Chapter 1

Matter and Measurement

I) Definition of Chemistry

Science which deals w. composition, structure and reactions of matter.

A) Matter

Anything that has mass
& occupies space.

1) Mass

measure of the quantity of matter

2) Weight

Result of gravitational attraction between matter

B) Composition

What matter is made of and how much of each component is present.

1) Several Ways of Expressing

a) by weight (mass)

b) by volume

c) Percent

d) Number of Moles

e) Number of Atoms

2) Macroscopic Level

Amounts that can be seen and weighed

a) Ex: 1/4 lb. cheeseburger

1) By weight (mass)

meat	4.0 oz
cheese	0.8 oz
roll	1.7 oz
	6.5 oz

b) Ex : 95% ethanol

95% ethanol & 5% water

3) Submicroscopic Level

described by numbers & types of atoms

Atoms: simple units of matter

Molecules: combinations of atoms

a) Qualitative

Ethanol consists of carbon, hydrogen & oxygen

b) Quantitative

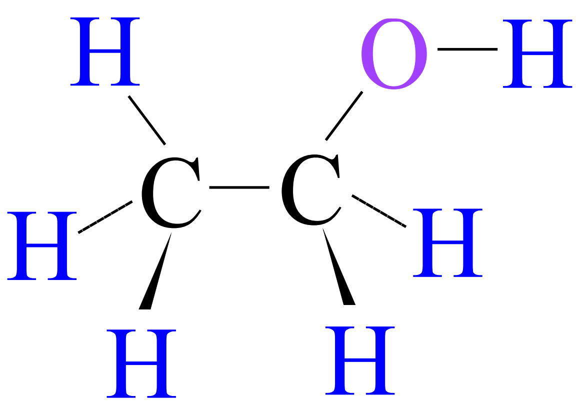
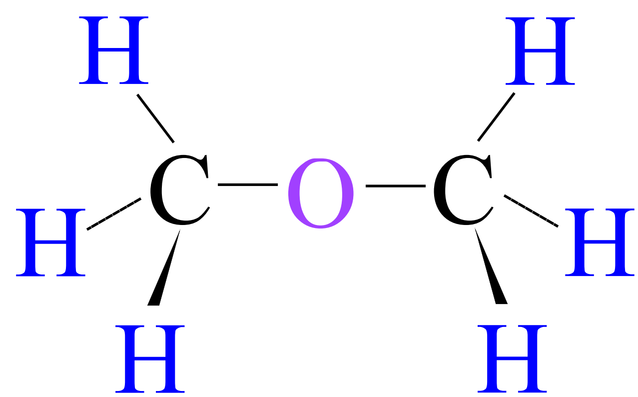
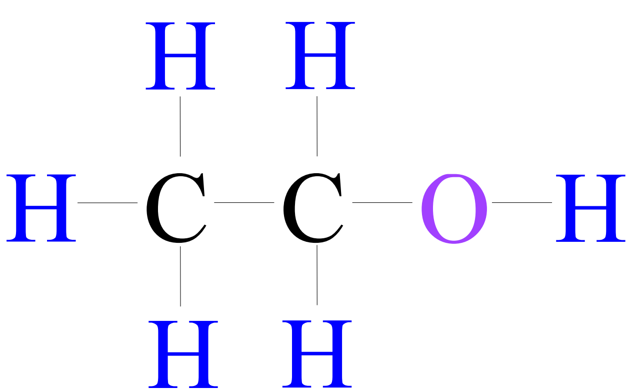
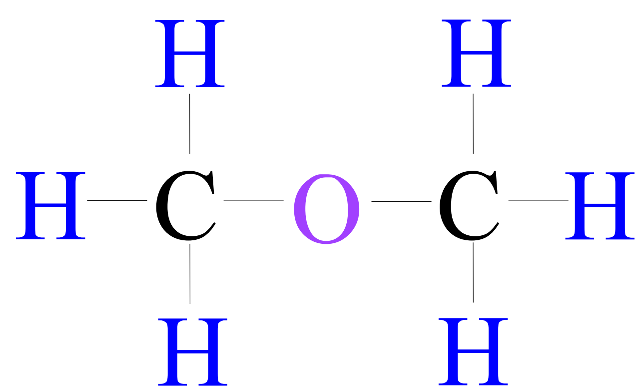
Ethanol: 2 C atoms, 6 H atoms 1 O atom

Formula: C_2H_6O

C) Structure

Arrangement of components & how they are held together, or bonded

Ethanol	Dimethyl Ether
C_2H_6O	C_2H_6O
or	or
C_2H_5OH	CH_3OCH_3

D) Reactions

Changes in composition
& structure.

1) What products are formed?

2) How much of each product?

3) How fast the change occurs?

4) What energy changes
accompany the reaction?

$$2\,H_2 + O_2 \longrightarrow 2\,H_2O + \text{heat}$$

II) Scientific Method

A) Experiment (Record Observations)

1) Careful recordings & analysis of data under controlled conditions

2) Reproducible - exp. never performed just once

B) Draw a Conclusion - Law

Concise statement about a basic relationship or regularity of nature drawn from observations.

- true for all cases examined

Law of Gravity $$F = G\frac{m_1 m_2}{r^2}$$

C) Model (Explanation)

Idea that explains or correlates a number of facts

- explains how and why

1) Hypothesis

Tentative model

- test with new experiments

2) Theory

Model that has been tested many times & not disproved

- best idea that agrees with all known facts.

III) States of Matter

Gas	Liquid	Solid
No definite volume or shape	Constant volume	Definite volume
fills container & takes its shape	shape of container	Definite shape
Highly compressible	Slightly compressible	Incompressible
Great expansion when heated	expands slightly when heated	expands very slightly when heated

IV) Physical and Chemical Properties

A) Physical Property

can be determined ***WITHOUT*** changing the identity of the substance.

Ex : physical state, color, odor, m.p., b.p., density, specific heat

B) Chemical Property

describes a reaction with or conversion into another substance

Ex : flammability

C) Extensive & Intensive Prop.

1) Extensive Property

Depends on sample size.

Ex : mass, volume, heat content

2) Intensive Property

Do NOT depend on sample size.

Ex : color, melting point,
boiling point, density,
specific heat

V) Physical & Chemical Changes

A) Physical Changes

Change in appearance without change in identity

1) Ex: change in state

	melting → / ← freezing	
Solid	⇌	Liquid

	vaporization → / ← condensation	
Liquid	⇌	Gas

	sublimation → / ← deposition	
Solid	⇌	Gas

B) Chemical Changes (Reactions)

Converts a substance into a chemically different substance.

- change in composition &/or structure

$$2\ K(s) + 2\ H_2O(\ell) \longrightarrow 2\ KOH(aq) + H_2(g)$$

VI) Pure Substances and Mixtures

A) Pure Substances

uniform in properties throughout

1) Characteristics

a) constant (fixed) composition

b) distinct intensive properties

c) NOT separable by
physical methods

Elements and Compounds

2) Elements

Substances that can NOT be decomposed into simpler substances by chemical means

118 known elements

Symbols used to identify

- 1 or 2 letters

C ≡ carbon

Co ≡ cobalt

Ca ≡ calcium

a) Periodic Table

Elements arranged in order of increasing atomic number

- properties of elements correlate w. position in periodic table

1) Periods

horizontal rows

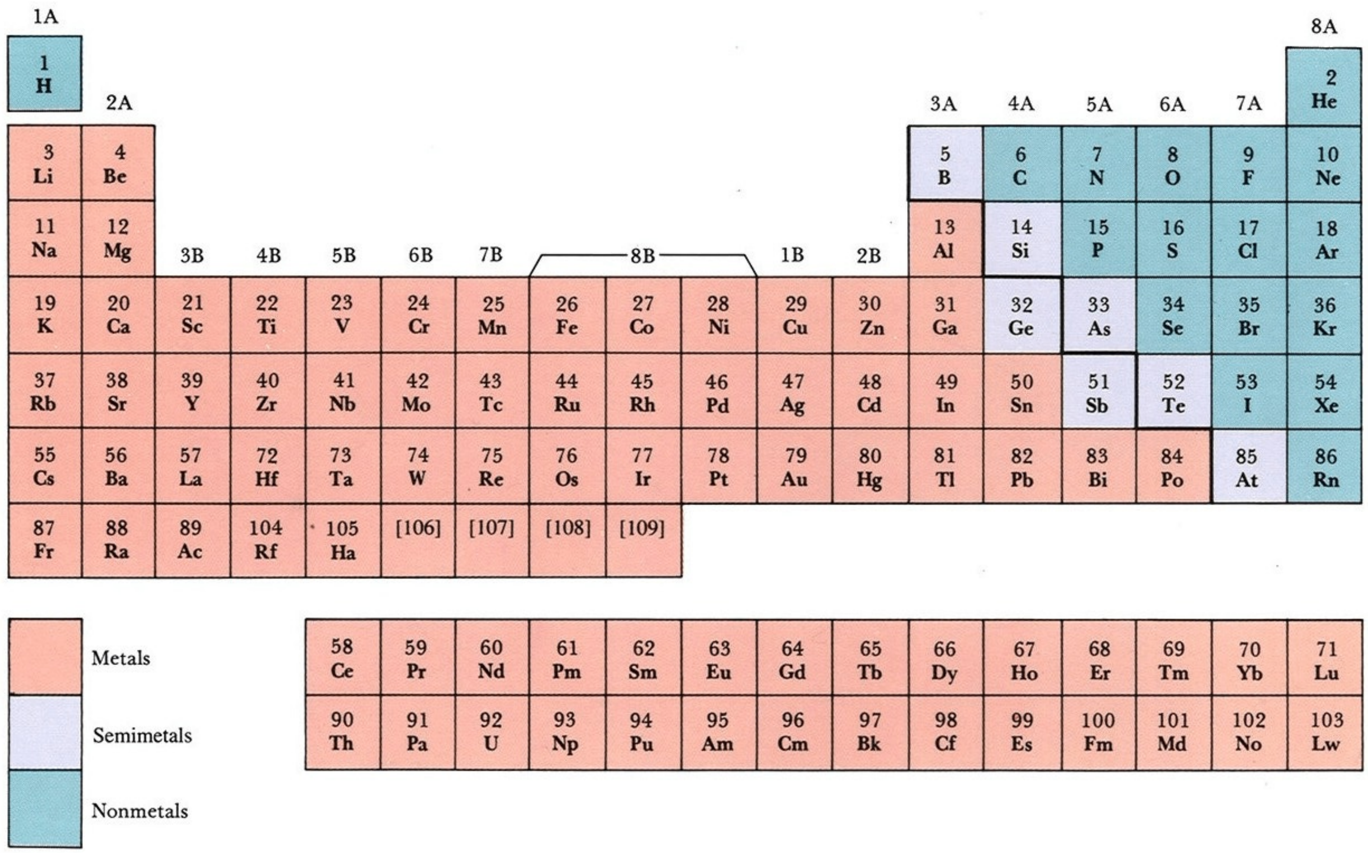
- gives information about atomic structure

2) Groups

vertical columns

- elements in groups have similar physical & chemical properties

Transparency 13 Figure 2.16 Periodic table divided into metals, nonmetals, and semimetals



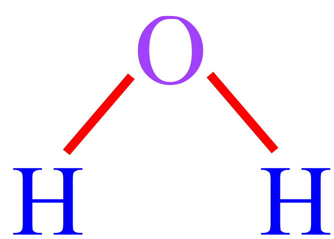
CHEMISTRY: THE CENTRAL SCIENCE
by Brown/Le May/Bursten

3) Compounds

Composed of 2 or more elements, chemically combined

- separable into its elements by chemical means

Ex : H_2O



11.2% hydrogen
88.8% oxygen

a) Law of Definite Proportions

elements in a compound are combined in definite proportions by mass

B) Mixtures

2 or more substances NOT chemically combined.

1) Characteristics

a) variable composition

b) separable by physical methods

c) components retain their own properties (chem. identities)

Ex: water-ethanol mixture

5% - mostly water
95% - mostly ethanol
50% - equal amounts

2) Heterogenous Mixture

Consists of parts that are unlike

- do NOT have same composition, properties & appearance throughout

Ex: sand & salt
Raisin Bread

3) Homogenous Mixture

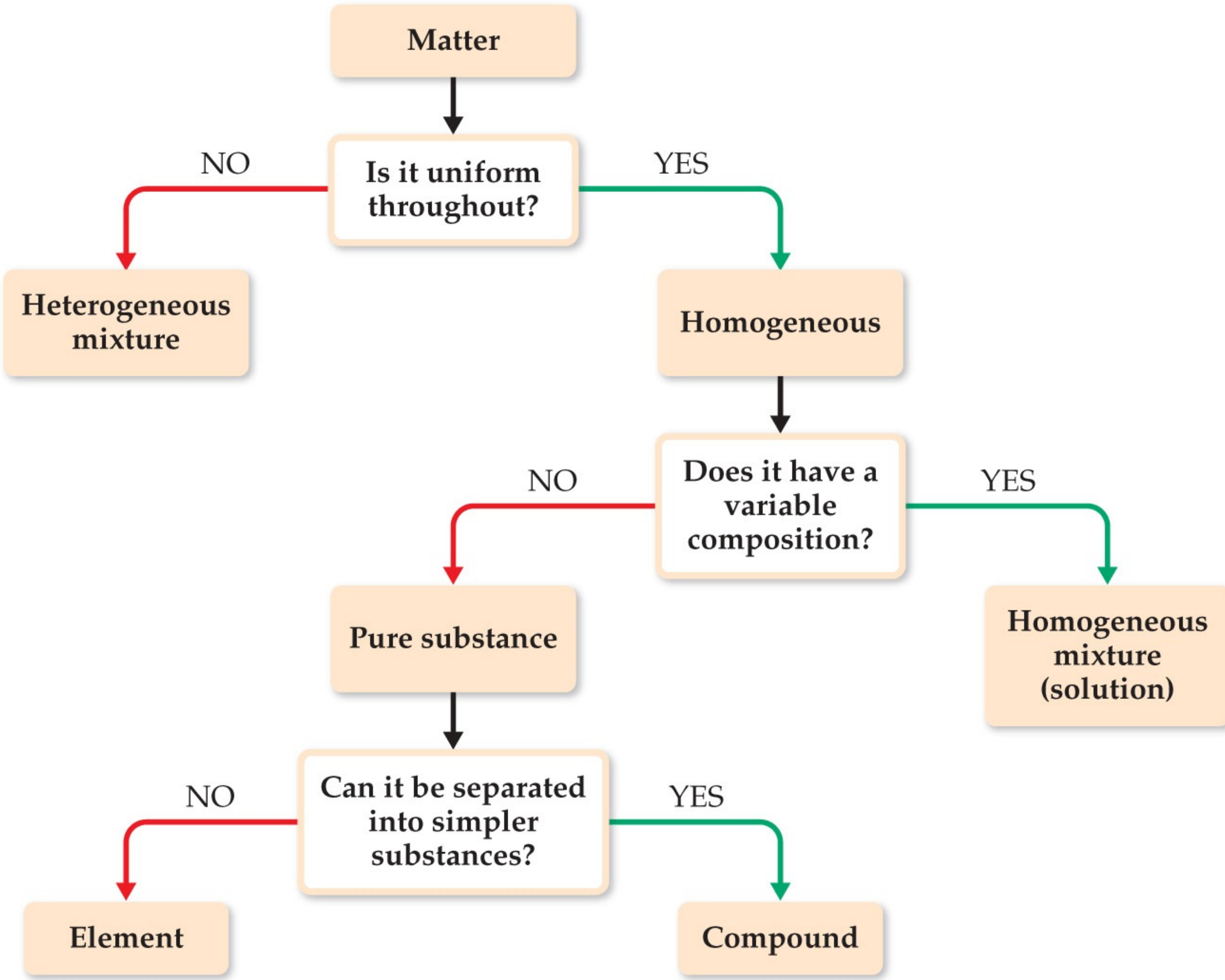
Prop. are uniform throughout

- down to the molecular level

Solutions

a) Ex:

gaseous solution: Air
liquid soln: 95% ethanol
solid solution: brass


Matter
Is it uniform throughout?
NO
YES
Heterogeneous mixture
Homogeneous
Does it have a variable composition?
NO
YES
Pure substance
Homogeneous mixture (solution)
Can it be separated into simpler substances?
NO
YES
Element
Compound

VII) Nature of Energy

Energy ≡ Capacity to do work

Mechanical Work

$w = F \times d$

Heat energy

- energy used to cause the temperature of an object to inc.

A) Kinetic & Potential Energy

1) Kinetic Energy

$$KE = \frac{1}{2} m v^2$$

Energy due to motion

SI units:

$$\text{Energy} = \text{kg}\ (\text{m/s})^2 = \text{J}$$

2) Potential Energy

Energy stored in an object by virtue of its position or composition

Chemical energy is due to composition of substances

Chemical energy released when bonds are formed.

Chemical energy is consumed when bonds are broken.

VIII) Units of Measurement

International System, SI units:

- have base units from which all other units are derived

Table 1.4

mass	length	time	temp
kg	m	s	K

Base units for length & mass are part of metric system

- employs factors of 10

Prefixes: indicate size of unit relative to base unit

Selected SI Prefixes

Prefix	Abbrev.	Meaning	Example
Mega-	M	10^6	1 megameter (Mm) = 1 x 10^6 m
Kilo-	k	10^3	1 kilometer (km) = 1 x 10^3 m
Deci-	d	10^{-1}	1 decimeter (dm) = 0.1 m
Centi-	c	10^{-2}	1 centimeter (cm) = 0.01 m
Milli-	m	10^{-3}	1 millimeter (mm) = 0.001 m
Micro-	μ[a]	10^{-6}	1 micrometer (μm) = 1 x 10^{-6} m
Nano-	n	10^{-9}	1 nanometer (nm) = 1 x 10^{-9} m
Pico-	p	10^{-12}	1 picometer (pm) = 1 x 10^{-12} m
Femto-	f	10^{-15}	1 femtometer (fm) = 1 x 10^{-15} m

[a] This is the Greek letter Mu (pronounced "mew")

A) Mass

kilogram, kg

$1\ \text{kg} \equiv 10^3\ \text{g}$

$1\ \text{kg} \cong 2.205\ \text{lb}$

$1\ \text{lb} \cong 453.6\ \text{g}$

B) Length

meter, m

$1\ \text{in} \equiv 2.54\ \text{cm}$

$1\ \text{m} \cong 1.0936\ \text{yd}$

C) Volume

SI unit is m^3

Commonly use liter, L

$$1\ L \equiv 1\ dm^3$$

$$(1\ dm \equiv 10\ cm)$$

$$1\ L = (10\ cm)^3 = 10^3\ cm^3$$

$$1\ L \equiv 10^3\ mL$$

$$\therefore\ 1\ mL = 1\ cm^3$$

D) Temperature

Must specify temp. when making quantitative measurements

1) Celsius Scale

°C - commonly used

Fahrenheit, °F, scale used in public (USA)

°F	°C	
212	100.0	b.p. of H_2O
98.6	37.0	body temperature
32.0	0.0	f.p. of H_2O

$$y\ °C = \frac{100\ °C}{180\ °F}(x\ °F - 32\ °F)$$

$$y\ °C = \frac{5\ °C}{9\ °F}(x\ °F - 32\ °F)$$

or

$$y\ °F = \frac{9\ °F}{5\ °C}(x\ °C) + 32\ °F$$

a) Ex : Convert 25 °C to °F

2) Kelvin Scale

SI base unit is kelvin, K

Must be used in most cases in chemistry

Absolute scale:

0 K : lowest possible temp.

$\Delta T_K = \Delta T_{^\circ C}$ (unit same size)

0 °C = 273.15 K

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

E) Density

Mass per unit volume

$$D = \frac{m}{V}$$

SI unit is kg/m^3

Solids	Liquids	Gases
g/cm^3	g/mL	g/L

1) Specific Gravity

$$\text{Sp. Gr.} = \frac{D_{substance}\ (g/mL)}{D_{water}\ (g/mL)}$$

No units

H_2O : $D = 1.0$ g/mL

Ethanol : $D = 0.79$ g/mL

sp. gr. = 0.79

F) Units of Energy

$w = F \times d$

$= (m \times a) \times d$

$= (kg \times m/s^2) \times m$

⇓

$= (kg \bullet m^2)/s^2 \quad = N \times m$

= joule, J (SI unit)

calorie (cal)

original def: amt. of energy req. to raise temp. of 1g of water by 1 °C, from 14.5 °C to 15.5 °C

1 cal = 4.184 J

Cal - nutritional calorie

1 kcal

1) Kinetic Energy

$$KE = \frac{1}{2} m v^2$$

SI units:

$$\text{Energy} = \text{kg}\,(\text{m/s})^2 = \text{J}$$

IX) Measurement & Significant Figures

Uncertainties always exist
in measured quantities.

A) Precision

Degree of reproducibility of
repeated measurements

i.e. - How close are to each other

Depends on skill of measurer

1) Ex: Measure width of
notebook paper (in cm)

21.32 21.33 21.32 21.31

avg. width = 21.32 cm

good precision

B) Accuracy

How close measurement is to true value

Paper's true width is 21.59 cm

Numbers in previous ex. have poor accuracy

Depends on quality of the measuring device

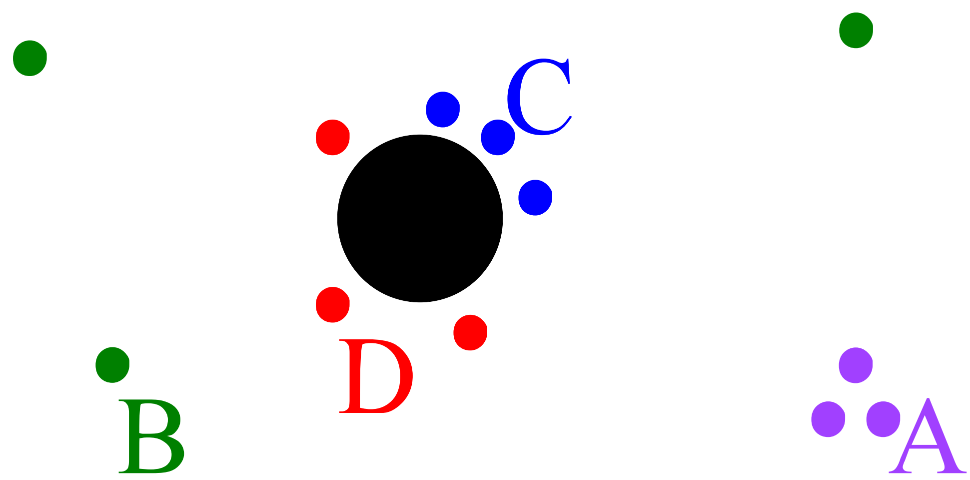
1) Ex: remeasure paper with a "better" ruler (in cm)

21.54	21.61	21.56	21.65

Avg. = 21.59 cm

good accuracy, poor precision

Ex:



A (•) - good precision
poor accuracy

B (•) - poor precision
poor accuracy

C (•) - good precision
good accuracy

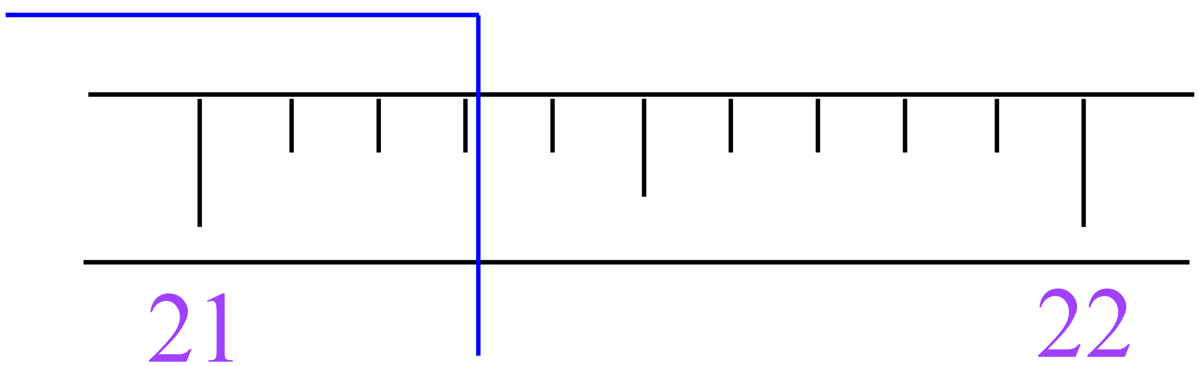
D (•) - "poor" precision
good accuracy

C) Significant Figures

ALL digits we know **exactly**
plus **one** we **estimate**.

Calibration of instrument determines number of significant figures (sig. fig.)

- previous measurements used a ruler marked in tenths of a cm (mm)



D) Exact Numbers

Infinite number of sig. fig.

1) By Count

Count the number of people in the room

- Integers

2) By Definition

1 dozen ≡ 12 items

1 yd ≡ 3 ft

1 lb ≡ 16 oz

1 in ≡ 2.54 cm

E) Significant Figures Rules

1) ALL nonzero digits ARE sig.

1,542 3.456

2) Captive zeros: zeros between sig. digits ARE sig.

20.6 20.06

3) Leading zeros: zeros to left of first nonzero digit are NOT sig.

- locate decimal point

0.401 0.004

4) Trailing zeros: zeros to right of last non-zero digit

a) Number ends in zero to right of decimal point

- zeros ARE sig.

0.040 400.0

b) Number ends in zero to left of decimal pt.

- zeros generally NOT sig.

400 4100

f) Scientific Notation

Express a number as a coefficient times a power of 10.

$$A \times 10^{n}$$

1 non-zero digit to left of decimal pt.

$400 = 4 \times 10^2$

4.0×10^2

4.00×10^2

Entering in calculators:

4 [EE] or [EXP] 2

F) Sig. Fig. in Calc. - Rounding Off

Result of a calc. must reflect accuracy of original measurements

1) Multiplication & Division

Answer must contain same # of sig. fig. as quantity w. least # of sig. fig.

a) Ex 1: Divide 907.2 by 453.6

b) Ex 2: Determine volume of a box that measures 3.6 cm by 2.45 cm by 10.0 cm.

1) Rounding Rule 1

If leftmost number to be discarded is < 5,

round down

i.e. - last number to be retained is unchanged

∴ Answer should be:

2) Addition & Subtraction

Last place in answer is last place common to ALL numbers

a) Ex 3: Add 4, 1.45, 12.4 & express answer to correct number of sig. fig.

```
  4
  1.45
 12.4
------
 17.85
```

1) Rounding Rule 2

If leftmost number to be discarded is > 5 or 5 followed by non-zero digits,

round up

i.e. last number retained is inc. by 1

b) Ex 4: Find the difference between 12.4 and 4

$$\begin{array}{r} 12.4 \\ -\ \ 4 \\ \hline 8.4 \end{array}$$

c) Ex 5: Add 9.8 and 9.94

$$\begin{array}{r} 9.8 \\ +\ 9.94 \\ \hline 19.74 \end{array}$$

d) Ex 6: Subtract 2.78 from 3.18

$$\begin{array}{r} 3.18 \\ -\ 2.78 \\ \hline 0.40 \end{array}$$

e) Ex 7: Find diff. between 12.3 & 1.45

$$\begin{array}{r} 12.3 \\ -\ 1.45 \\ \hline 10.85 \end{array}$$

1) Rounding Rule 3

If number to be discarded is 5, or 5 followed by zeros,

round even

i.e. - leave last digit to be retained unchanged if even, increase by 1 if it is odd

∴ Answer is:

f) Ex 8: Round each of the following to 2 sig. fig.

1.45 ⇛

1.550 ⇛

1.452 ⇛

X) Dimensional Analysis
(Factor Unit Method)

Solve problems by carrying units throughout the calculations

- just converting units by using conversion factors

Conversion Factor

A number having two or more units associated with it

Numerically equivalent to 1

information given in one type of unit X conv. factor = same info in a different type of unit

A) Ex 1: A local donut shop sells donuts for \$4.49 a dozen. You want 3 dozen donuts. How much will it cost?

change units

dozen ⇒ dollars

1 dozen ≡ \$4.49

Can write 2 conv. factors

$$\frac{1 \text{ dozen}}{\$4.49} = 1 \qquad \frac{\$4.49}{1 \text{ dozen}} = 1$$

Convert 3 dozen to ? dollars :

B) Ex 2: Convert 0.34 cm to µm

? cm = 1 µm

1 cm ≡ 10^{-2} m
or
10^{2} cm ≡ 1 m

1 µm ≡ 10^{-6} m
or
10^{6} µm ≡ 1 m

$$? \,\mu m = 0.34 \text{ cm} \times \frac{10^{-2}\text{ m}}{1\text{ cm}} \times \frac{1\,\mu m}{10^{-6}\text{ m}}$$

Note : Conversions within a system are exact by definition.

C) More Complicated Conversions

1) Ex 1: A good pitcher can throw a fastball at a speed of 90.0 mi/hr. How long will it take (in sec) to reach home plate 60.5 ft away?

60.5 ft $\Rightarrow$? sec

Have 90.0 mi/hr

Must convert units in both numerator and denominator

1 mi $\equiv$ 5280 ft   1 hr $\equiv$ 3600 s

2) Ex 2: A pool measures 60.500 ft by 30.500 ft by 10.0000 ft. How many cubic meters of water can the pool hold?

3) Ex 3: What volume will 50.0 g of ether occupy? The density of ether is 0.71 g/mL

Density can be used as a conversion factor between mass and volume