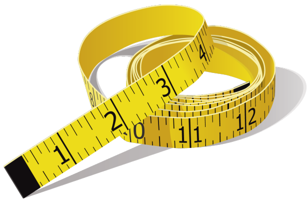
1 Measurements

MEASURING is an important part of our everyday lives, and very probably you took several measurements today. You might now be sipping a cup of coffee, or perhaps you checked the outside temperature on a street thermometer. You might be planning to bake a cake and need to use a scale and a cup to measure the flour and sugar. A cup, a thermometer, or a scale are measuring devices. It is critical to know how to accurately measure properties and, more importantly, how to transform measurements using prefixes and unit conversions. By learning how to measure and perform operations with units, you will gain experience performing basic chemistry calculations.



© PngImg

GOALS

1. Switch unit prefixes
2. Identify units and prefixes
3. Add and remove prefixed
4. exchange prefixes
5. Calculate significant figures
6. Carry out density calculations

1.1 Units of Measurements and systems of units

You probably heard the term liter, kilogram or meter. These are units of measurement. Units can be classified into different *systems of units*. For example, the unit *meter* belongs to a different system than the unit *mile*. In particular here we will address two main systems: the English System used in the US and the Metric System used by most of the industrialized world. The *Metric System* (MS) is used by scientists throughout the world and is the most common measuring system based on the meter. The *International System of Units* (SI) adopted the metric system in 1960 in order to provide additional uniformity for units used the sciences. This chapter will be mostly based on the SI units. In the following we will introduce some common units.

Discussion: (a) Discuss why is chemistry important for your career objective by listing three reasons that links chemistry with your career objective, (b) You have a glass filled with water and ice to its rim. If the ice melts will the water overflow the glass? Explain your reasoning.

`Length` What is your height? Length refers to distance and both the metric and SI unit of length is the meter (m). A smaller unit of length would be the centimeter (cm) that is commonly used in chemistry. The most important units of length are: meter, inch and mile.

`Mass` What is you weight? The mass of an object is a measure of the quantity of material it contains. You may be more familiar with the term weight than with mass. However, mass and weight are not exactly the same, as weight is a measure of the gravitational pull on an object. It differs depending on your location in the earth–in particular the height of your location. In the metric system, the unit for mass is the gram (g). The SI unit of mass, the kilogram (kg), is used for larger masses such as body weight. Pound, lb, is another unit of mass. The most important units of mass are: g, kg and lb.

`Temperature` How is the weather today? Is it cold or hot? You use a thermometer to measure temperature and for example assess how hot an object is, or how cold it is outside, or perhaps to determine if you have a fever. Temperature tells us how hot or

▼Scales measure mass

© www.wallpaperflare.com

▼Watches are used to measure time

© www.wallpaperflare.com

▼Beakers can carry a liquid volume

© wikipedia

▼Thermometers measure temperature

© PngImg

▼pipets are used in chemistry practice to add an exact volume of liquid

© www.weberscientific.com

cold an object is. Temperature can be measured in numerous units such as Celsius (°C), Fahrenheit (°F), or kelvins (K).

`Time` How long is your commute to from home work? It might take you hours to go to work, or maybe minutes. You probably think of time as years, days, minutes, or seconds. Of all these units, the International System of units (SI, abbreviated from the French Systeme international) uses seconds (s) to measure time. Still, time can be measured in s, min, or h and during this chapter we will learn how to convert units of time.

`Volume` How much milk do you usually buy? Maybe a gallon. Volume is the amount of space that a substance occupies. A liter (L) is commonly used to measure volume. The milliliter (mL) is more convenient for measuring smaller volumes of fluids in hospitals and laboratories. Gallon is still used in every-day life. L, mL and gallon are units of volume. Units of volume are in general cubic units, so for example one liter is the same as one dm^3. We will cover cubit units further in this chapter.

`Concentration` Even though we will devote a whole chapter to solutions and concentration, it felt important to introduce here the unit molarity. In chemistry, the unit molarity (M) refers to the concentration of a solution. That is the larger this number, the larger molarity, the more concentrated a solution will be. In other words, there will be more substance in the solution.

Sample Problem 1

State the type of measurement indicated in each of the following:

(a) 1ft(foot) (b) 20Kg (c) 3L (d) 300K

SOLUTION

(a) length; (b) mass; (c) volume; (d) temperature;

STUDY CHECK

State the type of measurement indicated in each of the following: (a) 800°F (b) $1m^3$ (c) 3m (d) 67s

Table ?? Different unit systems

Measuremts	Metric System	International System (SI)	English System
Length	Meter (m)	Meter (m)	Foot (ft)
Volume	Liter (L)	Cubic meter (m^3)	Gallon (gal)
Mass	Gram (g)	Kilogram (kg)	Pound (lb)
Time	Second (s)	Second (s)	Second (s)
Temperature	Celsius (°C)	Kelvin (K)	Fahrenheit (°F)

1.2 Scientific notation

Numbers in science can often be very large or tinny numbers. For example, the mass of the earth is 5900000000000000000000000000 grams, whereas the size of an atom is of the order 0.00000000001cm. Scientific notation is a standard notation extensively used among the scientific community to simplify numbers. Numbers not expressed in scientific notation are referred to as numbers in full notation (e.g. 0.234g). Numbers expressed in scientific notation place the decimal after the first significant digit while being accompanied by a power of ten (e.g. 2.34×10^{-1}g). In order to express a number (e.g. 345000g) in scientific notation, we just need to leave the first digit followed by the decimal point which was moves in this case to the left 5 places, giving 3.45×10^5g. The power of ten can contain a positive or a negative exponent. Positive exponents represent large numbers, whereas negative exponents represent numbers smaller than 1. Below are two examples of full and scientific notation. On one hand a number smaller than one:

$$0.000134\text{g} \quad \text{(full notation)} \qquad 1.34 \times 10^{-4}\text{g} \;\; \text{(scientific notation)}$$

On one hand a number smaller than one:

$$4500000\text{g} \quad \text{(full notation)} \qquad 4.5 \times 10^{6}\text{g} \;\; \text{(scientific notation)}$$

1.3 Prefixes & Conversion Factors

Let's consider the following measurements: 1 **k**m, 2 **c**m, and 3 m that can be read as one kilometer, two centimeters, and three meters. The word kilo (**k**) and centi (**c**) are called prefixed whereas meter (m) is a simple unit. Kilometer is larger than meter, whereas centimeter is smaller than a meter. Prefixes such as kilo or centi are attached to units in order to make numbers more manageable. For example, the radius of the earth is 6356 km, and this number is easier to handle than 6356000m. At the same time, we can attach any prefix to different units. Hence, we can talk about a centimeter (**c**m) but also about a centisecond (**c**s) or centiliter (**c**L). All these units have the same prefix. Table **??** lists some of the metric prefixes, their symbols, and their decimal values.

Table ?? Different prefixes

Prefix	Symbol	Meaning	Value
peta	P	1000000000000000	1×10^{15}
tera	T	1000000000000	1×10^{12}
giga	G	1000000000	1×10^{9}
mega	M	1000000	1×10^{6}
kilo	k	1000	1×10^{3}
hecto	h	100	1×10^{2}
deca	da	10	1×10^{1}
–	–	1	1×10^{0}
deci	d	0.1	1×10^{-1}
centi	c	0.01	1×10^{-2}
milli	m	0.001	1×10^{-3}
micro	μ	0.00001	1×10^{-6}
nano	n	0.000000001	1×10^{-9}
pico	p	0.000000000001	1×10^{-12}
femto	f	0.000000000000001	1×10^{-15}

`How to identify prefixes?` Look for example at the measurement 2 cm. Centi (c) is the prefix and means 1×10^{-2} and meter (m) is the unit which refers to length. Another example, 7 kg means kilogram. Kilo (k) is the prefix and means 1×10^{3}, whereas gram (g) is the unit that refers to mass. The prefix refers to the first letter whereas the unit refers to the last letter.

Would you prefer to be paid a kilodollar, a dollar or a centidollar? A unit with a prefix can be bigger or smaller than the plain unit–this is the unit without prefix–, depending on the prefix. The following prefixes make the unit smaller: deci, centi, milli, micro, nano, pico and femto. For example a fs (femtosecond) is smaller than a s (second). Differently, the following prefixes make the unit larger: Tera, Giga, Mega. For example a Tb (terabyte) is larger than a b (byte). byte is a unit used in computer science.

How to write unit equalities and conversion factors Unit equalities are simple expressions that relate a unit with a unit with prefix. For example: one centimeter (cm) is 1×10^{-2}m. Hence we can write this as a unit equality:

$$\boxed{1cm = 1 \times 10^{-2}m} \text{ unit equality}$$

Let's compare cm and m. The first, cm, is a unit with a prefix, whereas m is simple a unit of length without a prefix. In order know how many m are there in a cm we need to write down a conversion factor. Think about prefixes as synonymous of a number. In this way, centi stands for 1×10^{-2}, so

$$\boxed{\frac{1cm}{1 \times 10^{-2}m} \text{ or } \frac{1 \times 10^{-2}m}{1cm}} \text{ conversion factor}$$

Sample Problem 2

Complete each of the following equalities and conversion factors:

(a) 1dm=______m

(b) 1km=______m

(c) $\dfrac{1nm}{\_\_\_\_\_\_m}$

(d) $\dfrac{\_\_\_\_\_\_m}{1cm}$

SOLUTION

(a) 1dm=$1 \times 10^{-1}m$; (b) 1km=$1 \times 10^{3}m$; (c) $\dfrac{1nm}{1 \times 10^{-9}m}$; (d) $\dfrac{1 \times 10^{-2}m}{1cm}$;

STUDY CHECK

Second is a unit of time. Complete each of the following equalities and conversion factors involving seconds:

(a) 1cs=______s (b) $\dfrac{\_\_\_\_\_\_s}{1Ts}$ (c) $\dfrac{\_\_\_\_\_\_s}{1Ms}$

1.4 Using Conversion Factors

Unit equalities in the form of conversion factors are used to convert a unit into another. Sometimes one wants to get rid of a prefix, such as when we transform centimeter (cm) into meter (m). Sometimes, one wants to convert a prefix into another prefix. An example would be converting centimeters (cm) to millimeters (mm). Let's work on some example.

Removing or adding prefixes Imagine that you need to remove a prefix from a unit, and convert 3 km (we will call this one the original unit) in meters (this is the final unit). First, you would need the conversion factor corresponding to the prefix (centi) from Table **??**. Then you need to arrange the conversion factor placing the prefix at the bottom of the fraction. This will cancel out the prefix in the original unit and in the bottom part of the conversion factor, hence leaving the final unit on top of the conversion factor. The arrangement would be:

$$3\cancel{km} \times \frac{1 \times 10^3 m}{1\cancel{km}} = 3000m$$

Imagine now that you need to add a prefix into a unit, and convert 4000 m in km. The same would apply for this case, but now you will have to arrange the conversion factor so that the prefix is on the top:

$$4000\cancel{m} \times \frac{1\text{ km}}{1 \times 10^3\cancel{m}} = 4km$$

Sample Problem 3

The length of a textbook page is 20cm. Convert 20cm to meters, expressing the result in scientific notation.

SOLUTION

In order to convert 20cm into meters, we need to remove the prefix (centi) leaving the unit (meter) without any prefix. We will use the conversion factor that relates m to cm: $\frac{1 \times 10^{-2} m}{1cm}$ or $\frac{1cm}{1 \times 10^{-2} m}$. We will arrange the conversion factor so that cm cancels giving m and hence we will use $\frac{1 \times 10^{-2} m}{1cm}$:

$$20\cancel{cm} \times \frac{1 \times 10^{-2} m}{1\cancel{cm}} = 2 \times 10^{-1} m$$

The original units and on the bottom of the conversion factor cancel and we get meters, the final unit.

STUDY CHECK

Convert 100m to km, expressing the result in scientific notation.

Switching prefixes In order to switch a prefix into another prefix, such as transforming 30 millimeters (30 mm) into centimeters (cm), you will need two different conversion factors: the first conversion factor will remove the original unit (mm) introducing an intermediate unit, meters (m), whereas the second conversion factor will remove the intermediate meter and introduce the final unit (cm). You will get the conversion factors from Table **??**. You will arrange the first conversion factor so that the original unit cancels out with the bottom of the first conversion factor, giving you an intermediate unit. You will arrange the second conversion factor so that the intermediate unit cancels out with the bottom of the second conversion factor giving the final unit. For this example:

$$30\cancel{mm} \times \frac{1 \times 10^{-3}\cancel{m}}{1\cancel{mm}} \times \frac{1cm}{1 \times 10^{-2}\cancel{m}} = 3cm$$

Sample Problem 4

The length of a textbook page is 20cm. How many mm correspond this length, expressing the result in scientific notation.

SOLUTION

We want to convert 20 cm into mm, that is, we are switching prefixed. In order to do this, you need two conversion factors: $\dfrac{1\times10^{-2}m}{1cm}$ and $\dfrac{1\times10^{-3}m}{1mm}$. You will have to arrange the number (20cm) and the two conversion factors in the following form:

$$20\cancel{cm}\times\frac{1\times10^{-2}\cancel{m}}{1\cancel{cm}}\times\frac{1mm}{1\times10^{-3}\cancel{m}}=2\times10^{3}mm$$

STUDY CHECK

Convert 100mm to km, expressing the result in scientific notation.

`Square or cubic units` How big is your apartment? You might be living in a $750ft^2$ loft in Brooklyn or in a larger house Upstate. Often times we encounter cubic or square units such as cubic centimeter (cm^3) or square feet (ft^2). The equivalencies for cubic or square units should take into account the unit power (power of two or power of three). If $1cm = 1\times10^{-2}m$, for square units the relation should be squared and $1cm^2 = 1\times(10^{-2})^2m^2 = 1\times10^{-4}m^2$. Another example, for the case of mm and mm^3:

$$\boxed{\frac{1mm}{1\times10^{-3}m}\quad\text{and}\quad\frac{1mm^3}{1\times10^{-9}m^3}}$$

Let us work on an example in which we want to convert $30m^2$ into m^2:

$$30\cancel{m^2}\times\frac{1cm^2}{1\times10^{-4}\cancel{m^2}}=3\times10^{5}cm^2$$

90

Sample Problem 5

How many m^2 is $20cm^2$, expressing the result in scientific notation.

SOLUTION

In order to convert $20cm^2$ to square meters, we need to remove the centi prefix and that will give us the unit square meter without any prefix. We will use the conversion factor that relates m^2 to cm^2: $\dfrac{1\times10^{-4}m^2}{1cm^2}$ or $\dfrac{1cm^2}{1\times10^{-4}m^2}$.

$$20\cancel{cm^2}\times\frac{1\times10^{-4}m^2}{1\cancel{cm^2}}=2\times10^{-3}m^2$$

STUDY CHECK

Convert $100m^3$ to dm^3, expressing the result in scientific notation.

`Units of volume` Units such as L or mL are units of volume. As volume is a three-dimensional property, those units somehow have to be related to the units of length. In fact one liter is the same as one dm^3 and one ml is the same as one cm^3. In the allied health field, the units *mL* is also written as *cc* as in cubic centiliters.

$$\boxed{1L = 1dm^3\ \text{ and }\ 1mL = 1cm^3}$$

Let us work on an example in which we want to convert $30cm^3$ into L:

$$30\cancel{cm^3}\times\frac{1\cancel{mL}}{1\cancel{cm^3}}\times\frac{1\times10^{-3}L}{1\cancel{mL}}=3\times10^{-2}L$$

Sample Problem 6

Convert 30 m^3 into L, expressing the result in scientific notation.

SOLUTION

In order to convert m^3 into L we just need to remember that the L actually refers to dm^3, therefore is connected to meter. We will first convert m^3 into dm^3 and then dm^3 into L.

$$30\cancel{m^3} \times \frac{1\cancel{dm^3}}{1 \times 10^{-3}\cancel{m^3}} \times \frac{1L}{1\cancel{dm^3}} = 3 \times 10^4 L$$

STUDY CHECK

Convert 40L to cm^3, expressing the result in scientific notation.

Table ?? Table containing some common unit equalities

Unit	Equality
Inches (in)-centimeters (cm)	2.54 cm = 1 in
miles (mi)-meters (m)	1 mi = 1609.34m
minutes (min)-hours (h)	60 min = 1 h
minutes (min)-seconds (s)	60 s = 1 min
pound (lb)-grams (g)	454 g = 1 lb
cubic centimeter (cm^3)-mililiters (mL)	1 mL = 1cm^3
Liter (L)-cubic decimeters (dm^3)	1 L = 1dm^3
drops-mililiters* (mL)	1 mL = 15 drops

* There are several definitions of a drop

Using other equalities How many hours is 300 minutes, or how many centimeters is 2 inches? Some of the units conversion are not based on a power of ten relationship and do not contain prefixes such as kilo or centi. Table **??** lists some of the common equalities that can be easily converted into conversion factor. As an example, the unit equivalency between hours and minutes is $60min = 1h$ and the conversion factor would be $\frac{60min}{1h}$ or $\frac{1h}{60min}$.

Sample Problem 7

Convert 20 in to cm, expressing the result in scientific notation.

SOLUTION

We want to convert 20 inches into centimeters. The relationship between Inch and centimeter is given in Table **??**. In order to do this, you need the conversion factor: $\frac{1in}{2.54cm}$ or $\frac{2.54cm}{1in}$. You will have to arrange the number (20 in) and the conversion factor in the following form:

$$20\cancel{in} \times \frac{2.54cm}{1\cancel{in}} = 5.080 \times 10^1 cm$$

STUDY CHECK

Convert 200mL to drops, expressing the result in scientific notation.

1.5 Significant Figures

Exact numbers results from counting. For example, think about how many eggs are there in your refrigerator, there might be three and this number is an exact number. Differently, numbers that results from a measurement are called measured values and they are subject to uncertainty–in another words error. For example, if you weight a single egg in an scale depending of the type of scale you used you will measure 70g or 71g or maybe 70.8g. The mass of an egg is a measured property and hence some of the digits of the measurement are uncertain. The goal of this section is, given a value, calculate the number of significant figures of a number (we will refer to significant figures as SF, or SFs). Another goal is to estimate significant figures in calculation in order to express the result with the right number of digits and significant figures.

`Significant figures of numbers` In general all numbers different than zero are significant and for example the number 123 has three significant figures. Similarly, the number 45 has two significant figures. Zeros are also significant except when:

Exception 1 *A zero is not significant when placed at the beginning of a decimal number.*

For example, the number 0.123 has three significant figures, as the first zero is not significant. Similarly, the number 0.002340 has four significant figures as the first three zeros are not significant but the last zero it is. Mind the rule affects only the zeros at the beginning. A final example:

$$0.032 \;\; \text{(2SF)}$$

Exception 2 *A zero is not significant when used as a placeholder in a number without a decimal point.*

For example, the number 1000 has only one significant figure, and the number 3400 has two. Let us consider more examples. The number 120 has two significant figures, as according to the second rule the last zero is not significant. Differently, the number 1203 has four significant figures, as the zero in between two numbers is not affected by neither the first nor the second rule. A final example,

$$3200 \;\; \text{(2SF)}$$

Exception 3 *A zero in a number expressed on scientific notation is significant*

For example the zero in 3.0×10^{-2} is significant, and the number has 2SFs. A final example:

$$3.2020 \times 10^{2} \;\; \text{(5SF)}$$

Sample Problem 8

Indicate the number of significant figures in the following numbers: 123, 4567, 1200, 340, 0.001, 0.023 and 0.0405.

SOLUTION

123 has three significant figures, whereas 4567 has four SF. 1200 has only 2SF as the last two zeros are not significant, and 340 has only 2SF as the last zero is not significant. 0.001 has only one significant figure as the first 3 zeros are not significant and 0.023 has only two SFs. Finally, 0.0405 has threee SFs as the first two zeros are not significant but the zero between 4 and 5 is indeed significant.

STUDY CHECK

Indicate the number of significant figures (SFs) in the following numbers: 4560, 0.123, 1000 and 0.0030.

`Significant figures in calculations` There are two different rules that allow you to express the result of a calculations with the correct number of figures.

Rule 1 $(+\ -)$ *For additions or subtractions, the results has the same number of decimal places as the number with the least decimal places in the calculation.* For example:

$$34.3451 + 34.5 = 68.8 \ (+\ -\text{less decimals})$$

If you add $34.3451 + 34.5$ you will obtain 68.8451, however, as 34.3451 has four decimal places (4DP) and 34.5 has one decimal place (1DP), the result of adding both numbers will have to have only one decimal place, therefore 68.8451 needs to be rounded to 68.8 (1DP). Overall, we have:

$$34.3451\ (4\text{DP})\ + 34.5\ (1\text{DP})\ = 68.8\ (1\text{DP})$$

Rule 2 $(\times\ \div)$ *For multiplications and divisions, the number of significant figures of the result should be the same as the least number of significant figures involved.* For example, if you carry the following multiplication:

$$4500 \times 342 = 1500000 \ (\times\ \div \text{less SFs})$$

the number 4500 (2SF) has two significant figures, whereas the number 342 (3SF) has three significant figures. If we multiply both numbers the results should contain just two significant figures. The result of multiplying 4500×342 is 1539000 (4SF), however, this number needs to be rounded into two significant figures into 1500000 (2SF). Overall we have:

$$4500\ (2\text{SF})\ \times 342\ (3\text{SF})\ = 1500000\ (2\text{SF})$$

Sample Problem 9

Do the following calculation with the correct number of figures.

$$\frac{88.5 - 87.57}{345.13 \times 100}$$

SOLUTION

We will analyze each number indicating the number of SF and Digits (DP): 88.5(3SF, 1DP), 87.57(4SF, 2DP), 345.13(6SF, 2DP) and 100(1SF, 0DP). The result of doing the addition needs to be rounded to one single decimal place: $88.5 - 87.57 = 0.93 \simeq 0.9$. After that we have only multiplications and divisions and hence we will now focus on the number of SFs:

$$\frac{0.9\ (1\text{SF})}{345.13\ (5\text{SF})\ \times 100\ (1\text{SF})}$$

The result of this operation needs to be rounded to one SF:

$$\frac{0.9}{345.13 \times 100} = 2.6077 \times 10^{-5} \simeq 3 \times 10^{5}\ (1\text{SF})$$

STUDY CHECK

Do the following calculation with the correct number of figures: $(24.56 + 2.433) \times 0.013$

1.6 Matter

Matter–the material of the universe–represents anything with mass that occupies space. It has different levels of organization and complexity. We can classify matter in terms of composition. Some substances are made of a single component whereas others contain multiple components. At the same time, some substances are made of many components while they appear to be made of a single component.

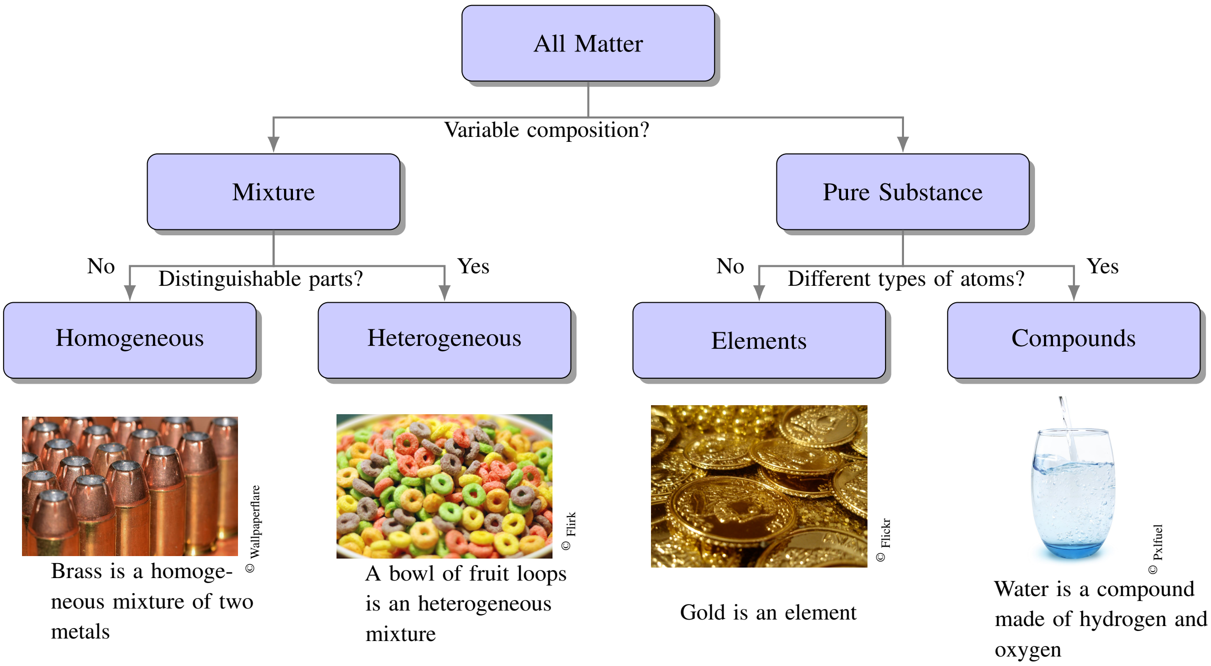


Figure ?? Classification of the matter

`Solids, liquids & gases` *Solid* has a well-defined shape and volume. Think about an ice cube, for example, that is made of water in the form of ice. In an ice cube, attractive forces keep the shape of the cube constant. The water molecules in the ice are arranged in rigid patterns and they can only vibrate in the solid. A *liquid*, on the other hand, has a well-defined volume. However, liquids do not have a constant shape, as their shape will depend on the shape of the container. Think about water that is a liquid. You can find many different bottle shapes. In all of them, the molecules of water will arrange to occupy the shape of the container. The volume of the liquid–the amount of space the liquid occupies–will be constant but not the shape. In a liquid, the particles move randomly but are still attached. A *gas* does not have a well-defined shape or volume. In a gas, the particles are randomly distributed and barely interact with each other, as they move at high speeds, taking the shape and volume of their container. The process of boiling and freezing represent *physical changes* and during physical changes, these substances involved do not change their composition. When water boils, both steam and water are made of the same component, H_2O. When a substance undergoes a *chemical process*, it will change its composition. For example, when burning a piece of paper, paper made of carbon, hydrogen, and hydrogen becomes ash made mostly of carbon. As such, during a chemical process, a substance (e.g. paper) becomes a new substance (e.g. carbon).

`Pure Substances and mixtures` On one hand, *pure substances* have definite composition, being only made of a single component. For example, water and gold are pure substances. However, there are two different types of pure substances: elements and compounds. *Elements* are composed of only one type of atom. Examples are silver, iron, and aluminum. They all contain one type of substance, and for example, iron is only made of iron atoms. *Compounds* are combinations of different elements. For example, water, H_2O is a compound made of a combination of hydrogen and oxygen atoms. On the other hand, *mixtures* are physical combinations of different pure substances. Mixtures have variable composition. For example, air is a mixture of oxygen and nitrogen. Wood, soda pop, or soil are all mixtures. Mixtures can be homogeneous or heterogeneous. In a *homogeneous mixture*–also known as solutions–the composition is uniform throughout the sample. An example of a homogeneous mixture is salty water, a solution of salt and water. *Heterogeneous mixtures* are mixtures in which the components are not uniformly distributed throughout the sample. An example would be a chocolate chip cookie in which you can differentiate the dough and the chocolate.

Sample Problem 10

Classify as element, compound, homogeneous mixture, heterogeneous mixture:
(a) An iron nail (b) Milk (c) Sugar (d) miso soup

SOLUTION

(a) An iron nail is an element as it is only made of iron, a single material; (b) Milk is a homogeneous mixture as it is made of water, fat, protein even though you only see a single substance; (c) Sugar is a compound made of carbon and other constituents; (d) miso soup is a mixture of water, fat and other chemicals and therefore is a mixture. As you can differentiate its constituents we call this heterogeneous mixture.

STUDY CHECK

Classify as element, compound, homogeneous mixture, heterogeneous mixture:
(a) muscle milk (b) water (c) a gold ring (d) rice & beans

Separation Techniques Chemists and scientists often need to separate the components of a mixture. There are many analytical techniques used to separate and identify the components of a mixture. These techniques exploit the differences in the chemical or physical properties of the components of the mixture to separate the different elements. The techniques used to separate mixtures are called *separation techniques*. Distillation is an important separation technique used to separate compounds with different volatility–different tendencies to generate vapor. For example, a mixture containing water and gasoline can be distilled as the components of the mixture largely differ in their volatility. The distillation setup consists of a distilling flask containing the mixture to be distilled, a condenser, and a receiving flask. When the mixture is heated the most volatile component will separate first and with the help of the condenser–a cooled tube–will cool down into liquid and deposit on the receiving flask. It is important to understand that only mixtures of liquids of mixtures of solids and liquids can be distilled. Another important separation technique is *filtration*. This technique is used to separate mixtures of solids and liquids. The mixture to be filtered is poured into a mesh–for example a filter–that retains the solid and lets the liquid pass. *Chromatography* is another separation technique. Indeed, the name represents a series of methods that uses two different phases, a mobile phase, and a stationary phase, to separate the components of a mixture. The principle behind this technique is the differences in affinity of the mixture components towards the mobile and stationary phase. The stationary phase is a solid whereas the mobile phase can either be a liquid (paper chromatography) or a gas (gas chromatography). Mixture components with a high affinity for the mobile phase will move faster through the chromatographic system. *Thin layer chromatography* (TLC) relies on capillarity, which is the tendency of liquid substances to rise on the surface of a material. In this technique, a drop of a liquid solution containing different substances (the sample) is deposited on a rectangular piece of filter paper, close to the bottom edge. This paper is called the stationary phase. The bottom end of the paper is immersed in a liquid called the mobile phase, to a point that is just below the spot where the sample was placed. Due to capillarity, the mobile phase will move up along the stationary phase. When the mobile phase reaches the sample, the different components of the mixture will begin to migrate, carried away by the mobile phase. The chemical compounds forming the sample will move with the mobile phase, but as different chemicals have different tendencies to stick to the mobile phase, they will cover different distances in the stationary phase. The different heights achieved by the different substances would allow you to identify those chemicals. A component of the mixture with a high affinity to the mobile phase will migrate more than a component with a higher affinity to the stationary phase. The distance traveled by a component is referred to as the distance traveled by the mobile phase is a measure of that affinity between the chemical and the mobile phase. The *retention factor* R_f of a given chemical as is defined as:

$$R_f = \frac{\text{distance traveled by the chemical}}{\text{total distance traveled by mobile phase}}$$

The R_f value of a substance is characteristic of that substance. When dealing with mixtures one has to calculate the R_f for each pure component separately to then compare the retention factors with the ones obtained in the mixture.

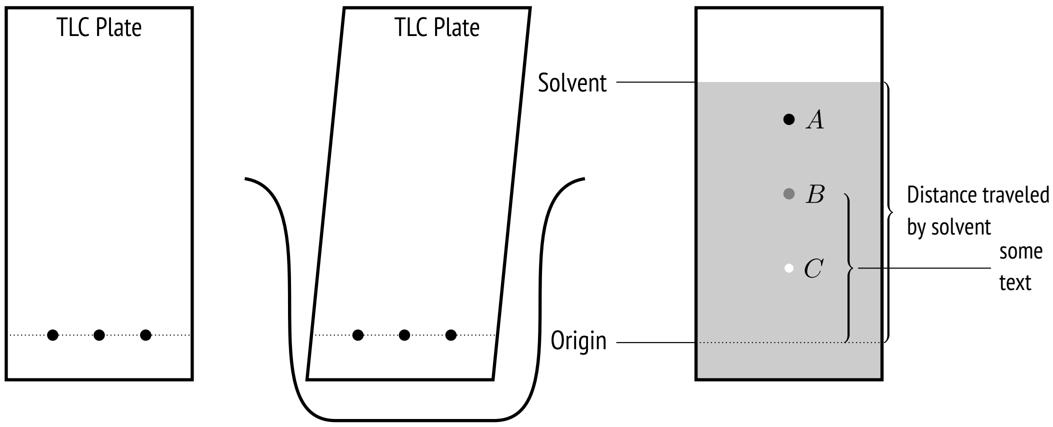


Figure ?? TLC set up. (Left) Solid phase with spots (Center) Solid phase inside a beaker containing the liquid phase (Right) Solid phase with separated mixture components.

1.7 Density

`Density` Density refers to the mass of a substance with respect to its volume. This is an unique property for each substance. For example, the density for copper is 8.92 $g \cdot ml^{-1}$ and for gold is 19.3 $g \cdot ml^{-1}$. By measuring density only you would be able to differentiate copper than gold. The larger density the more compact is an object and that means the more mass per volume it has. The formula for density is

$$\boxed{\text{Density} = \frac{\text{Mass of substance}}{\text{Volume of substance}}} \tag{1.1}$$

Density and mixing A small piece of ice will float on water. The reason for that is density: density of ice (0.9g/mL) is smaller than density of water (1.0g/mL) and hence ice will stay on top of water. Objects with density larger than 1 g/mL will sink whereas objects with density smaller than this value will float. If you add a drop of vegetable oil to a glass of water, the drop will float. This is because the density of oil is smaller than 1g/mL.

Table ?? Density of some common substances at 273.15 K and 100 kPa

Substance	Density (g/mL)	Physical State
Helium	0.2	gas
Hydrogen	0.1	gas
Water	1.0	Liquid
Cooking oil	0.9	Liquid
Mercury	13.5	Liquid
Tetrachloroethene	1.6	Liquid
Gold	19.3	solid
Plastics	1.2	solid
Ice	0.916*	solid

*Ice is given at T <273.15 K

Sample Problem 11

In the figure, we mixed three liquids of density: A (0.5 g/mL), B(2 g/mL) and C(1 g/mL). Identify each liquid.

SOLUTION

The heavier the liquid, that is the larger density, the lower the liquid will arrange in the mixture. From top to bottom we have A, C and B.

STUDY CHECK

Indicate the order that the following invisible liquids will appear in a cylinder when mixed: (a) benzene(0.87 g/mL) and (b) water(1 g/mL)

Density and the volume of objects Density depends on volume and in particular the larger volume the smaller density. In the Figure below you can find formulas to calculate the volume of some common objects like an sphere or a cube

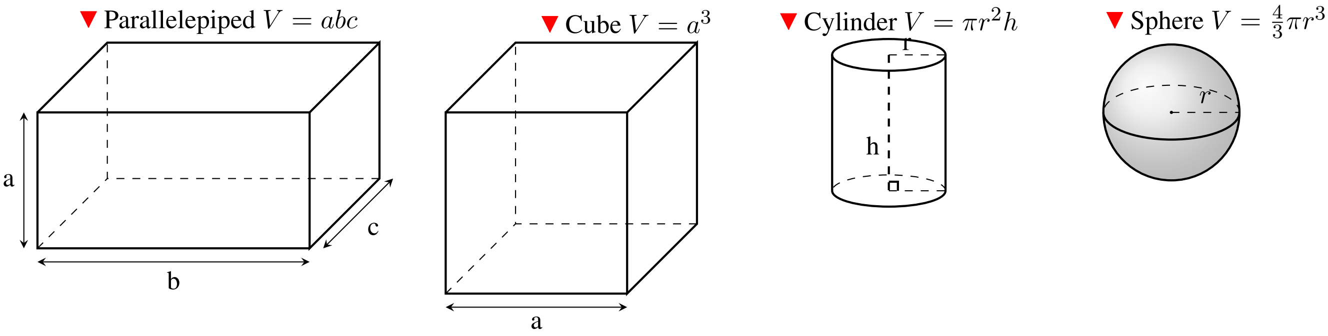


Figure ?? Volume of some objects

Sample Problem 12

After adding a 30g object into a cylinder filled of water, the level of water rises from 60mL to 90mL. Calculate the density of the object.

SOLUTION

Density is mass over volume. The mass of the object is 30g and its volume is (90-60)mL that is 30mL. Hence: $d = 30g/30mL = 1g/mL$.

STUDY CHECK

A lead weight used in the belt of a scuba diver has a mass of 226 g. When the weight is placed in a graduated cylinder containing 200.0 mL of water, the water level rises to 220.0 mL. What is the density of the lead weight (g/mL)?

240