Basics You Should Know

All of chemistry is based on measurements of some sort to determine how much of something is present. This dates back to the 1700's when Lavoisier discovered the Law of Conservation of Mass.

Law of Conservation of Mass: The total mass remains constant during a chemical reaction. No mass is lost or gained.

Examples:

a. 12.4 grams of oxygen reacts with hydrogen to produce 14.0 grams of water. How **many grams of hydrogen** reacted with the oxygen?

Mass in = mass out

12.4 g oxygen + g hydrogen = 14.0 g water
g hydrogen = 14.0 g water – 12.4 g oxygen = 1.6 g hydrogen

b. If 22.4 g of phosphorus reacts with 153.9 g of chlorine, **how many grams of product** are formed?

22.4 g + 153.9 g = 176.3 g of product.

Matter

States of matter:

Solid- definite shape and volume (neither changes)

Liquid- indefinite shape but definite volume (liquids conform to the shape of their container)

Gas- indefinite shape and volume (gases conform to the shape and volume of their container)

Changes in matter

Physical change- A change that does not alter the chemical compound present.

Examples:

Boiling, Freezing, dissolving, distilling

Chemical change - A change the does alter the compounds present.

Examples:

Any type of chemical reaction. Burning, decomposing

Physical property- a property that can be determined by observing a substance and does not require a chemical change.

Examples:

Color, temperature, hardness

Chemical property- a property that does require a chemical change to determine.

Examples:

Does or does not react with a chemical

Substances- A kind of matter that cannot be separated into its constituents by physical processes. Two kinds of substances:

Elements- simplest type of substance. Comprised of only one kind of atom. Cannot be separated into any components by physical or chemical processes.

Compound- A chemical combination of two or more elements. Can be broken down by chemical processes.

Law of definite proportions – A chemical compound will always have the same ratio of elements in it regardless of its source. For example, water will always contain 88.81% oxygen and 11.19% hydrogen. This is true if the water come from a biological process or an industrial one.

Mixtures- a kind of matter that can be separated by physical processes. Two kinds of Mixtures:

Homogeneous- A mixture that has uniform properties throughout

Examples:

Air, sugar water, milk (the kind bought in stores)

Heterogeneous- a mixture in which the properties change from one point in the mixture to the next.

Examples:

Sewage, concrete, gravel

Physical measurements

All of chemistry, as stated before, is based on measurements of some sort. When a measurement is taken there are two parts to it. The first is the number that tells us "how many." The second is the unit that tells us what we have measured. All measurements have some sort of error in them. This error, known as **random error**, arises because the instruments we use to make the measurements are not infinitely precise. A thermometer, for instance, has marking only every degree or tenth of a degree. We could not use this thermometer to measure a temperature to a precision of a millionth of a degree. This precision in measurements is shown by the number of **significant figures** a measurement has. The random error affects the precision of a number. Usually in the lab, a measurement is taken more than once. This is to assure that the measurements are correct. However, each of the measurements may not be exactly the same. How close each of these measurements is to each other is called the **precision** of the set of measurements. Sometimes the experimenter makes a mistake in reading his or her instruments. This is called **systematic error**. Systematic errors affect the accuracy of a measurement. The **accuracy** of a measurement is how close to the actual value is the measurement.

How to determine the number of significant figures in a measurement.

Rules:

1. All non-zero (1,…,9) digits are significant.

122.3 g	4 sig figs
2.4 miles	2 sig figs
124342435245787965.43 seconds	20 sig figs

2. Zeroes between non-zero digits are significant.

201.3 lbs	4 sig figs
2001 kg	4 sig figs

3. Leading zeroes are never significant.

0.00342 L	3 sig figs
0.0000000000000000000012 kg	2 sig figs

4. Trailing zeroes are significant only if the number is in **scientific notation**.

2100 mL	2 sig figs

2.100×10^3 mL 4 sig figs

The only exception to rule for is for relatively small temperatures such as 40°C.

Scientific Notation is a method of writing large or small numbers without having to write and keep track of all of the zeroes. It is usually in the form:

$C \times 10^M$

C is the Characteristic of the number and is usually a number between 1 and 10 ($1 \geq C > 10$)

M is the mantissa and denotes how many times we multiply the characteristic by 10 and can be any integer (positive or negative). If M is greater than 0 the number is bigger than 1. If M is less than 0, the number is less than 1.

$1.2 \times 10^2 = 120$

$1.2 \times 10^{-2} = 0.012$

Significant figures in calculations.

Adding or subtracting:

The answer cannot have more precision than the numbers with which you start. The answer will have the same number of *decimal places* as the number you start with that has the least number of decimal places.

Examples:

```
  12.34
 222.3
   1.2
 235.84    (calculator answer)
 235.8     (rounded answer with the correct number of sf)


    333.5674
  12345.67
 876543.2
 889222.4374   (calculator answer)
 889222.4      (rounded answer)
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Multiplying and dividing:

The answer should have the same number of *significant figures* as the number with the fewest sig figs with which you start.

Examples:

$$22.43 \times 2.445 = 54.84135 \text{ (calculator answer)}$$
$$= 54.84 \text{ (rounded answer)}$$

$$33.456 / 0.0032 = 10455 \text{ (calculator answer)}$$
$$= 1.0 \times 10^4 \text{ (rounded answer)}$$

When doing calculations involving combinations of multiplication and addition, keep track of the significant figures at each step and round your *final* answer to the correct number of significant figures.

$$(22.34 + 1.2) \times (12.3 - 1.0)$$
$$= (23.54) \times (11.3)$$
3 sf 3 sf
$$= 266.002 \text{ (calculator answer)}$$
$$= 266 \text{ (rounded answer)}$$

Exact numbers:

Exact Numbers have an infinite number of significant figures. They come from numbers that are counted (e.g., 5 books), or defined (12 inches in 1 foot).

Units

SI units

SI units are the units used for most scientific applications. In the SI there are base unit and prefixes which modify the size of the base unit.

Base units:

Length	Meter (m)
Mass	Kilogram (kg)
Time	Second (s)
Temperature	Kelvin (K)
Amount	Mole (mol)
Electrical Current	Ampere (A)

Prefixes:

Mega (M)	10^6
Kilo (k)	10^3
Centi (c)	10^{-2}
Milli (m)	10^{-3}
Micro (μ)	10^{-6}
Nano (n)	10^{-9}
Pico (p)	10^{-12}

Additional unit: Å Angstrom = 10^{-10} m.

In chemistry metric temperature is sometimes used (°C) .

Temperature conversions:

Fahrenheit to Celsius:

$$T_F = \left(T_C \times \frac{9°F}{5°C}\right) + 32°F$$

Celsius to Fahrenheit:

$$T_C = \left(T_F - 32°F\right) \times \frac{5°C}{9°F}$$

Celsius to Kelvin:

$$T_C = \left(T_K + 273.15K\right) \times \frac{1°C}{1K}$$

Kelvin to Celsius:

$$T_K = (T_C - 273.15°C) \times \frac{1K}{1°C}$$

Derived Units

Area is a unit derived from the unit for length. Area is length $\times$ length so the SI unit for area is m^3.

Volume is a unit derived from the unit for length also. Volume is length^3 so the SI unit for volume is m^3. However 1 m^3 is a rather large volume so we use the metric unit for volume, which is the liter. Some useful conversions you should know are:

$1000 \text{ L} = 1 \text{ m}^3$
$1 \text{ mL} = 1 \text{ cm}^3$

Density is also a derived unit. It is derived from mass and volume units. The SI unit for density is kg/m^3. This unit can be a little cumbersome to use also so we will usually use units of g/mL, g/cm^3, or g/L.

Density can also be used as a conversion factor. It will convert between the mass and volume of an object.

Example:

A bar of gold has a volume of 1.00 L and a density of 19.3 g cm^{-3}. What is the **mass of the gold in pounds**?

$$1.00\text{L} \times \frac{1\text{mL}}{10^3\text{L}} \times \frac{1\text{cm}^3}{1\text{mL}} \times \frac{19.3\text{g}}{1\text{cm}^3} \times \frac{1\text{lb}}{454\text{g}} = 42.5\text{lb}$$

The above method of solving problems is called **dimensional analysis**. It is a useful method because it is self-correcting. When the units for the numbers are included in the calculation all of the units cancel out except for the units desired. If the correct units do not come out in the end, a mistake was made somewhere along the line.

Examples:

A room measures 12.4 ft by 13.2 ft by 8.1 ft. What is the **volume of the room in mL**?

Volume is length $\times$ width $\times$ height

$$(12.4\,ft \times 13.2\,ft \times 8.1\,ft) \times \left(\frac{12in}{1\,ft}\right)^3 \times \left(\frac{2.54cm}{1in}\right)^3 \times \frac{1mL}{1cm^3} = 37542701.75mL \text{ (calc answer)}$$

$$= 3.8 \times 10^7\,mL$$

(Now for a hard one) The earth has a radius of 6051.9 km. It has a density of 5.515 g cm^{-3}. What is the **mass of the earth in pounds**?

Volume of a sphere =

$$\frac{4\pi r^3}{3} = \frac{4(3.14159)(6051.9km)^3}{3} = 9.184605859\times10^{11}\,km^3 \text{ (5 sig figs)}$$

$$9.184605859\times10^{11}\,km^3 \times\left(\frac{10^3 m}{1km}\right)^3 \times\left(\frac{1cm}{10^{-2} m}\right)^3 \times\frac{5.515g}{1cm^3}\times\frac{1lb}{453.9g} = 1.116\times10^{25}\,lb$$