UNIT 01: Matter & Measurement

What is chemistry?

Chemistry can be described as the science that deals with matter, and the changes that matter undergoes. It is sometimes called the central *science* because so many naturally occurring phenomena involve chemistry and chemical change.

Scientific problem solving

Scientific (logical) problem solving involves three steps;

1. State the problem and make observations. Observations can be *quantitative* (those involving numbers or measurement) or *qualitative* (those not involving numbers).
2. Formulate a possible explanation (this is known as a *hypothesis*).
3. Perform experiments to test the hypothesis. The results and observations from these experiments lead to the modification of the hypothesis and therefore further experiments.

Eventually, after several experiments, the hypothesis may graduate to become a *theory*. A theory gives a universally accepted explanation of the problem. Of course, theories should be constantly challenged and may be refined as and when new data and new scientific evidence comes to light.

Theories are different to *laws*. Laws state what general behavior is observed to occur naturally. For example, the *law of conservation of mass* exists since it has been consistently observed that during all chemical changes mass remains unchanged (i.e., it is neither created nor destroyed).

Measurements

Measurements, and subsequently calculations applied to those measurements, allow the determination of some of the quantitative properties of a substance; for example, mass and density.

Scientific notation

Measurements and calculations in chemistry often require the use of very large or very small numbers. In order to make handling them easier, such numbers can be expressed using *scientific* *notation*. All numbers expressed in this manner are represented by a number between 1 and 10 which is then multiplied by 10, raised to a particular power.

The number of places the decimal point has moved determines the power of 10. If the decimal point has moved to the left then the power is positive, if it has moved to the right then it is negative.

For example, the number 42000.0 is converted to scientific notation by using the number 4.2. In the process the decimal point has moved four places to the *left*, so the power of 10 used is +4.

$$42000.0 = 4.2 \times 10^4$$

The number 0.00012 is converted to scientific notation by using the number 1.2. In the process the decimal point has moved four places to the *right*, so the power of 10 used is -4.

$$0.00012 = 1.2 \times 10^{-4}$$

Task 01a

1 Convert the following numbers to scientific notation.

(a) 24500
(b) 356
(c) 0.000985
(d) 0.222
(e) 12200

2. Convert the following scientific notation numbers to non-scientific notation numbers.

(a) 4.2×10^{3}
(b) 2.15×10^{-4}
(c) 3.14×10^{-6}
(d) 9.22×10^{5}
(e) 9.57×10^{2}

SI units

Units tell us the scale that is being used for measurement. Prefixes are used to make writing very large or small numbers easier. Common SI (*System International*) units and prefixes are given below.

Base quantity	Name of unit	Symbol
Mass	Kilogram	kg
Length	Meter	m
Time	Second	s
Amount of substance	Mole	mol
Temperature	Kelvin	K

Prefix	Symbol	Meaning
Giga	G	10^{9}
Mega	M	10^{6}
Kilo	k	10^{3}
Deci	d	10^{-1}
Centi	c	10^{-2}
Milli	m	10^{-3}
Micro	μ	10^{-6}
Nano	n	10^{-9}
Pico	p	10^{-12}

Converting units and dimensional analysis (the factor label method)

One unit can be converted to another unit by using a conversion factor. Application of the simple formula below will allow the conversion of one unit to another. This method of converting between units is called *dimensional analysis* or the *factor-label method*.

(unit a) (conversion factor) = unit b

The conversion factor is derived from the equivalence statement of the two units. For example, in the equivalence of 1.00 inch = 2.54 cm, the conversion factor will either be,

$$\frac{2.54\text{ cm}}{1.00\text{ inch}} \quad \textit{or} \quad \frac{1.00\text{ inch}}{2.54\text{ cm}}$$

The correct choice is the one that allows the cancellation of the unwanted units. For example, to convert 9.00 inches to cm, perform the following calculation

$$\frac{9.00\ \cancel{\text{inch}}}{} \text{ x } \left(\frac{2.54\text{ cm}}{1.00\ \cancel{\text{inch}}}\right) = 22.86\text{ cm}$$

To convert 5.00 cm into inches, perform the following calculation

$$\frac{5.00\ \cancel{\text{cm}}}{} \text{ x } \left(\frac{1.00\text{ inch}}{2.54\ \cancel{\text{cm}}}\right) = 1.97\text{ inches}$$

Task 01b

1. ***Convert the following quantities from one unit to another, using the following equivalence statements; 1.000 m = 1.094 yd, 1.000 mile = 1760 yd, 1.000 kg = 2.205 lbs***

   (a) 30 m to miles
   (b) 1500 yd to miles
   (c) 206 miles to m
   (d) 34 kg to lbs
   (e) 34 lb to kg

2. ***In each case below, which is the larger quantity?***

   (a) A distance of 3.00 miles or 3000. m.
   (b) A mass of 10.0 kg or 25 lbs.

Temperature

There are three scales of temperature that you may come across in your study of chemistry. They are Celsius (^oC), Fahrenheit (^oF) and Kelvin (K). The following conversion factors will be useful.

Temperature Conversion factors	
Celsius to Kelvin	$\boxed{\text{T in K} = \text{T in } ^oC + 273}$
Kelvin to Celsius	$\boxed{\text{T in } ^oC = \text{T in K} - 273}$
Celsius to Fahrenheit	$\boxed{\text{T in } ^oF = (1.8\ (\text{T in } ^oC)) + 32}$
Fahrenheit to Celsius	$\boxed{\text{T in } ^oC = \dfrac{(\text{T in } ^oF - 32)}{1.8}}$

Task 01c

1. ***Convert the following temperatures from one unit to the other.***

   (a) ***263 K to ^oF***
   (b) ***38 K to ^oF***
   (c) ***13 ^oF to ^oC***
   (d) ***1390 ^oC to K***
   (e) ***3000 ^oC to ^oF***

2. ***When discussing a <u>change</u> in temperature, why will it not matter if the change is recorded in Celsius or Kelvin?***

Derived units

All other units can be derived from base quantities. One such unit that is very important in chemistry is volume. Volume has the unit, length$^{3.}$ Common units for volume are liters (L) or milliliters (mL).

$$1.000 \text{ mL} = 1.000 \text{ cm}^3$$

and

$$1.000 \text{ L} = 1000. \text{ mL} = 1000. \text{ cm}^3 = 1.000 \text{ dm}^3$$

Density is the ratio of the mass to volume.

$$\boxed{\text{density} = \frac{\text{mass}}{\text{volume}}}$$

This relationship is particularly useful when dealing with liquids in chemistry. Liquids are most conveniently measured by pouring them into, say, a graduated cylinder. The graduated cylinder records a volume, not a mass. In order to calculate the mass of a known volume of a liquid (assuming the density is known) the relationship below can be applied.

$$\boxed{\text{mass} = (\text{density})\ (\text{volume})}$$

Assuming that density has the units of g/L, volume has units of L, and by using dimensional analysis, it can be seen that the resultant unit for mass in this case is g.

$$\left(\frac{g}{\cancel{L}}\right)\frac{(\cancel{L})}{} = g$$

Uncertainty, significant figures and rounding

When reading the scale on a piece of laboratory equipment such as a graduated cylinder or a buret, there is always a degree of uncertainty in the recorded measurement. The reading will often fall between two divisions on the scale and an estimate must be made in order to record the final digit. This estimated final digit is said to be *uncertain* and is reflected in the recording of the numbers by using +/-. All of the digits that can be recorded with certainty are said to be *certain*. The certain and the uncertain numbers taken together are called *significant figures*.

Determining the number of significant figures present in a number

1. Any non-zero integers are always counted as significant figures.
2. Leading zeros are those that precede all of the non-zero digits and are never counted as significant figures.
3. Captive zeros are those that fall between non-zero digits and are always counted as significant figures.
4. Trailing zeros are those at the end of a number and are only significant if the number is written with a decimal point.
5. Exact numbers have an unlimited number of significant figures. (Exact numbers are those which are as a result of counting e.g., 3 apples or by definition e.g., 1.000 kg = 2.205 lb).
6. In scientific notation the 10^x part of the number is never counted as significant.

Determining the correct number of significant figures to be shown as the result of a calculation

1. When multiplying or dividing. Limit the answer to the same number of *significant figures* that appear *in the original data with the fewest number of significant figures*.
2. When adding or subtracting. Limit the answer to the same number of *decimal places* that appear *in the original data with the fewest number of decimal places*.

i.e., don't record a greater degree of significant figures or decimal places in the calculated answer than the weakest data will allow.

Rounding

Calculators will often present answers to calculations with many more figures than the significant ones. As a result many of the figures shown are meaningless, and the answer, before it is presented, needs to be rounded.
In a multi-step calculation it is possible to leave the rounding until the end i.e., leave all numbers on the calculator in the intermediate steps, or round to the correct number of figures in each step, or round to an extra figure in each intermediate step and then round to the correct number of significant figures at the end of the calculation. In most cases in the AP chemistry course you will leave numbers on the calculator and round at the end.

Whichever method is being employed, use the simple rule that if the digit directly to the right of the final significant figure is less that 5 then the preceding digit stays the same, if it is equal to or greater than 5 then the preceding digit should be increased by one.

Task 01d

1. ***Determine the number of significant figures in the following numbers.***

   (a) *250.7*
   (b) *0.00077*
   (c) *1024*
   (d) *4.7 x* 10^{-5}
   (e) *34000000*
   (f) *1003.*

2. ***Use a calculator to carry out the following calculations and record the answer to the correct number of significant figures.***

   (a) *(34.5) (23.46)*
   (b) *123 / 3*
   (c) *(2.61 x* 10^{-1}*) (356)*
   (d) *21.78 + 45.86*
   (e) *23.888897 - 11.2*
   (f) *6 - 3.0*

Accuracy and precision

Accuracy relates to how close the measured value is to the actual value of the quantity. *Precision* refers to how close two or more measurements of the same quantity are to one another.

Task 01e

1. ***Consider three sets of data that have been recorded after measuring a piece of wood that is exactly 6.000 m long.***

	SET X	***SET Y***	***SET Z***
	5.864 m	***6.002 m***	***5.872 m***
	5.878 m	***6.004 m***	***5.868 m***
Average Length	***5.871 m***	***6.003 m***	***5.870 m***

(a) Which set of data is the most accurate?
(b) Which set of data is the most precise?

Percentage error

The data that are derived in experiments will often differ from the accepted, published, actual value. When this occurs, a common way of expressing accuracy is *percentage error*.

$$\text{Percentage Error} = \left|\frac{(\text{Actual Value - Calculated Value})}{\text{Actual Value}}\right| \times 100$$

States of matter and particle representations

All matter has two distinct characteristics. It has mass and it occupies space. Properties associated with the three states of matter, and the behaviors of the particles that make up each, are summarized below.

SOLIDS

Have a definite shape and definite volume.
The particles in a solid are packed tightly together and only vibrate relatively gently around fixed positions.

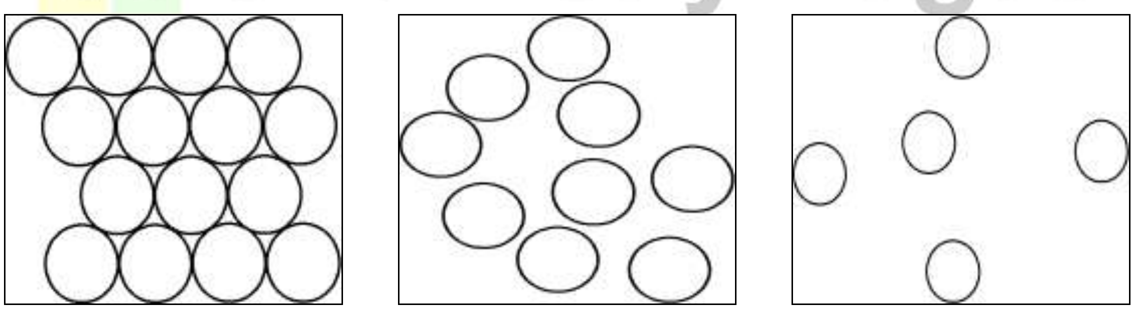
LIQUIDS

Have no shape of their own but take the shape of their container. A liquid has a definite volume.
The particles in a liquid are free to move around one another.

GASES

Have neither a definite shape nor a definite volume.
The particles in a gas spread apart filling all the space of the container available to them and interactions between the particles are considered to be negligible.

The circles in the diagrams below represent the relative positions and movements of the particles in the three states of matter. Expect to see many such *particulate representations* during the AP course.



Physical and chemical changes and properties

All matter exhibits physical and chemical properties by which it can be classified. Examples of *physical properties* are color, odor, density, hardness, solubility, melting point, and boiling point.

Chemical properties are those exhibited when a substance reacts with other substances. Examples of chemical properties are reactions with acids and bases, oxidation and reduction (REDOX) and a huge number of other chemical reactions. Changes in which the physical or chemical properties of a substance are altered are considered physical or chemical changes, respectively.

Physical change

If some aspect of the physical state of matter is altered, but the chemical composition remains the same, then the change is considered to be a physical change. The most common physical changes are changes of state. These are summarized below.

SOLID	→	LIQUID	**Melting**
LIQUID	→	GAS	**Boiling**
GAS	→	LIQUID	**Condensing**
SOLID	→	GAS	**Sublimation**
GAS	→	SOLID	**Reverse sublimation or deposition**
LIQUID	→	SOLID	**Freezing**

In solids, the particles have relatively little energy and vibrate around fixed positions. If a solid is heated, the particles gain energy, move around move, and eventually gain enough energy to break away from their fixed positions and form a liquid. Continued heating leads to the liquid particles gaining sufficient energy to break away from one another and form a gas. In a gas the particles move freely and with relatively large amounts of energy.

Chemical change

In a chemical change, which is often called a chemical reaction, the atoms of a substance are rearranged to form new substances. A chemical change requires that the new substance or substances formed have a different chemical composition to the original substance or substances. Chemical changes are often accompanied by observable changes such as color changes and energy changes.

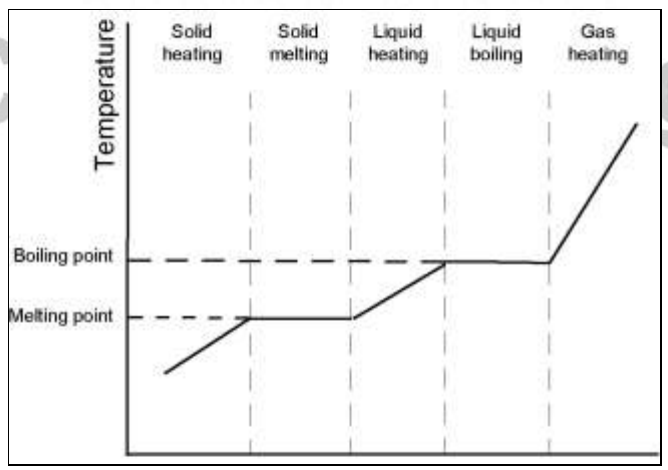
Heating and cooling curves

It is possible to investigate the changes from one phase to another when the pressure is kept constant and energy is added or removed. In these cases only the temperature is changed and heating and cooling curves result. Starting with a solid below its melting point the following effects can be observed.

1. The temperature of the solid increases at a constant rate until it begins to melt.
2. When melting begins, the temperature is constant until the solid has all turned to liquid.
3. The temperature of the liquid increases at a constant rate until it begins to boil.
4. When boiling begins, the temperature is constant until the liquid has all turned to gas.
5. The temperature of the gas increases at a constant rate.

In summary, energy is either being used to change the temperature but not the phase, or it is being used to change the phase and not the temperature. A plateau represents a stage when two phases exist together with one another and the phase change is occurring.

Heating curve



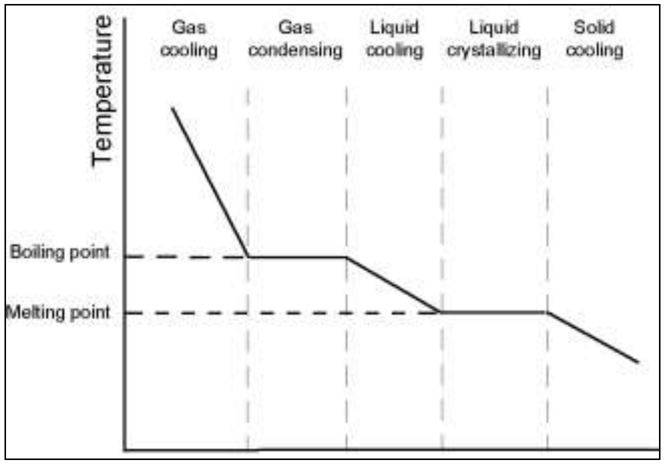
Energy added

In the regions where the temperature of the solid, liquid or gas is being increased, the amount of energy being added is, $q = m\,c\,\Delta T$. where, q = energy, m = mass, c = specific heat capacity of the substance and $\Delta T = T_{final} - T_{initial}$, i.e., the change in temperature.

Where the solid is melting, the amount of energy being added is, $q = (\Delta H_{fusion})(\text{moles})$. where, ΔH_{fusion} is the molar enthalpy of fusion and is the energy absorbed when 1 mole of a solid melts.

Where the liquid is boiling the amount of energy being added is, $q = (\Delta H_{vaporization})(\text{moles})$, where, $\Delta H_{vaporization}$ is the molar enthalpy of vaporization and is the energy absorbed when 1 mole of a liquid vaporizes.

Cooling curve



Energy removed

A cooling curve shows the same process as the heating curve only in reverse, where the energy is released rather than being absorbed.

Notes:

1. The mole is a unit used in chemistry to denote the "amount" of a substance present. This is discussed in much greater depth later in the course, but for now it is sufficient to think of it simply as an 'amount'.

2. In the equations above;

   q = Enthalpy or Energy or Heat (although technically very different, we will tend to use these words interchangeably). Energy that is absorbed (taken in) is referred to as being part of an ENDOTHERMIC change and is given a positive sign, and energy that is released (given out) is referred to as being part of an EXOTHERMIC change and is given a negative sign. Again, more later in the course.

   m = mass

   c = specific heat capacity, a constant that is different for different substances and is defined as the energy required to raise the temperature of 1 g of a substance by 1 degree Celsius.

   ΔT = change in temperature

ΔH_{fusion} = the enthalpy of fusion, a constant that is different for different substances and is defined as the energy required to convert 1 mole of solid to one mole of liquid and vice-versa.

$\Delta H_{vaporization}$ = the enthalpy of vaporization, a constant that is different for different substances and is defined as the energy required to convert 1 mole of liquid to one mole of gas and vice-versa.

Elements, atoms, mixtures and compounds

An *element* is defined as a substance that cannot be broken down into other substances by chemical means. Any single element is comprised of only one type of *atom*. The elements are displayed on the periodic table.

A *compound* is formed when a number of these elements bond together. Compounds always have a fixed composition of atoms, i.e., they always contain the same, definite amount of each element's atoms. For example, a water molecule always contains two hydrogen atoms bonded to one oxygen atom, and it always has the formula, H_2O. When the ratio of each type of atom is fixed within a compound, so is the ratio of the masses of the atoms. If that ratio changes, then the chemical formula changes, and the substance ceases to be water. All pure substances are either elements or compounds.

Unlike a pure compound or element, a *mixture* has varying composition and is made up of a number of pure substances. Mixtures are either;

Homogeneous, with a uniform in composition throughout a given sample but with a composition and properties that vary from one sample to another, for example, a solution of salt water taken from different bodies of water in different locations, or

Heterogeneous, with separate, distinct regions within the sample with a composition and properties that vary from one part of the mixture to another, for example, a chocolate chip cookie.