ in one of a number of steps, according to the reaction:

    $Ca_3(PO_4)_{2(s)} + 3H_2SO_{4(l)} \rightarrow 2H_3PO_{4(l)} + 3CaSO_{4(s)}$

    a) State the stoichiometry of the reaction in terms of moles of each reactant reacting and moles of each product being formed.
    b) State the stoichiometry of the reaction in terms of grams of each reactant reacting and grams of each product being formed.
17. Natural lithium consists of 7.42 per cent $^{6}_{3}Li$, actual mass 6.015 12 u; and 92.58 per cent $^{7}_{3}Li$, actual mass 7.016 00 u. Calculate the average atomic mass of natural lithium.

18. Given the following two sets of measurements:

| Set A | Set B |
|---|---|
| 14.56 g | 17.98 g |
| 14.48 g | 16.15 g |
| 14.64 g | 19.37 g |
| 14.56 g | 18.61 g |
| 14.60 g | 17.50 g |

The accepted value for set A measurements is 13.85 g, and the accepted value for set B measurements is 18.02 g.

a) Which set of measurements has better precision? Explain.

b) Which set of measurements has better accuracy? Explain.

c) Calculate the percent error for each set of measurements.

## DO AND SEND

### *STOICHIOMETRY RESEARCH ASSIGNMENT*

1. Research an industrial process (different from any mentioned in this unit) that produces a chemical.

a) Include the balanced chemical equations for the reactions.

b) State the stoichiometry for each chemical reaction in terms of how many moles of each reactant reacted and how many moles of each product was formed.

c) State the stoichiometry for each chemical reaction in terms of how many grams of each reactant is reacting and how many grams of each product is being formed.

# EXPERIMENT 1
# DO AND SEND

## TITLE

Limiting Reagent in a Chemical Reaction

## PURPOSE

To study a chemical reaction that illustrates the concept of limiting reagent

## CHEMICALS

One box of baking soda (sodium hydrogen carbonate), 2.0 L container of white vinegar (5% acetic acid by volume)

## EQUIPMENT

Measuring spoons, large bowl or tall 500 mL glass

## DATA TABLE

See above right

| Action | Observations |
|---|---|
| **Part A** Addition of 5 mL of baking soda to 120 mL of vinegar | |
| **Part B** Addition of 5 mL more of the baking soda | |
| **Part C** Addition of 5 mL more of the baking soda | |
| **Part D** Addition of 5 mL more of the baking soda | |
| **Part E** Addition of 5 mL more of the baking soda | |
| **Part F** Addition of 60 mL of vinegar to the above mixture | |
| **Part G** | |
| **Part H** | |

## PROCEDURE

The balanced chemical equation for the reaction between the sodium hydrogen carbonate in baking soda and the acetic acid in vinegar is

$$NaHCO_{3(aq)} + CH_3COOH_{(aq)} \rightarrow NaCH_3COO_{(aq)} + CO_{2(g)} + H_2O_{(l)}$$

where (aq) indicates dissolved in water, (g) indicates gas state, and (l) indicates liquid state.

The limiting reagent is the reactant that is completely used up in a reaction. Since it is all used, it limits the amount of product that can be made.

## PART A

The experiment should be done in the sink as some of the liquid may bubble out of the container. Add 120 mL of vinegar to a tall 500 mL glass or large bowl. Add 5 mL of baking soda to the vinegar and swirl the liquid until there is no fizzing or bubbling observed. The fizzing or bubbling is from carbon dioxide gas being produced. Record your observations in the table data.

## PART B

Add another 5 mL of baking soda to the container in part A, swirl the liquid until there is no more fizzing or bubbling. Record your observations in the data table.

## Part C

Add another 5 mL of baking soda to the container in part B, swirl the liquid until there is no more fizzing or bubbling. Record your observations in the data table.

## Part D

Add another 5 mL of baking soda to the container in part C, swirl the liquid until there is no more fizzing or bubbling. Record your observations in the data table.

## Part E

Add another 5 mL of baking soda to the container in part D, swirl the liquid until there is no more fizzing or bubbling. Record your observations in the data table.

## Part F

Add 60 mL of vinegar to the container in part E, swirl the liquid until there is no more fizzing or bubbling. Record your observations in the data table.

## Part G

Add another 60 mL of vinegar to the container in part F, swirl the liquid until there is no more fizzing or bubbling. Record your observations in the data table.

## Part H

Keep adding 60 mL measures of vinegar to the container and swirling until there is no fizzing or bubbling taking place. Record your observations in the data table for each additional vinegar measure.

## Questions

1. At what point could the vinegar be considered the limiting reagent in the experiment? Be specific and indicate how much baking soda and vinegar had been mixed at that point. Why was the vinegar considered the limiting reagent at that point?
2. At what point could the baking soda be considered the limiting reagent in the experiment? Be specific and indicate how much baking soda and vinegar had been mixed at that point. Why was the baking soda considered the limiting reagent at that point?
3. Why did fizzing still occur even after all the solid baking soda disappeared?
4. What are three questions you have as a result of doing this experiment?

## Conclusion

What conclusions did you make as a result of doing this experiment?