INTRODUCTION

- Welcome to CH110 Group C

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- Lecture times: Monday 08:00-10:00 (ST5), Thursday 10:30-12:30 (ST 6)

INTRODUCTION

Course objectives

- to provide a firm foundation in concepts and principles of selected General Chemistry topics
- to instill in students an appreciation of the usefulness of chemistry in our daily life.

Assessment

Continuous assessment 40%, final examination 60%

Continuous assessment: Tests 30% , practicals 10%

Note: Repeaters are to do the whole course which includes laboratory work.

INTRODUCTION

Learning Resources

1. Prescribed text book:

   Chemistry (8^{th} Edition) by Steven S Zumdahl and Susan A Zumdahl, Brooks/Cole Centage Learning, Belmont, CA, USA, 2010 (available in bookshop at K275, electronic copy will be made available through class representative)

2. Recommended textbooks see list in handout
3. Internet
4. Chemistry journals

INTRODUCTION

Term 1 Topics (6 June- 14 August 2015)

- Introduction to chemistry
- Stoichiometry
- Reactions in solution
- Gases

Test 1: 11 Sept 2015

Term 2 topics (31 Aug-6 Nov 2015)

- Thermochemistry
- Atomic structure and periodicity
- Chemical bonding and molecular geometry

INTRODUCTION

- Organic chemistry

  Test 2 4 December 2015

Term 3 topics (23 Nov 2015-29 Jan 2016)

- Kinetics
- Chemical equilibrium
- Acids and Bases
- Electrochemistry

Final examinations – February 2016

Chemistry

- In this science we study **matter**, its **properties**, and its **behavior-**changes it undergoes.
- A **property** is any characteristic that allows us to recognize a particular type of matter.
- Examples of matter: paper, pen, air, water, rocks, food, clothes, bacteria, viruses, etc.
- Experiments have shown that matter is composed of various combinations of about 100 basic elements.
- Chemistry relates properties of matter to its composition Chemistry enables us to understand the properties of matter in terms of **atoms** and **molecules**, the building blocks of elements

Why study chemistry

Provides understanding of our world

- health care – antibiotics, ARVs, vitamins,
- Development of new materials-plastics, polymers,
- protection of environment-global warming, ozone layer
- provision of food-chemical fertilizers and pesticides
- Transport- car battery, fuel – petrol, diesel, biofuels

Why study chemistry-continued

Chemistry enables us to answer questions like

- Why does ice melt and water evaporate when heated?
- How does a battery generate electricity?
- Why does keeping food in a fridge slow down its spoilage?
- How do our bodies use food to maintain life?

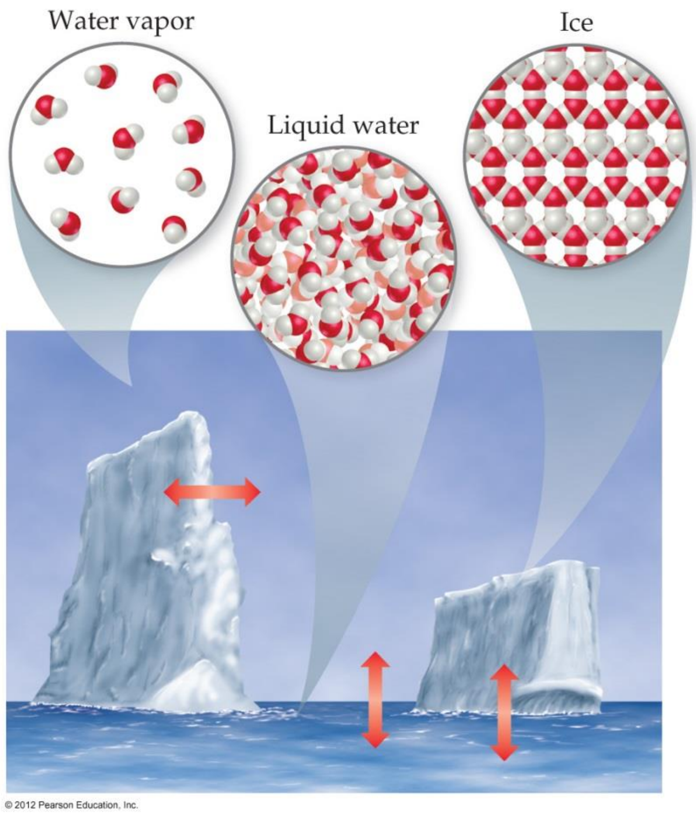
Matter

- We define **matter** as anything that has mass and takes up space.

- Examples of matter: paper, pen, air, water, rocks, food, clothes, bacteria, viruses, etc.

- Two principal ways of classifying matter – according to its physical state (solid, liquid or gas) and according to its composition (element, compound or mixture)

States of Matter



There three states of matter

Gas (vapour)-has no fixed volume or shape, can be readily compressed or expanded to occupy any volume

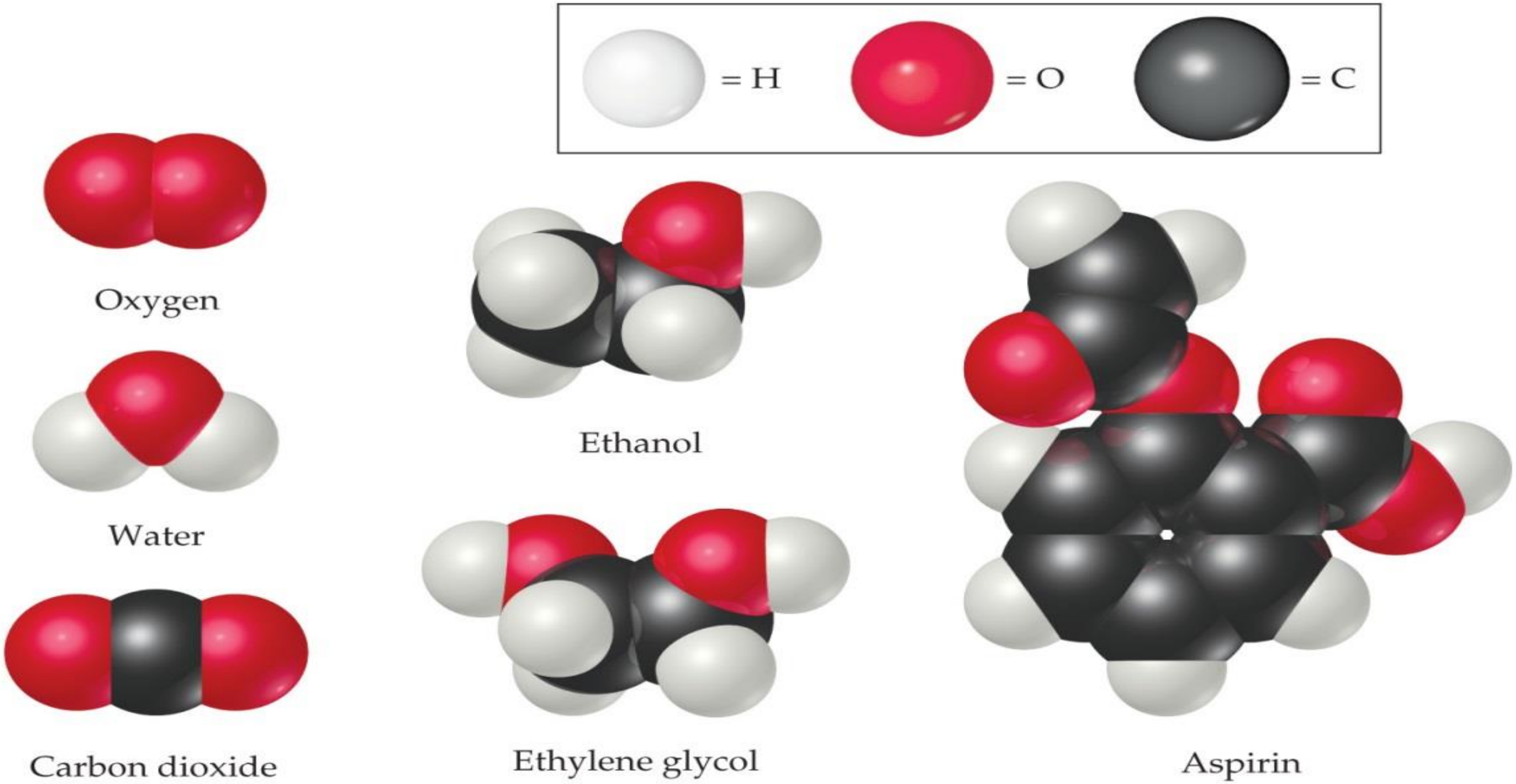
Liquid-has distinct volume but takes shape of container

Solid-has distinct volume and shape

Pure substance

- Each different pure kind of matter is called a ***substance.***
- A **pure substance** (substance) is a form of matter that has distinct properties and a composition that does not vary from sample to sample.
- Water and salt (sodium chloride) are pure substances but sea water (contains water and salt) is not.
- Other examples of substances include: oxygen, copper, iron, limestone (calcium carbonate)
- Substances are either **elements** or **compounds**

Pure substance



- **Atoms** are the building blocks of matter.
- Each **element** is made of the same kind of atom.
- A **compound** is made of two or more different kinds of elements.

ELEMENTS

- There are currently about 117 known elements
- Smallest particle of an element is called an atom
- Each **element** is made of the same kind of atom.
- No systematic way of naming elements- some are from ancient times e.g. copper from Cyprus, gold from an old English word meaning yellow. Some from a characteristic property of element e.g. Chlorine derived from Greek meaning yellow green. Others honour names of places or people e.g. americium, berkelium, californium, einsteinium and curium.

ELEMENTS

- The International Union of Pure and Applied Chemistry (IUPAC) is the international body that currently among other duties, approves names for elements.
- Each element is represented by a chemical symbol made up of one or two letters
- Many of the symbols are the first one or two letters of the element’s name:

hydrogen H	carbon C	nitrogen N	oxygen O
helium He	aluminium Al	nickel Ni	silicon Si

ELEMENTS

- Other symbols are taken from the element's name in Latin, German or Greek:

Copper Cu (from cuprum)   gold Au (form aurum)

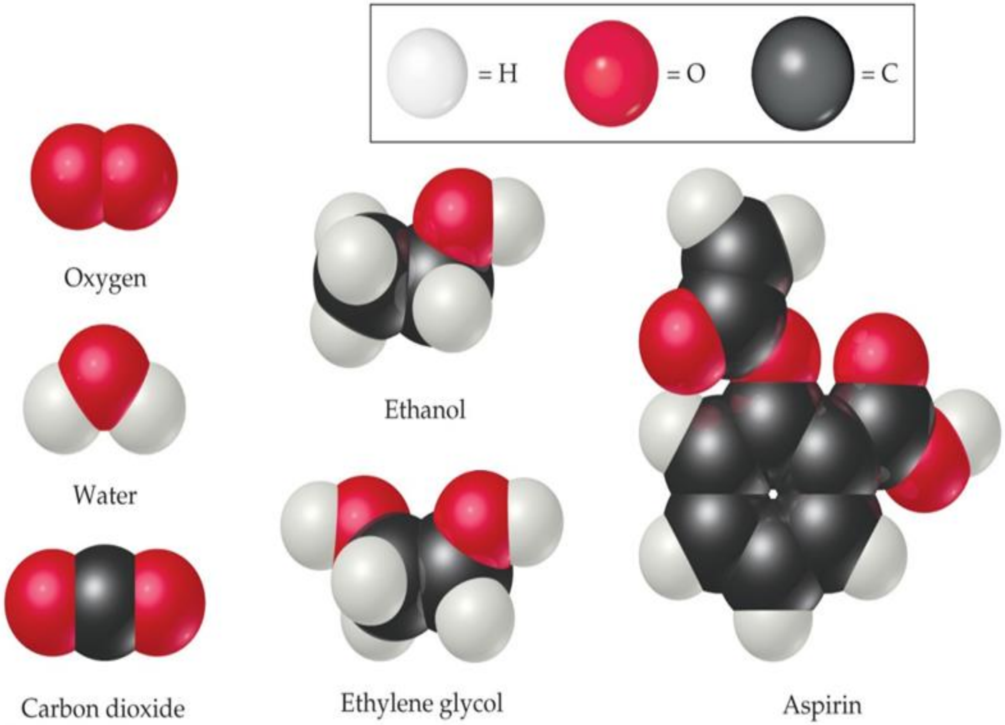
iron Fe (from ferrum)   Potassium K (from kalium)

Sodim Na (from natrium) Silver Ag (from argentum)

- NOTE: The first letter in a chemical symbol is always uppercase and the second letter lowercase. Thus Co and CO mean different things. Co – chemical symbol of the element cobalt whereas CO is the formula of a compound of carbon (C) and oxygen (O)

COMPOUNDS

- Most elements interact with each other to form compounds.
- The observation that elemental composition of a pure compound is always the same is known as **the law of constant composition** or **law of definite proportions**.



Mixtures

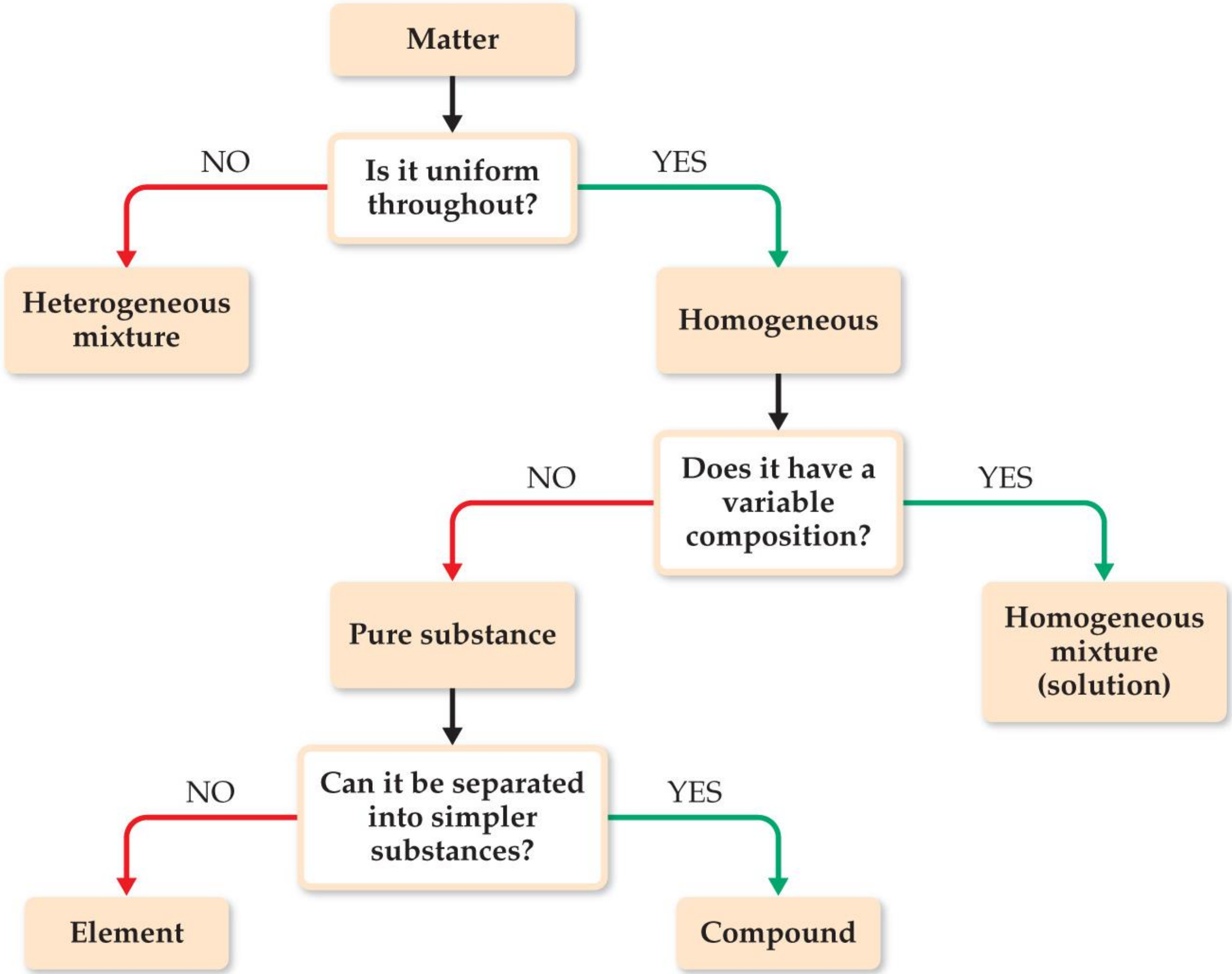
- Most of matter consists of mixtures of different substances, e.g. air (oxygen, nitrogen and other gases), sea water(water, sodium chloride and other salts), brass (copper and zinc)
- Differences between mixtures and compounds are listed in table below

Mixture	**Compound**
-Composition is variable e.g. salt and sand mixture can have variable composition -Properties are related to those of its components, salt is soluble in water sand is not. -can be readily separated using physical techniques	- Composition is fixed e.g. water, ratio of hydrogen to oxygen is fixed -properties are unlike those of components, hydrogen and oxygen are gases whilst water is a liquid. - cannot be separated using physical techniques.

Mixtures

- A mixture is classified as ***heterogeneous*** if its components can be identified with the naked eye or a microscope e.g. salt/sand, milk, rocks.
- In a ***homogeneous*** mixture the components are thoroughly mixed at molecular level and cannot be identified even with a powerful microscope e.g. sugar/water, beer, air.
- Homogeneous mixtures are called ***solutions.***
- The figure below summarizes the classification of matter into elements, compounds and mixtures

Classification of Matter



Properties and Changes of Matter

Types of Properties

- Physical Properties…
  - Can be observed without changing a substance into another substance.
    - Boiling point, density, mass, volume, etc.

- Chemical Properties…
  - Can *only* be observed when a substance is changed into another substance.
    - Flammability, corrosiveness, reactivity with acid, etc.

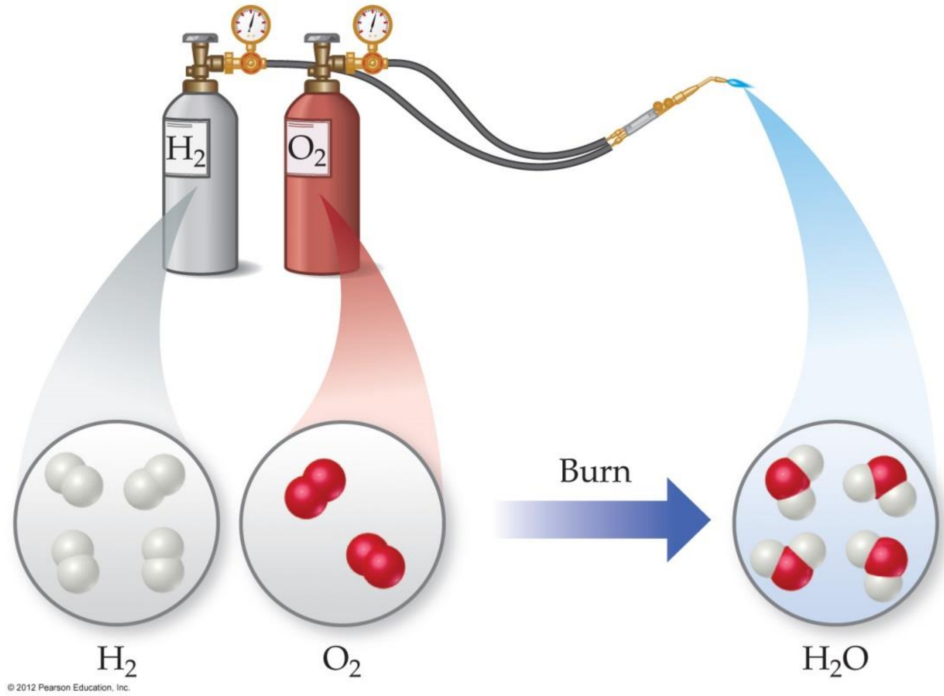
Types of Properties

- Intensive Properties…
  - Are independent of the amount of the substance that is present.
    - Density, boiling point, color, etc.
- Extensive Properties…
  - Depend upon the amount of the substance present.
    - Mass, volume, energy, etc.

Types of Changes

- Physical Changes
  - These are changes in matter that do not change the composition of a substance.
    - Changes of state e.g. solid→liquid, temperature, volume, etc.
- Chemical Changes
  - Chemical changes result in new substances.
    - Combustion, oxidation, decomposition, etc.

Chemical Reactions

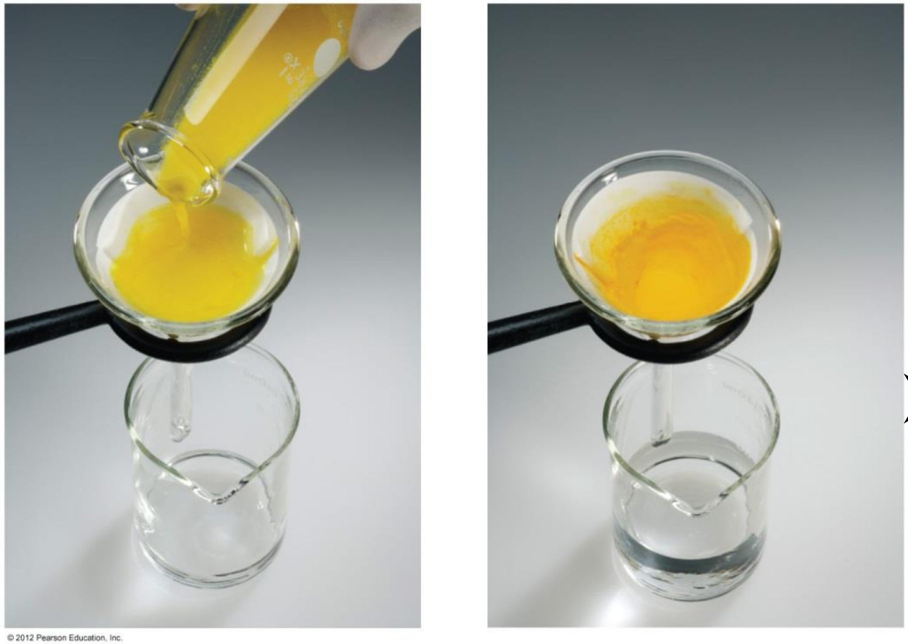


In the course of a chemical reaction, the reacting substances are converted to new substances.

Separation of Mixtures

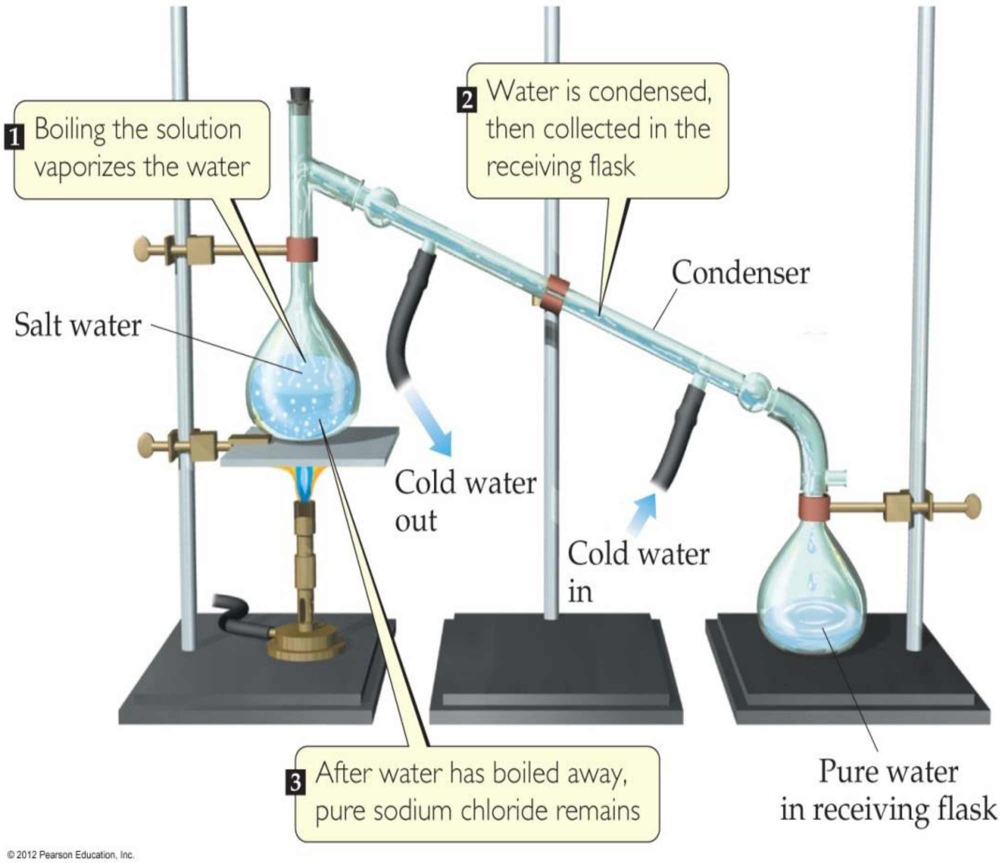
- Since components in a mixture retain their properties, they can be separated by taking advantage of their properties.
- A mixture of gold and iron fillings can be separated by sorting by colour or using a magnet which would attract iron and leave behind the gold.
- A mixture of sand and sugar can be separated by adding water, sugar dissolves sand does not. Resulting mixture of water, sand and sugar can be separated by **filtration** followed by **distillation**

Filtration



- In filtration, solid substances are separated from liquids and solutions using a mesh or filter paper.
- Sand remains on the mesh while sugar solution passes through

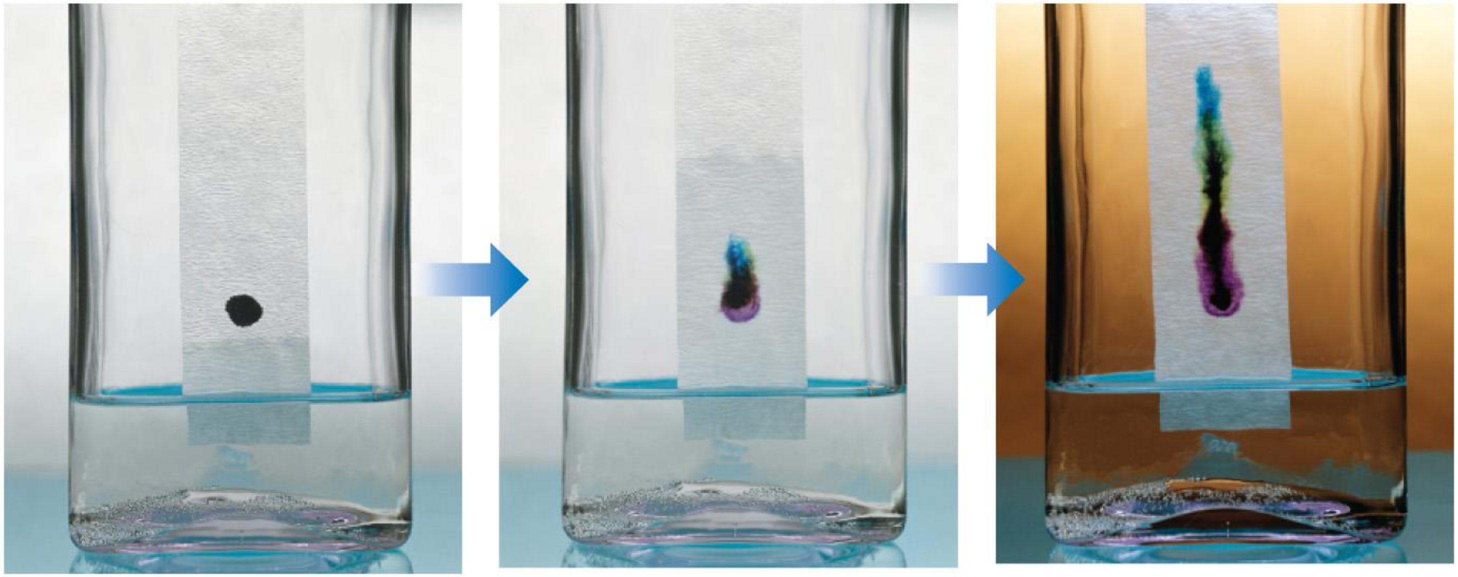
Distillation



- Distillation uses differences in the boiling points of substances to separate a homogeneous mixture into its components.
- Sugar solution water evaporates and sugar is left behind.

Chromatography

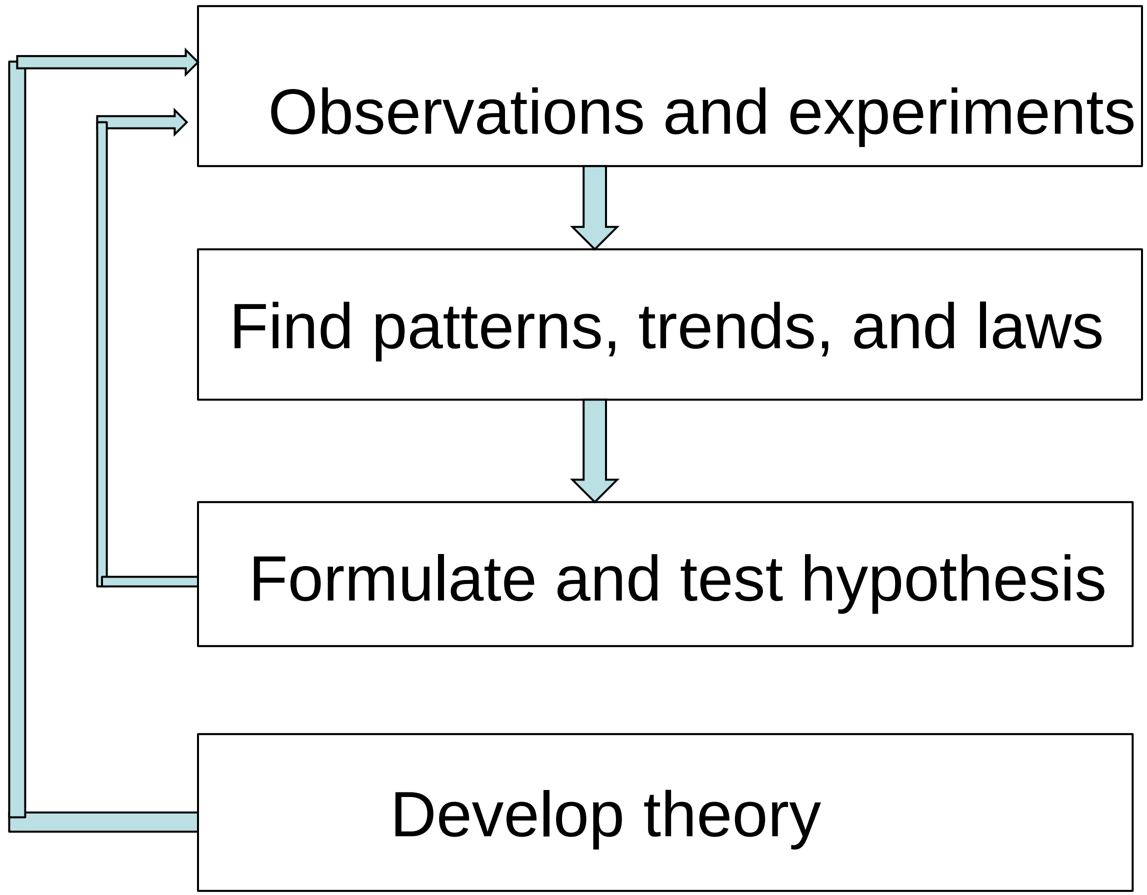
- This technique separates substances on the basis of differences in ability to adhere to a solid such paper and starch.
- E.g. ink, can be separated into its componenents



SCIENTIFIC METHOD

1. Making observations
   - **Qualitative** e.g. sky is blue, ice floats on water
   - **Quantitative** e.g. water boils at 100 ^{o}C, (number and unit). A quantitative observation is called a **measurement**.
2. Find patterns, trends, and laws – a law is a concise verbal or mathematical equation that summarizes a range of observations e.g. energy cannot be created or destroyed.
2. Formulating hypothesis-A **hypothesis** is a possible explanation of an observation
3. Performing experiments to test if hypothesis is valid
4. Development of a theory- A **theory** is an interpretation, a possible explanation of why nature behaves in a particular manner.
5. Theory is tested by making more observations and modified as needed

SCIENTIFIC METHOD



Units of Measurement

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

- *Système International d'Unités*
- A different base unit is used for each quantity.

SI Units

- The unit of mass (the quantity of material) is 1 kilogram (1 kg).
- The unit of length (the distance between two points) is 1 meter (1 m).
- The unit of time (the duration of an event) is 1 second (1)
- Distinguish between mass and weight of an object:
  - mass refers to ***quantity of matter*** it contains
  - weight refers to ***gravitational pull*** it experiences
  - Mass and weight are proportional to each other but are not identical, e.g. an astronaut would have the same mass on Earth and Mars but his weight would be less on Mars than on Earth because Mars has a lower gravity.

Metric System

Prefixes convert the base units into units that are appropriate for the item being measured.

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

[a]The watt (W) is the SI unit of power, which is the rate at which energy is either generated or consumed. The SI unit of energy is the joule (J); $1\ J = 1\ kg \cdot m^2/s^2$ and $1\ W = 1\ J/s$.

[b]Greek letter mu, pronounced "mew."

DERIVED UNITS

- The kilogram is out of line; it is a base unit but it already has a prefix (k),i.e.1 kg = 10^3 g.
- In calculations replace the prefixes by the power of 10 they represent and then proceed numerically.
- Note that an expression like cm^3 means $(cm)^3$:

$$1\ cm^3 = (1\ cm) \times (1\ cm) \times (1 cm)$$
$$= (10^{-2}\ m) \times (10^{-2}\ m) \times (10^{-2}\ m)$$
$$= 10^{-6}\ m^3$$

- Multiply and divide units like numbers e.g. 6 g divide by 2 g gives 3, the unit cancel: $\frac{6\ g}{2\ g} = 3$

DERIVED UNITS

- A box of sides 1.0 m, 2.0 m, and 3.0 m has a volume, V:

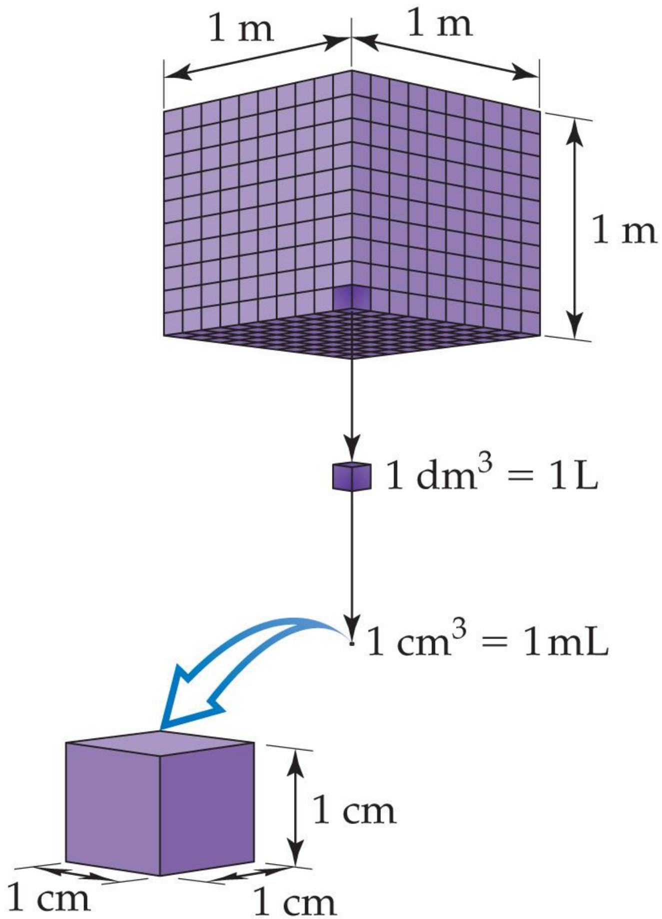
$$V = 1.0 \text{ m} \times 2.0 \text{ m} \times 3.0 \text{ m} = 6 \text{ m}^3$$

Units derived in this way are called **derived units**

- It is common in chemistry to report volumes of liquids in litres (L), where I L is exactly 10^3 cm^3.
- Note that 1 mL (1 millilitre) is exactly the same as 1 cm^3.

Volume

- The most commonly used metric units for volume are the liter (L) and the milliliter (mL).
  - A liter is a cube 1 decimeter (dm) long on each side.
  - A milliliter is a cube 1 centimeter (cm) long on each side.



DERIVED UNITS

- A derived unit may include a base unit raised to a negative power; e.g. the unit for **density,**

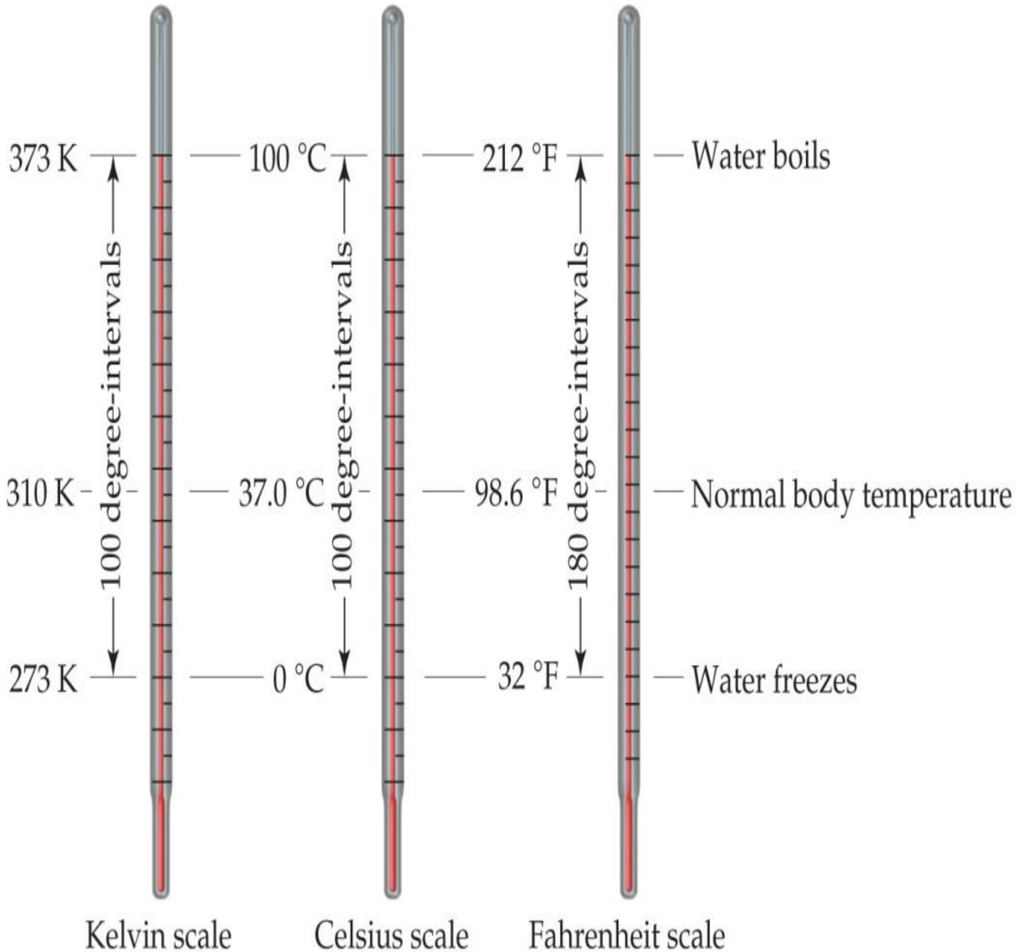
$$\boldsymbol{density = \frac{mass}{volume} = \frac{1\,kg}{1\,m^3} = kgm^{-3}}$$

- Density of water is about 1 g cm^{-3} or 1 kg L^{-1}.

Temperature

- Temperature is a measure of the hotness and coldness of an object.
- Temperature determines direction of heat flow: heat flows from hot to cold
- In scientific measurements, the Celsius and Kelvin scales are most often used.
- The Celsius scale is based on the properties of water.
  - 0 °C is the freezing point of water.
  - 100 °C is the boiling point of water.

Temperature



- The kelvin is the SI unit of temperature-must be used in all calculation.
- It is based on the properties of gases.
- There are no negative Kelvin temperatures, 0 K is the lowest possible temperature-absolute temperature scale.
- K = °C + 273.15

Temperature

Example

Ethylene glycol freezes at -11.5 ^oC. What is the freezing point in K?

Solution

We use

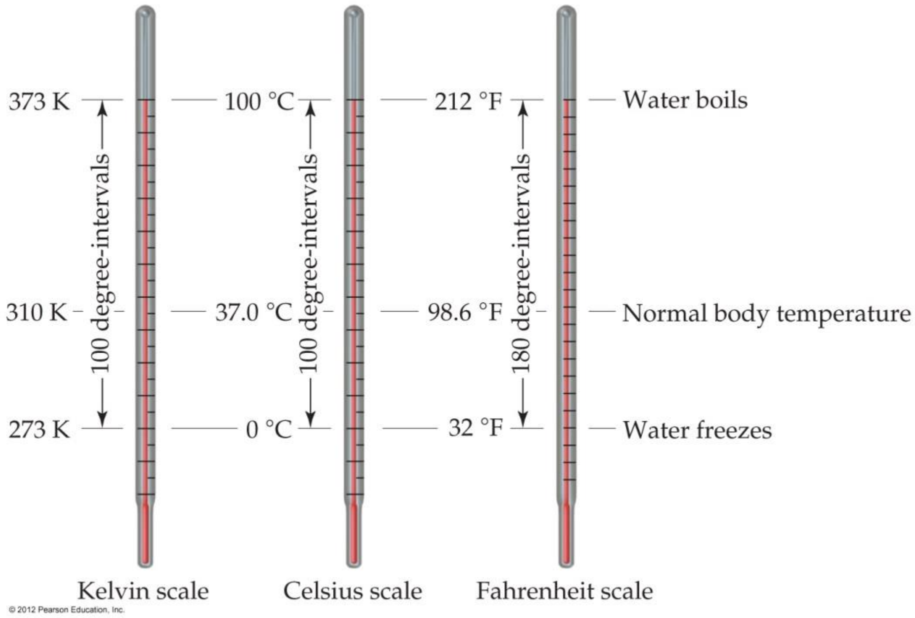
$$K = {}^\circ C + 273.15$$

$$= -11.5\ oC + 273.15$$

$$= 261.7\ K$$

NOTE: do not use the word degree or the degree sign (^{o}) with the kelvin unit

Temperature

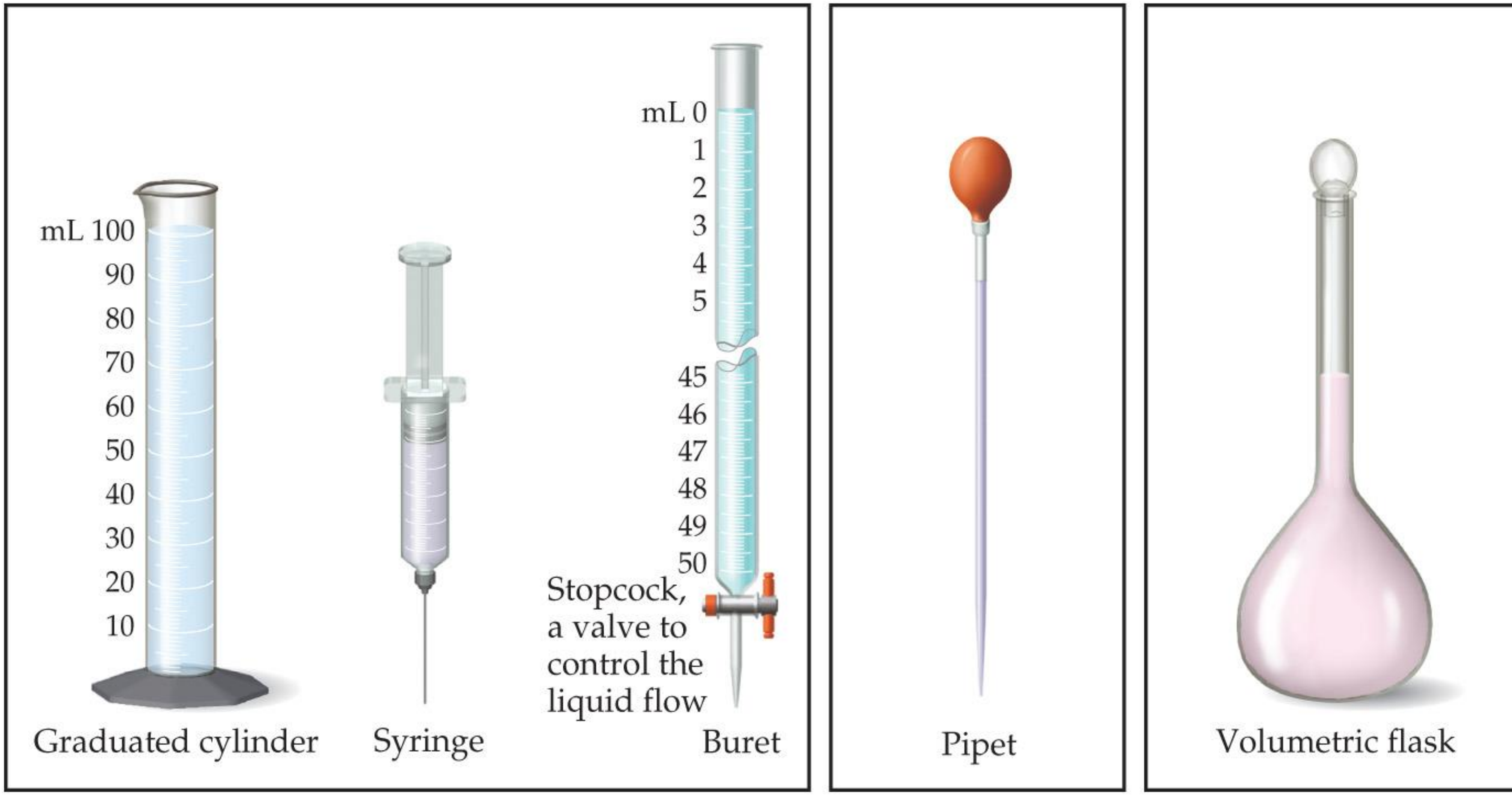


- The Fahrenheit scale is not used in scientific measurements.
- °F = 9/5(°C) + 32
- °C = 5/9(°F − 32)

Uncertainty in Measurement

Uncertainty in Measurements

Different measuring devices have different uses and different degrees of accuracy.



These deliver **variable** volumes

Pipet **delivers** a **specific** volume

Volumetric flask **contains** a specific volume

Significant Figures

- The term **significant figures** refers to digits that were measured.
- Convention: the last digit in the data is imprecise to the extent of ± 1 of that figure.
- **Example:** A measurement reported as 1.2 cm^3 means that the volume lies between 1.1 cm^3 and 1.3 cm^3 and a mass reported as 1.78 g means that it lies between 1.77 g and 1.79 g.
- The digits in a reported figure are called the **significant figures. Hus** 1.2 cm^3 has 2 significant figures and 1.78 g has 3 significant figures

Significant Figures

The significance of zeroes in a measurement can be a source of difficulty: some zeroes are measured and hence significant, others are serve to mark decimal point.

1. All nonzero digits are significant.
2. Zeroes between two significant figures are themselves significant, e.g. 80.2 kg has 3 significant figures.
3. Zeroes at the beginning of a number are never significant, e.g. 0.0025 g (2 significant figures).
4. Zeroes at the end of a number are significant if a decimal point is written in the number, 0.020 cm (2 significant figures).

Significant Figures

- Use exponential notation to clearly indicate whether zeroes at the end of a number without a decimal point are significant not, e.g. 10 300 g can be:
  - 1.03×10^4 (3 significant figures)
  - 1.030×10^4 (4 significant figures)
  - 1.0300×10^4 (5 significant figures)
- In science we distinguish between the results of **measurements** which are always uncertain and the results of **counting,** which are exact, e.g.12 eggs means exactly 12 eggs and not a number somewhere between 11 and 13.
- When rounding calculated numbers, we pay attention to significant figures so we do not overstate the accuracy of our answers.

Significant Figures

- When addition or subtraction is performed, answers are rounded to the least significant **decimal place**.

examples

$$
\begin{array}{r}
20.42 \\
1.322 \\
\underline{83.1} \\
104.842
\end{array}
$$

Result is 104.8

- When multiplication or division is performed, answers are rounded to the number of digits that corresponds to the ***least*** **number of significant figures** in any of the numbers used in the calculation.

$$\frac{1.78\ g}{1.2\ cm^3} = 1.483333333 gcm^3$$

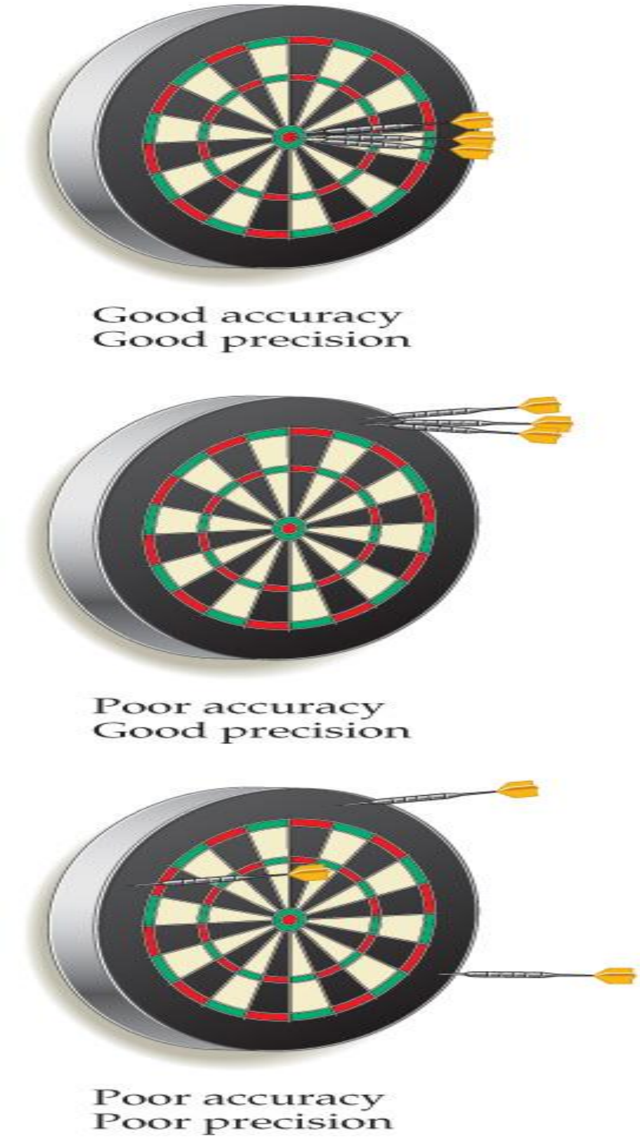
Result is 1.5 gcm^3.

Accuracy versus Precision

- **Accuracy** refers to the proximity of a measurement to the true value of a quantity.
- **Precision** refers to the proximity of several measurements to each other.
- Scientists perform several measurements or “trials” and compute the average value and standard deviation:

Standard deviation is a measure of precision-how close to each other.

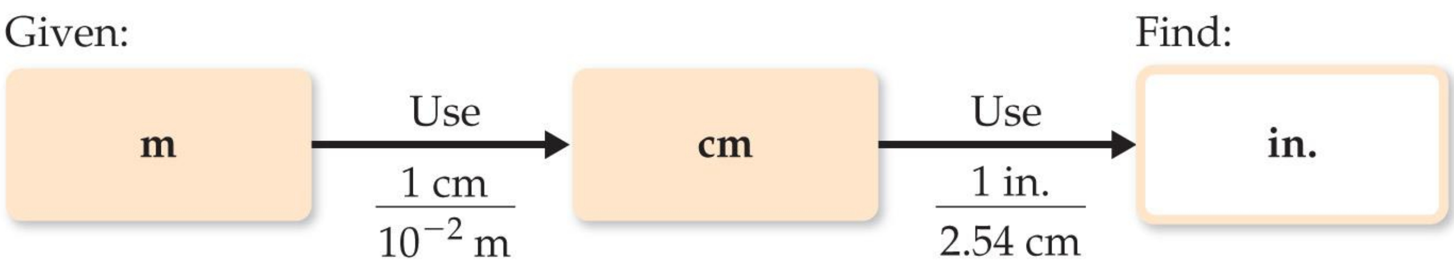
Average value is a measure of accuracy – how close to the true value



Dimensional Analysis

- We use **dimensional analysis** to convert one quantity to another.
- Most commonly, dimensional analysis utilizes **conversion factors** (e.g., 1 in. = 2.54 cm)

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



Dimensional Analysis

Use the form of the conversion factor that puts the sought-for unit in the numerator:

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

Conversion factor

Dimensional Analysis

For example, to convert 8.00 m to inches,

– convert m to cm, 1 m = 100 cm

– convert cm to in, 1 in = 2.54 cm

$$8.00\ \cancel{m} \times \frac{100\ \cancel{cm}}{1\ \cancel{m}} \times \frac{1\ \text{in.}}{2.54\ \cancel{cm}} = 315\ \text{in.}$$

Dimensional Analysis

Example: A woman has a mass of 115 lb, what is her mass in kg? Given the following:

1 lb = 453.6 g   1 kg = 10^3g

Solution

$$\text{mass in kg} = 115\ \cancel{\text{lb}} \times \frac{453.6\ \cancel{\text{g}}}{1\ \cancel{\text{lb}}} \times \frac{1\ \text{kg}}{10^3\ \cancel{\text{g}}}$$

= <u>52.2 kg</u>

Dimensional Analysis

Example

Earth's oceans contain about 1.36 x 10^9 km^3 of water. Calculate the volume in litres.

Solution

First find the conversion factors.

1 L = 1 dm^3 = 10^{-3} m^3 (we used 1 dm = 10^{-1} m)

1 km = 10^3 m hence $(1\ km)^3 = (10^3 m)^3 = 10^9\ m^3$

1 km^3 = 10^9 m^3

Volume in litres = $1.36x10^9 km^3 x \frac{10^9 m^3}{1km^3} x \frac{1L}{10^{-3} m^3}$

= 1.36 x10^{21} L