Chemical Principles 50:160:115

Fall 2016

Chemistry is *easy*…IF:	don't fall behind
	understand, not just memorize
	do problems
	remember things from one chapter to the next
Proficient in:	Explanations at the atomic/molecular level
	Qualitative explanations (illustrations / examples / words)
	Quantitative explanations using math
	Simple (algebraic) calculations

Notes are available at: **http://roche.camden.rutgers.edu/**

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Chapter 1: Matter and Measurement (Ch1 Chang, Ch1 Jespersen)

Chemistry is: The Science of Transformation.

Explanations at the Atomic and Molecular Level.

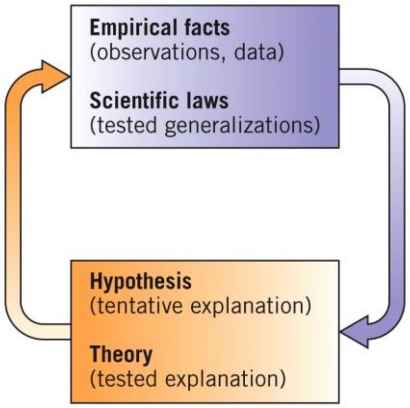
The Molecular Science.

The Scientific Method

The scientific method is a way to ask and answer questions by making *observations* and doing *experiments*.

The steps of the scientific method are to:

- · Ask a question
- · Do background research (find out what is *known*)
- · Construct a hypothesis (*tentative explanation*)
- · **Test** your hypothesis by doing an experiment
- · Analyze your data and draw a **conclusion**
- · *Communicate* your results



It is important for your experiment to be a *fair test*. A "*fair test*" occurs when you change only **one** factor (variable) and keep all other conditions the same.

The *cyclic* process means science is always *evolving*.

Facts, Laws, Hypotheses and Theories

Fact: *Observations* about the world around us. E.g. “It’s bright outside.”

Hypothesis: A *proposed explanation* for a phenomenon made as a starting point for further investigation.

E.g. “It’s bright outside because it is likely that the sun is out.”

Theory: A *well-substantiated explanation* acquired through the scientific method and repeatedly *tested* and *confirmed* through *observation* and *experimentation*.

E.g. “When the sun is out, it makes it bright outside.”

Law: A *statement* based on repeated *experimental observations* that describes some phenomenon of nature. Proof that something happens and **how** it happens, but <u>not</u> *why* it happens.

E.g. **Law** of Conservation of Energy, states that energy cannot be created or destroyed in an isolated system.

Chemistry requires *precision*.

You will be learning a new *language* - scientific terms have very precise meanings; attention to detail is very important.

Classification of Matter

Matter – anything that *occupies space* and has *mass*.

Substance – a form of *matter* that has a *definite* composition and *distinct* properties.

Mixture – a combination of two or more *substances* in which the substances retain their distinct identities.

Homogenous mixture – is a mixture where the composition is the *same* throughout (e.g. air, diet coke).

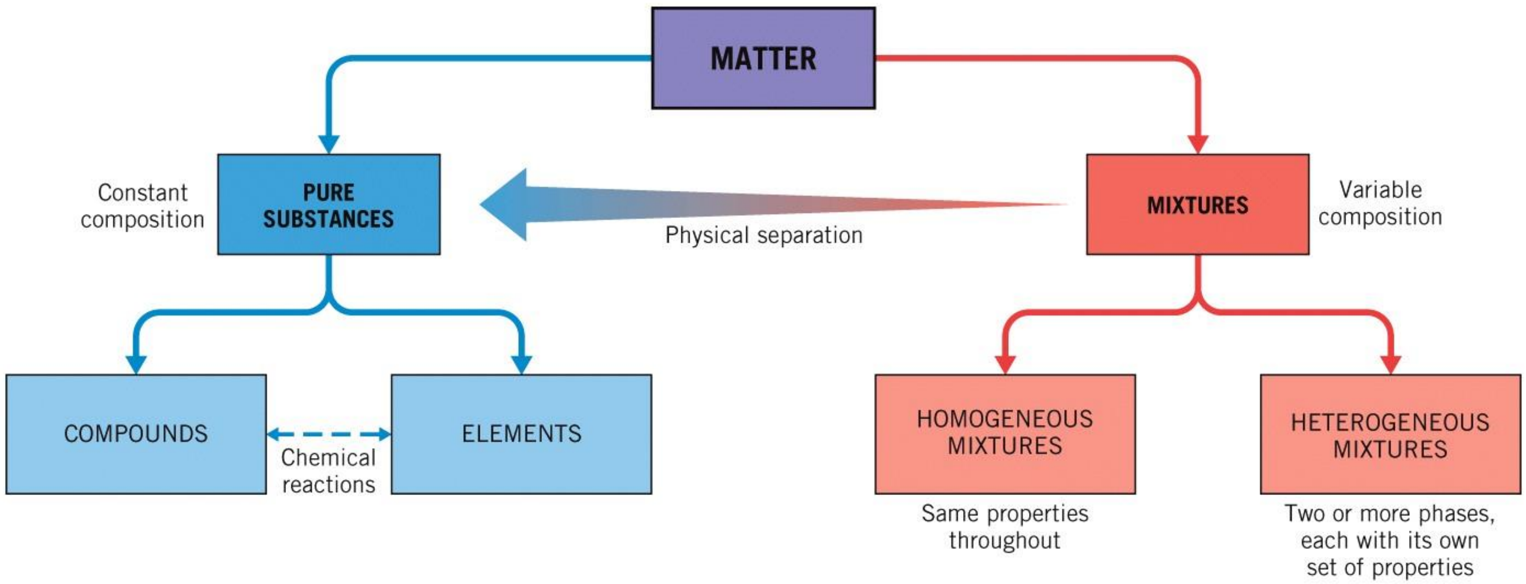
Heterogeneous mixture – is a mixture where the composition is **not** uniform throughout (e.g. diet coke with ice).

Element – a *substance* that **cannot** be separated into *simpler* substances by chemical means.

The elements are named, and often referred to by their chemical symbols: e.g. **H**, **O**, **N**, **C**, **Li**, **Cl**, **Br**, **Hg**, **Au**.

Compound – a *substance* composed of atoms of two or more *elements* united in fixed proportions.

Relationship between Elements, Compounds, and Substances

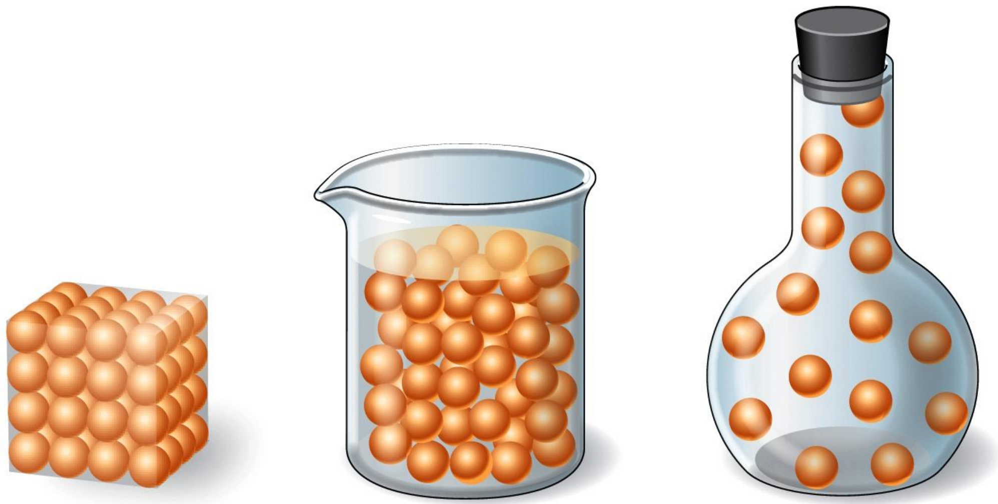


Air is a (homogeneous) mixture.

Physically separate it into pure substances (N_2, CO_2, O_2, etc.).

Chemically separate the compounds into elements ($CO_2 \rightarrow C$ and O_2).

Three *States* of Matter



Solids have fixed shape and volume.

Particles are close together.

Liquids have fixed volume, but adopt the container's shape.

Particles are close together.

Gases expand to fill the container.

Particles are far apart.

Solid → Liquid is *melting*, and Liquid → Gas is *boiling*.

Properties

Extensive Properties depend on the *amount* of *matter*.

For example: Mass (the *quantity of matter in a given sample of substance*).

Volume of a substance.

Intensive Properties do **not** depend on the amount of matter.

For example: Density.

Melting or Boiling Points.

Physical Properties can be measured (and observed) *without* changing the composition or identity of a substance.

E.g. Temperature, melting point, color, amount, odor, solubility, electrical conductivity.

Chemical Properties require that a chemical change occur to observe the property.

E.g. Iron will rust, coal is combustible, silver jewelry can tarnish.

Physical Change means the matter is *rearranged*, but the internal structure is **not** affected.

For example: Boiling water (water molecules are forced away from each other when the liquid changes to vapor, but the molecules are still H_2O).

Dissolving sugar in water (sugar molecules are dispersed within the water, but the individual sugar molecules are unchanged).

Dicing potatoes (cutting usually separates molecules without changing them).

Chemical Change requires the formation of **new** chemical substances.

For example: Iron rusting (iron oxide forms).

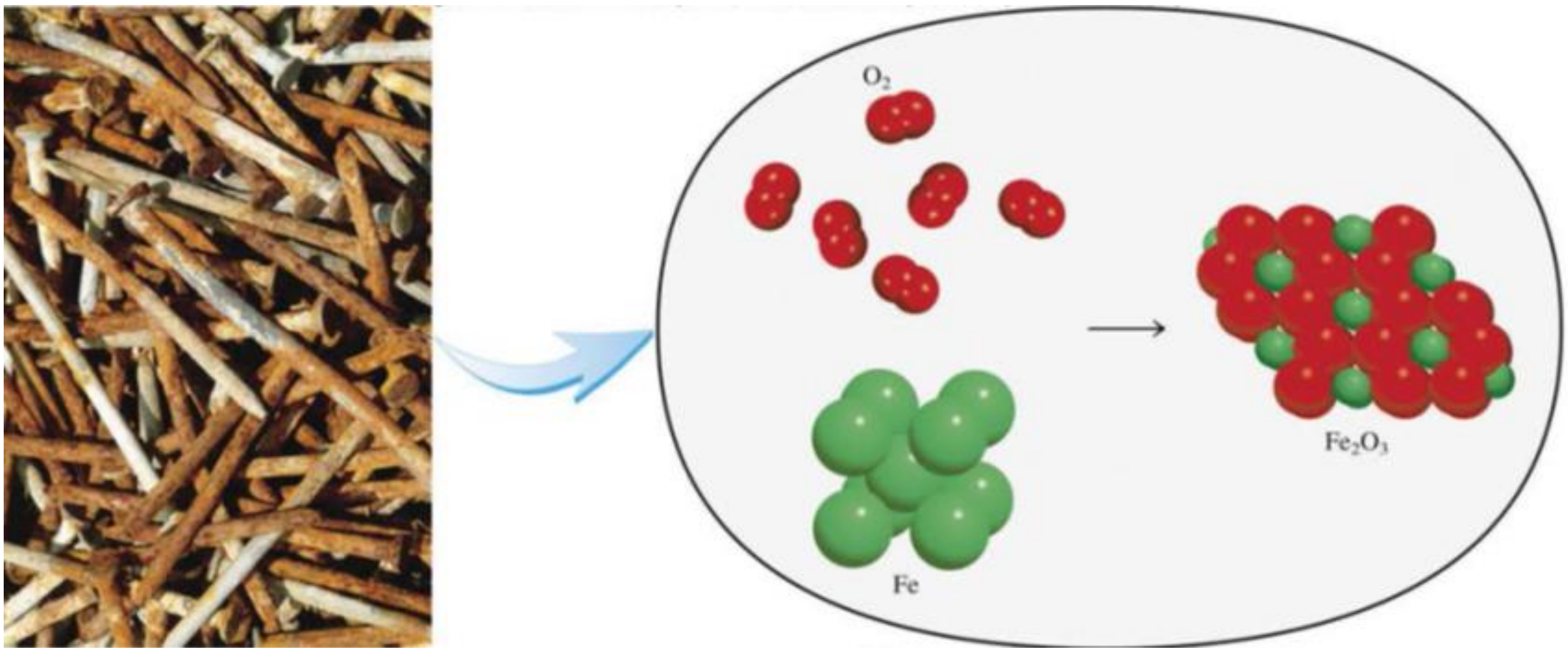
Gasoline burning (water vapor and carbon dioxide form).

Eggs cooking (fluid protein molecules uncoil and crosslink to form a network).

Suntanning (vitamin D and melanin is produced).

A *chemical change* is also known as a Chemical Reaction.

Iron nails rusting is a *chemical process* (iron converted to iron oxide), therefore it is a *chemical property* of iron that it will rust.



Macroscopic refers to something we can *directly observe* (with the naked eye).

Microscopic refers to *indirect observation*, where a tool/aid/instrument is needed to assist in the observation.

Units – we need *something* as a reference or standard (for meaningful comparisons and *communication*).

Do I weigh 13.5 stone, or 189 lbs or 85.7 kg ?

International System of Units (SI Units)

There are 7 base units, we will use 6 of them (*luminous intensity in candela, cd, is not covered*).

Physical quantity	**Name of Unit**	**Abbreviation**
Mass	Kilogram	kg
Length	Meter	m
Temperature	Kelvin	K
Amount of Substance	Mole	mol
Time	Second	s
Electric Current	Ampere	A

In science we deal with both very *large* and very *small* numbers, so we use prefixes *before* our units.

Prefix	**Abbreviation**	**Meaning**	**Example**
Giga	G	10^{9}	Gigabyte
Mega	M	10^{6}	Megawatt
Kilo	k	10^{3}	Kilogram
Deci	d	10^{-1}	Deciliter
Centi	c	10^{-2}	Centimeter
Milli	m	10^{-3}	Millimeter
Micro	µ	10^{-6}	Micrometer
Nano	n	10^{-9}	Nanometer
Pico	p	10^{-12}	Picomole

There are also common *non-SI units* often used in Chemistry:

Measurement	Unit	Abbreviation	Value in SI Units
Length	angstrom	Å	$1\ \text{Å} = 0.1\ \text{nm} = 10^{-10}\ \text{m}$
Mass	atomic mass unit	u (amu)	$1\ \text{u} = 1.66054 \times 10^{-27}\ \text{kg}$ (rounded to six digits)
	metric ton	t	$1\ \text{t} = 10^3\ \text{kg}$
Time	minute	min.	$1\ \text{min.} = 60\ \text{s}$
	hour	h	$1\ \text{h} = 60\ \text{min.} = 3600\ \text{s}$
Temperature	degree Celsius	°C	$T_K = t_{°C} + 273.15$
Volume	liter	L	$1\ \text{L} = 1000\ \text{cm}^3$

Or in everyday language:

Measurement	English Unit	English/SI Equality[a]
Length	inch	1 in. = 2.54 cm
	yard	1 yd = 0.9144 m
	mile	1 mi = 1.609 km
Mass	pound	1 lb = 453.6 g
	ounce (mass)	1 oz = 28.35 g
Volume	gallon	1 gal = 3.785 L
	quart	1 qt = 946.4 mL
	ounce (fluid)	1 oz = 29.6 mL

Measuring Temperature

Fahrenheit – common in everyday use (e.g. weather, body temperature), but **not** in science.

Water freezes at 32 ^{o}F, and boils at 212 ^{o}F.

Celsius scale – commonly used in science (even though it is a *non-SI unit*).

Based on freezing point (0 °C) and boiling point (100 °C) of water.

Kelvin scale – based on absolute zero being 0 K, (*lowest possible temperature*), which is –273.15 °C.

Water freezes at 273.15 K and boils at 373.15 K.

Notice the Celsius and Kelvin have *equal size units*, but differ at their zero (start) point.

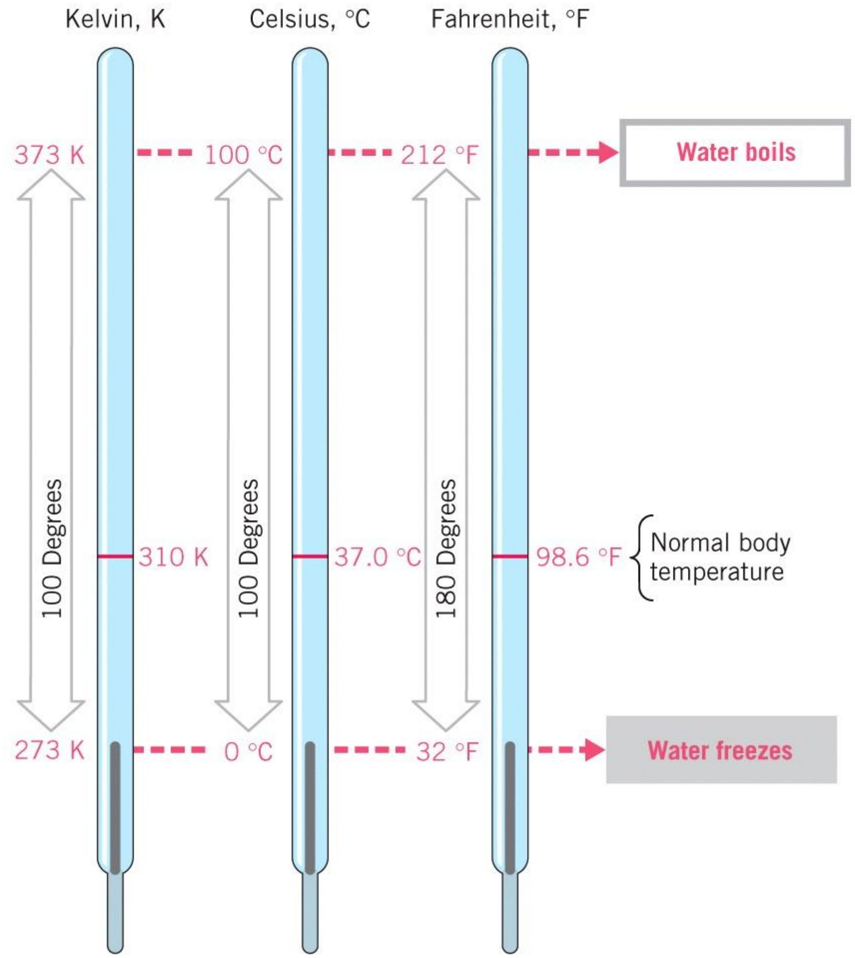
Conversions

$$K = °C + 273.15$$

$$°C = \frac{5}{9}(°F - 32)$$

$$°F = \frac{9}{5}°C + 32$$

Notice for Kelvin and Celsius, ΔT is the *same* numerical value (same size units).

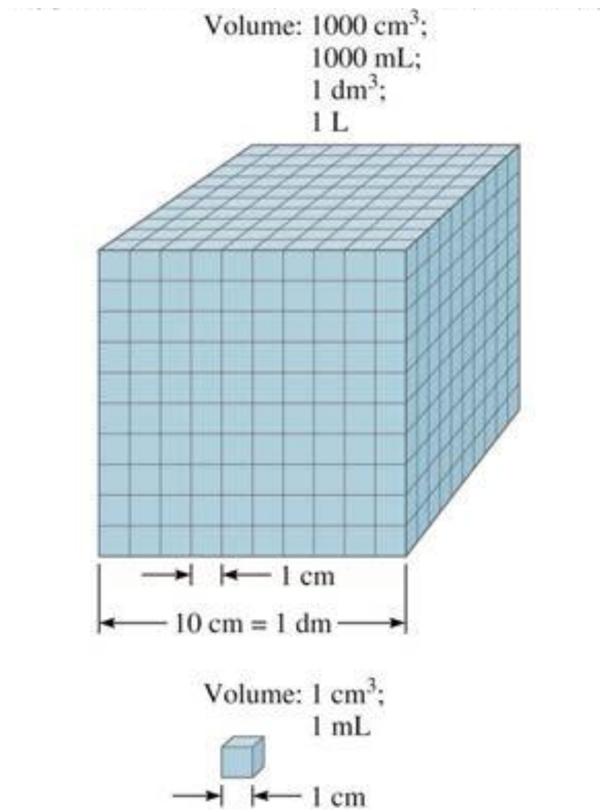


*(So a 100 K increase is the same as 100 ^oC increase – but **NOT** the same as a 100 ^oF increase).*

Derived Units: Measuring Volume

A derived unit is a unit comprised of a *combination* of the seven base units.

Length is an SI unit, so area (two lengths) and volume (three lengths) are *derived* units.



Typically we use cm^3 for the volume of solid items, and mL for liquids and gases.

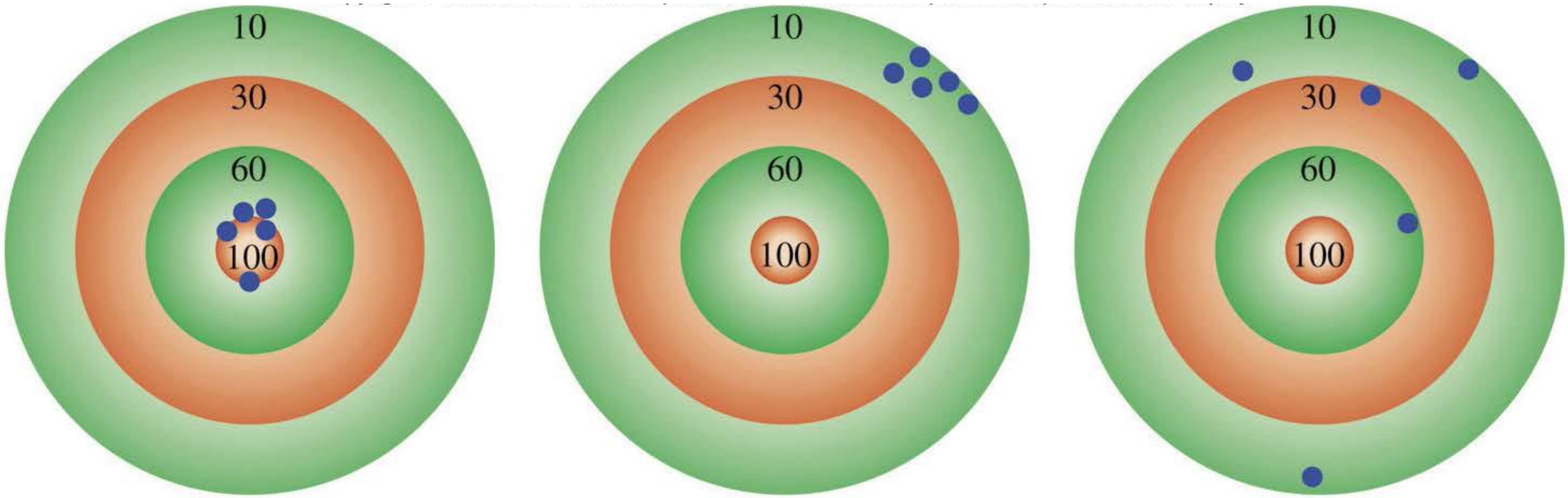
Derived Units: Measuring Density

Density is *mass* divided by *volume*, and therefore a derived unit.

Substance	Density (g/cm^3 or g/mL)
Air	0.0012
Gasoline	0.7
Water	1.0
Aluminum	2.8
Iron	7.87
Silver	10.50
Lead	11.34
Mercury	13.53
Gold	19.28
Osmium	22.59

Accuracy, Precision, and Significant Figures in Measurement

To be able to obtain our numbers or values, we need to *measure* them.



(a) (b) (c)

Accuracy – is the *closeness* to an accepted or known value.

Precision – is the *reproducibility* of results.

So (a) is accurate and precise, (b) is precise but not very accurate, (c) neither accurate nor precise.

Measured Numbers vs. Exact Numbers

Scientific *measurements* are reported so that every digit is *certain* except the last one, which is *estimated*.

So if the smallest division on a ruler is 0.1 cm, then your measurements will be to the 0.01 cm.

The first decimal place is *certain*, the second is *uncertain*.

This screw is measured as 5.10 cm long.

For our 5.10 cm value, the 1 is certain, the 0 is uncertain. It is obtained by making an *estimate*.

This number has three significant figures.

It is scientific convention that:

All digits in *measurement* up to and including the first estimated digit are **significant**.

*(*Alternatively, significant digits are the number of *certain digits plus the first uncertain digit*, or,

digits in measurement from the first non-zero number on ***left*** *to the first estimated digit on* ***right****).*

Exact numbers are numbers that have **no** uncertainty.

They come from: Counting of discrete objects. E.g. 3 eggs.

Defined quantities: 1 lb = 16 oz.

Integral numbers that are part of an equation: circumference of a circle = $2\pi r$.

(The 2 is an exact number).

We assume they have *infinite* number of significant figures, meaning they do not limit the number of significant figures in a calculation.

Guidelines to significant figures in a measured quantity

1. Nonzero digits are always significant.

   457 cm (3 sig. figs) 2.5 g (2 sig. figs)

2. Zeros between non-zero digits are always significant.

   1005 kg (4 sig. figs) 1.03 cm (3 sig. figs)

3. Zeros at the *beginning* of a number are **never** significant; they merely indicate the position of the decimal point.

   0.02 g (1 sig. fig) 0.0026 cm (2 sig. figs)

4. Zeros that fall **both** at the *end* of a number **and** *after* a decimal point are always significant.

   0.0200 g (3 sig. figs) 3.0 cm (2 sig. figs)

5. When a number *ends* in zero but contains no decimal point, the zeros *may* or *may not* be significant.

   130 cm (2 or 3 sig. figs) 10,300 g (3, 4, or 5 sig. figs)

The use of exponential (*scientific*) notation can remove the possible ambiguity seen in the last example…

Scientific Notation is the clearest way to present the number of significant figures *unambiguously*.

We use a number between **1** and **10** followed by the appropriate power of 10.

Using exponential notation, 10300 g can be written:

1.03×10^4 g (three significant figures)

1.030×10^4 g (four significant figures)

1.0300×10^4 g (five significant figures)

(Alternatively, you can put a decimal point after the zeros to indicate they are measured.

E.g. ***700.*** *indicates 3 sig figs).*

Rounding Numbers: Significant Figures in Calculations

In multiplication and division the result must be reported with the **same** number of significant figures as the measurement with the *fewest* significant figures.

1. If the leftmost digit to be removed is less than five, the *preceding* number is left *unchanged*.

    Rounding 7.248 to 2 sig. figs gives 7.2

2. If the leftmost digit to be removed is five or greater, the *preceding* number is *increased* by one.

    Rounding 4.735 to three significant figures gives 4.74

    Rounding 2.376 to two significant figures gives 2.4

In addition and subtraction the result **cannot** have more digits to the *right* of the decimal point than any of the original numbers.

Or put another way, the answer must have the *same* number of *decimal places* as the quantity with the fewest number of decimal places.

E.g.

$$\begin{array}{r} 20.4 \\ 1.322 \\ 83 \\ \hline 104.722 \end{array}$$

The sum of these numbers is correctly reported as rounded off to **105**

Converting from One Unit to Another

This is achieved via the Factor-Label Method of Solving Problems (also called Dimensional analysis method).

Given Quantity x Conversion Factor = Desired Quantity

This uses *conversion factors* usually found in lists inside the back cover of book/handouts etc.

E.g. 1 in. = 2.54 cm (exactly)

1 mile = 1.609 km

1 lb = 453.6 g

Problem: how tall in cm is AJR if his height is 75.0 inches ?

$$75\ \cancel{\text{in}} \times \frac{2.54\ \text{cm}}{1\ \cancel{\text{in}}} = 190.5\ \text{cm}$$

$$= 191\ \text{cm to 3 sig figs}$$

The inch units cancel, leaving the answer in cm.

Be aware of units *raised* to a *power*:

1 in = 2.54 cm

But $(1\ \text{in})^3 = (2.54\ \text{cm})^3$

Which is $1^3\ \text{in}^3 = 2.54^3\ \text{cm}^3$ (= 16.387064)

Therefore $1\ \text{in}^3 = 16.4\ \text{cm}^3$ (3 sig figs)

Summary of Dimensional Analysis

In using dimensional analysis to solve problems, we will always ask three questions:

1. What *data* are we given in the problem?
2. What *quantity* do we wish to obtain in the problem?
3. What *conversion factors* do we have available to take us from the given quantity to the desired one?

Problem: Convert 144 in^3 to mL.

$$144\ \cancel{in^3} \times \left(\frac{2.54\ \cancel{cm}}{1\ \cancel{in}}\right)^3 \times \frac{1\ mL}{1\ \cancel{cm^3}} = 2359.737216\ mL$$

$$= 2360\ mL\ (3\ sig\ figs)$$

$$= 2.36 \times 10^3\ mL$$

Problem: What is the volume (in cm^3) of a 63.4 g piece of metal with a density of 12.86 g/cm^3?

$$\text{Density} = \frac{\text{Mass}}{\text{Volume}}$$

Rearrange to get $$\text{Volume} = \frac{\text{Mass}}{\text{Density}}$$

$$= 63.4\ \cancel{g} \times \frac{cm^3}{12.86\ \cancel{g}} = 4.93\ cm^3\ (3\ sig\ figs)$$