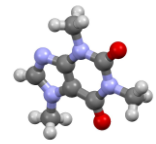
PHYSICAL SCIENCE 20 (PRE-AP CHEMISTRY)

INTRODUCTION

MATTER AND MEASUREMENT



THE STUDY OF CHEMISTRY

IT BEGINS…

I CAN…

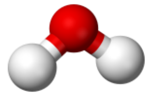
- Briefly **describe chemistry**, chemists and some major aspects of the field.
- **Classify matter** into organized and discreet groups.
- **Describe properties** of matter.
- Describe and **convert units** of measure with respect to **significant figures**.

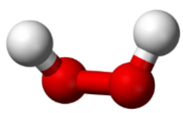
ATOMIC & MOLECULAR PERSPECTIVE

- **Matter** is the physical material of the universe – anything that has mass and takes up space.
- A **property** is any characteristic that distinguishes one piece of matter from the next (i.e., state, color, flammability)
- The matter around us is the different combinations of roughly 100 different **elements**.
- **Elements** are built of atoms, an incredibly small part of matter.
- **Molecules** (combination of two or more atoms) joined together in a specific shape.

COMBINATION OF ATOMS

- Even a small change atoms can cause a large change in properties.
  - Without water (H_2O) most life on Earth would nice be able to survive. Peroxide (H_2O_2) differs by only a single oxygen atom but can cause deadly acid burns and is toxic.

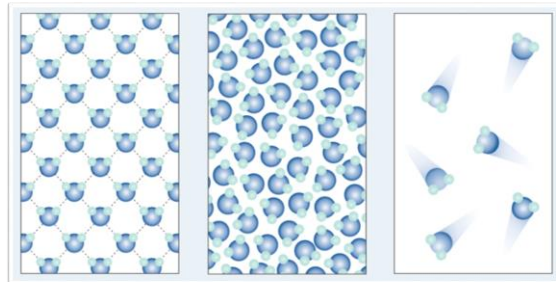




CLASSIFICATION OF MATTER

STATES OF MATTER

- There are three states of matter – solids, liquids and gases.
  - **Gas** – have no defined shape (fill container). Gases can be compressed to occupy a smaller volume or expand to fill larger volumes.
  - **Liquids** – have a specified volume independent of it's container but not definite shape, it takes the shape of it's container.
  - **Solid** – has a definite volume and shape.
- It is important to note that neither solids nor liquids exhibit much compressibility.



PURE SUBSTANCES

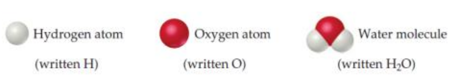
- A **pure substance** (usually just called a substance) is matter that has distinct properties regardless of sample.
  - Water* in Wascana Lake has the same properties as water in Lake Erie or the Amazon river.
- All substances can be broken down into **elements** or **compounds**.
  - Elements are substances that cannot be decomposes into simple elements (hydrogen, or oxygen)
  - Compounds are substances composed of two or more elements (two or more *atoms*).
    - **Mixtures** are combination of two or more substances that retain it's chemical identity.

ELEMENTS

- Currently there are **118 elements known** – though their abundance varies greatly.
  - 90% of the Earth crusts is made of only 5 lements: oxygen, silicon, aluminium, iron and calcium.
- The symbol for each of these elements is one or two letters, the first one always being capitalized.
- Most of the elements have named derived from English although some were named in foreign languages (e.g., Iron's symbol is Fe from the latin 'Ferrum' meaning iron)

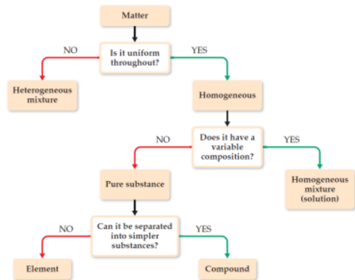
COMPOUNDS

- Most elements interact with other elements to form **compounds**.
  - Hydrogen gas (H_2) reacts with oxygen gas (O_2) to form water (H_2O).
- The compound created has **new chemical properties**, quite distinct from the component elements



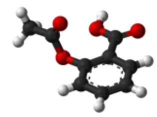
MIXTURES

- Most of the matter we encounter consists of **mixtures** of different substances.
- Each substance (component) in a mixture **retains it's chemical identity**.
- The characteristics of a mixture depend on the type of sample taken. For example the characteristics of a chocolate chip-cookie (mixture of organic molecules) can be 'dough' like or 'chocolate' like.
  - **Heterogeneous mixtures** are mixtures that are visibly different. (e.g., granite)
  - **Homogeneous mixtures** are mixtures where only one component is visible (e.g., coke)



EXAMPLES: CLASSIFICATION

- Aspirin is composed of 60.0% carbon, 4.5% hydrogen, and 35.5% oxygen by mass, regardless of the type of aspirin (brand-name or otherwise). Use the previous flow chart to classify aspirin.



EXAMPLE: CLASSIFICATION OF MATTER

Use the classification tree to classify the following:

- Tide (Pod) Laundry Detergent
- Stainless Steel
- Copper Wires
- Paper

PROPERTIES OF MATTER

PROPERTIES OF MATTER

- Every substance has unique **properties**.
- **Physical properties** can be determined without altering the identity or composition of the material (e.g., color, melting point, density).
- **Chemical properties** are those that describe how a substance changes (e.g., flammability, solubility)

INTENSIVE OR EXTENSIVE PROPERTIES

- Intensive Properties are those that do not depend on the amount of matter being examined.
  - 100 g of water will vaporize at the same *temperature* as 1500 grams will.
- Extensive Properties are those that depend on the amount of matter being examined.
  - The amount of heat energy required to vaporize 100 g of water is much less than the heat energy needed to vaporize 1500 grams.

PHYSICAL AND CHEMICAL CHANGES

- Matter can undergo two types of change, **physical or chemical**.
- **Chemical changes** (reaction) are those in a which component substances are rearranged into **chemically different substances**.
  - When iron (Fe) is in the presence of water and oxygen it begins to form rust (Fe_2O_3) changing the *color* from grey to brown-red.
- **Physical changes** occur when a substance changes it's appearance but not it's composition. **Chemically the material remains** the same after the change.
  - Water in the liquid phase is made of two hydrogens and a single oxygen. Water in the gaseous phase is made of the same elements, in the same ratio.

UNITS OF MEASURE

UNITS OF MEASURE

- Many properties of matter are **quantitative**, meaning they have numbers associated with them (e.g., *23* gram, *1.23* milliliters)
- In science, the metric system is used which is based on the **7 SI units**.

TABLE 1.4 • SI Base Units

Physical Quantity	Name of Unit	Abbreviation
Mass	Kilogram	kg
Length	Meter	m
Time	Second	s or sec
Temperature	Kelvin	K
Amount of substance	Mole	mol
Electric current	Ampere	A or amp
Luminous intensity	Candela	cd

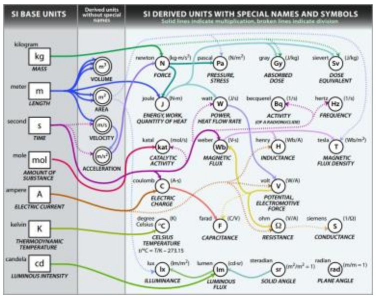


TABLE 1.5 • Prefixes Used in the Metric System and with SI Units

Prefix	Abbreviation	Meaning	Example
Peta	P	10^{15}	1 petawatt (PW) $= 1 \times 10^{15}$ watts[a]
Tera	T	10^{12}	1 terawatt (TW) $= 1 \times 10^{12}$ watts
Giga	G	10^{9}	1 gigawatt (GW) $= 1 \times 10^{9}$ watts
Mega	M	10^{6}	1 megawatt (MW) $= 1 \times 10^{6}$ watts
Kilo	k	10^{3}	1 kilowatt (kW) $= 1 \times 10^{3}$ watts
Deci	d	10^{-1}	1 deciwatt (dW) $= 1 \times 10^{-1}$ watt
Centi	c	10^{-2}	1 centiwatt (cW) $= 1 \times 10^{-2}$ watt
Milli	m	10^{-3}	1 milliwatt (mW) $= 1 \times 10^{-3}$ watt
Micro	μ[b]	10^{-6}	1 microwatt (μW) $= 1 \times 10^{-6}$ watt
Nano	n	10^{-9}	1 nanowatt (nW) $= 1 \times 10^{-9}$ watt
Pico	p	10^{-12}	1 picowatt (pW) $= 1 \times 10^{-12}$ watt
Femto	f	10^{-15}	1 femtowatt (fW) $= 1 \times 10^{-15}$ watt
Atto	a	10^{-18}	1 attowatt (aW) $= 1 \times 10^{-18}$ watt
Zepto	z	10^{-21}	1 zeptowatt (zW) $= 1 \times 10^{-21}$ watt

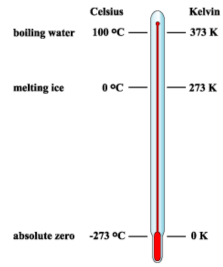
EXAMPLE: UNITS OF MEASURE

- What is the name of the unit that equals:
  - 10^{-9} gram
  - 10^{-6} second
  - 10^{-3} meter

MEASUREMENTS

- While there are many measurements that can be made this course will focus primarily on:
  - **Length** – the measurement of distance (base unit m)
  - **Mass** – the measurement of matter (base unit is Kg)
  - **Temperature** – measure of hotness or coldness (measured in Celsius or Kelvin)
    - The SI unit for temperature is Kelvin, which has the same scale but different starting point.
    - To covert C to Kelvin (or reverse)

K = °C + 273.15



DERIVED SI UNITS

- While there are only 7 base units in the SI system, there are **derived units** which are the combination of two or more SI units (e.g., meters per second, grams per litre). Examples include:
  - **Volume** – volume is a measurement of distance in three dimensions – a cube. A 1 cm by 1 cm by 1 cm is 1 cm^3 and is equivalent to a millilitre.
  - **Density** – defined by the a unit of mass per unit of volume according to the following formula. Common units include g/cm^3 or g/mL

$$\text{Density} = \frac{\text{mass}}{\text{volume}}$$

EXAMPLES: DETERMINING DENSITY

a. Calculate the density of mercury if 1.00 x 10^2 g occupies a volume of 7.36 cm^3

b. Calculate the volume of 65.0 g of liquid methanol if its density is 0.791 mg/mL.

c. What is the mass in grams of a cube of gold (density = 19.32 g/cm^3) if the length of the cube is 2.00 cm?

UNCERTAINTY IN MEASUREMENT

UNCERTAINTY IN MEASUREMENT

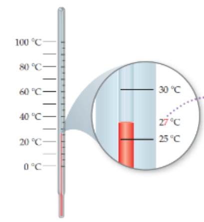
- There are two types of numbers, **inexact** and **exact**.
- **Exact** numbers are those with a defined quality. For example a dozen eggs is always 12 and a kilogram is 1000 g.
- **Inexact** numbers are those that are measured. Regardless of the measurement made, it is always inexact. Uncertainties will always exists in measured quantities.
  - Imagine 10 people are weighing the same dime. It is likely that there is some discrepancy between measurements due to human, equipment or random errors!

ACCURACY AND PRECISION

- **Precision** refers to how closely individual measurements align with each other.
- **Accuracy** refers to how closely each measurement is to the correct, or 'true' value.
  - In the lab we generally conduct trials and average the result and use standard deviation to determine the precision.

SIGNIFICANT FIGURES

- Suppose you measure something on an electronic scale capable of reading a value to he nearest 0.0001 g. You would report the mass as 2.3456 ± 0.0001g.
  - For the most part, all measurements do not express the ± as there is always uncertainty in the measurement.
- In the diagram to the right the temperature is between 25 and 30, estimated to be 27C, the second digit of this measurement being **uncertain**.



SIGNIFICANT FIGURES

- To determine the number of significant figures in a number use the following rules:
  1. All nonzero numbers are significant. (123 g = 3SF, 1838 kg = 4SF)
  2. Zeroes between nonzero numbers are significant. (e.g., 1005 kg = 4SF, 7.03 mL=3SF)
  3. Zeroes at the beginning of a number are never significant, they only indicate the position of a decimal point. (e.g., 0.02 g = (1 SF), 0.0072 s = 2 SF)
  4. Zeroes at the end of a number are significant if the number contains a decimal. (0.200 mL = 3 SF)

SIGNIFICANT FIGURES

- There can be some gray area with numbers that have a measured zero on the end. Remember, zeroes on the end of a number are only significant if there is a decimal.
  - When no decimal is placed in an experimental measurement the zeroes are considered to be non-significant. The number 10 300 grams can be written using scientific notation, each with a different number of SF based on the number of 'significant zeroes'

1.03×10^4 g	(three significant figures)
1.030×10^4 g	(four significant figures)
1.0300×10^4 g	(five significant figures)

EXAMPLE: SIGNIFICANT FIGURES

- How many significant figures are in each of the following?
  a. 4.003
  b. 6.023×10^{23}
  c. 5000

SIGNIFICANT FIGURES IN OPERATIONS

1. For addition and subtraction the result has to have the same number of decimal places as the value with the fewest decimal places. When the result contains more, round off to the appropriate digit.

This number limits	20.42	← two decimal places
the number of significant	1.322	← three decimal places
figures in the result →	83.1	← one decimal place
	104.842	← round off to one decimal place (104.8)

SIGNIFICANT FIGURES IN OPERATIONS

2. For multiplication and division the result contains the same number of significant figures as the term with the least. When it contains more than the term with the least number of significant figures it must be rounded. Be sure to round properly (use non-significant zeroes as placeholders!)

$$\text{Area} = (6.221\,\text{cm})(5.2\,\text{cm}) = 32.3492\,\text{cm}^2 \Rightarrow \text{round off to } 32\,\text{cm}^2$$

NOTES ABOUT SIGNIFICANT FIGURES

- For **addition and subtraction** is the number of **decimals** that determines the number of digits, while it is **significant figures** that determines the number of digits for **multiplication and division**.
- All **exact numbers** are considered have an **infinite number** of significant figures.
- When rounding numbers use the leftmost digit to be removed:
  - If the leftmost digit is less than 5, the preceding number is unchanged; (e.g., 7.248 rounded to two SF becomes 7.2)
  - If the leftmost digit is greater than 5, the preceding number is increased by 1. (e.g., 4.258 rounded to three SF is 4.26)

EXAMPLE: SIGNIFICANT FIGURES

- The width, length and height of a box are 15.5 cm, 27.3 cm, and 5.4 cm respectively. Calculate the volume of the box, using the correct number of significant figures in your answer.

EXAMPLE: SIGNIFICANT FIGURES

- A student is nearly late for Chemistry class and her classroom is 100.00 m away. She has 3 lates in the class and needs the incentive. It takes her 10.5 s to run to the classroom. What is her speed in meters per second (to the correct number of significant figures!).

EXAMPLE: SIGNIFICANT FIGURES

- A gas at 25 °C fills a container whose volume is 1.05×10^3 cm^3. The container plus gas has a mass of 837.6 g. The container, when emptied of all gas, has a mass of 836.2 g. What is the density of the gas?

EXPONENTIAL RULES

YOUR MATH TEACHERS WOULD BE SO PROUD

ADDITION AND SUBTRACTION

$(a\times10^m) + (b\times10^n) = (a+b)\times10^m$

- To add terms with scientific notation (without the use of a calculator) you can use the above law.
- Change terms with unlike exponents by moving the decimal, then add or subtract as needed.

EXAMPLE: ADDITION AND SUBTRACTION

1. $4.66 \times 10^{-19}\text{kg} + 2.1 \times 10^{-19}\text{kg}$
2. $5.0 \times 10^{-7}\text{mg} + 4 \times 10^{-8}\text{mg}$

MULTIPLICATION AND DIVISION

$(a \times 10^m)(b\times10^n) = (a \times b)\times10^{m+n}$

$(a \times 10^m) \div (b\times10^n) = (a \div b)\times10^{m-n}$

- To multiply or divide use the following exponent laws.
  - The exponents do not matter!

EXAMPLE: MULTIPLICATION AND DIVISION

1. $(1.4 \times 10^2)(3 \times 10^1) =$
2. $5.0 \times 10^{-6} \div (2.5 \times 10^2) =$

DIMENSIONAL ANALYSIS

DIMENSIONAL ANALYSIS

- Dimensional analysis is a systematic method of solving mathematical problems using conversion units to 'cancel' out units.
- Using dimensional analysis helps to ensure that solutions to problems yield the proper units.
- In order to use dimensional analysis correct, you must understand and use conversion units correctly.
- A conversion factor is a fraction whose numerator are the same quantity expressed in different units.

$$\frac{2.54\,\text{cm}}{1\,\text{in.}} \quad \text{and} \quad \frac{1\,\text{in.}}{2.54\,\text{cm}}$$

GENERAL FORM

- Using this form the numerator and denominator of a conversion factor are equal, multiplying any quantity by a conversion factor is equivalent to multiplying the number 1 and so does not change the intrinsic value of the quantity.

Desired unit

$$\text{Number of centimeters} = (8.50\ \cancel{\text{in.}})\frac{2.54\ \text{cm}}{1\ \cancel{\text{in.}}} = 21.6\ \text{cm}$$

Given unit

$$\cancel{\text{Given unit}} \times \frac{\text{desired unit}}{\cancel{\text{given unit}}} = \text{desired unit}$$

EXAMPLE: DIMENSIONAL ANALYSIS

Convert the following values using dimensional analysis:

a. 100 meters per second to km per s.

b. 10.5 cm to inches (2.54 cm = 1 inch)

c. 155 pounds to kg (1 kg = 2.21 lbs)

CONVERSIONS WITH MULTIPLE STEPS

- It is often necessary to use more than one conversion factor in a single calculation.
- To achieve this simply use more conversion factors, cancelling units from left to right.

$$\text{Number of inches} = (8.00\ \cancel{\text{m}})\left(\frac{1\ \cancel{\text{cm}}}{10^{-2}\ \cancel{\text{m}}}\right)\left(\frac{1\ \text{in.}}{2.54\ \cancel{\text{cm}}}\right) = 315\ \text{in.}$$

EXAMPLE: MULTI-STEP D.A.

- The average speed of a nitrogen molecule in air at 25 °C is 515 m/s. Convert this speed to miles per hour. (Note: 1 mile = 1.6km).

I CAN…

- Briefly **describe chemistry**, chemists and some major aspects of the field.
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