Chapter 1

Introduction: Matter and Measurement

Steps in the Scientific Method

1. Observations
   - quantitative
   - qualitative
2. Formulating hypotheses
   - possible explanation for the observation
3. Performing experiments
   - gathering new information to decide whether the hypothesis is valid

Outcomes Over the Long-Term

Theory (Model)
- A set of tested hypotheses that give an overall explanation of some natural phenomenon.

Natural Law
- The same observation applies to many different systems
- Example - Law of Conservation of Mass

Which statement explains the difference between a scientific law (X) and a scientific theory (Y)?

a. X is proven; Y is not proven.
b. X is not proven; Y is proven.
c. X tells what happens; Y explains why things happen.
d. X explains why things happen; Y tells what happens.

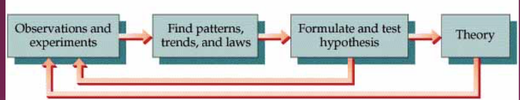
Which statement explains the difference between a scientific law (X) and a scientific theory (Y)?

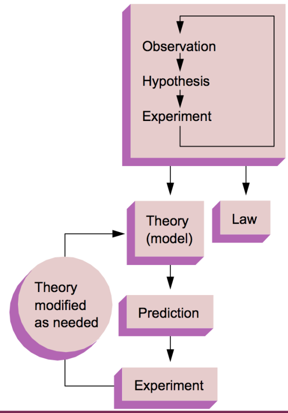
a. X is proven; Y is not proven.
b. X is not proven; Y is proven.
c. X tells what happens; Y explains why things happen.
d. X explains why things happen; Y tells what happens.

Law v. Theory

A law summarizes what happens; a theory (model) is an attempt to explain <u>why</u> it happens.

The Scientific Method





Nature of Measurement

Measurement - quantitative observation consisting of 2 parts

Part 1 - number
Part 2 - scale (unit)

Examples:

20 grams
6.63×10^{-34} Joule seconds

International System (le Système International)

Based on metric system and units derived from metric system.

The Fundamental SI Units

Physical Quantity	Name	Abbreviation
Mass	kilogram	kg
Length	meter	m
Time	second	s
Temperature	Kelvin	K
Electric Current	Ampere	A
Amount of Substance	mole	mol
Luminous Intensity	candela	cd

TABLE 1.5 Selected Prefixes Used in the Metric System

Prefix	Abbreviation	Meaning	Example
Giga	G	10^9	1 gigameter (Gm) $= 1 \times 10^9$ m
Mega	M	10^6	1 megameter (Mm) $= 1 \times 10^6$ m
Kilo	k	10^3	1 kilometer (km) $= 1 \times 10^3$ 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) $= 1 \times 10^{-6}$ m
Nano	n	10^{-9}	1 nanometer (nm) $= 1 \times 10^{-9}$ m
Pico	p	10^{-12}	1 picometer (pm) $= 1 \times 10^{-12}$ m
Femto	f	10^{-15}	1 femtometer (fm) $= 1 \times 10^{-15}$ m

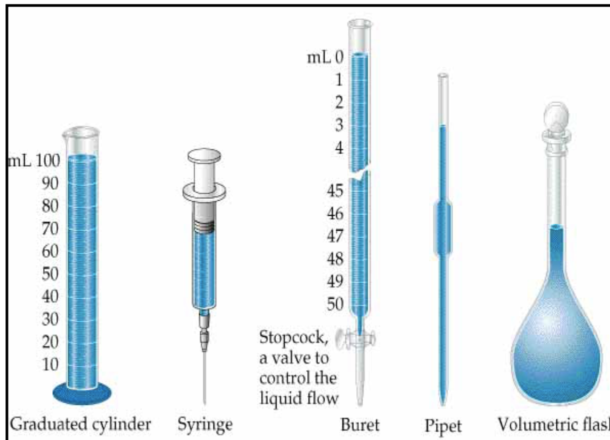
[a] This is the Greek letter mu (pronounced "mew").

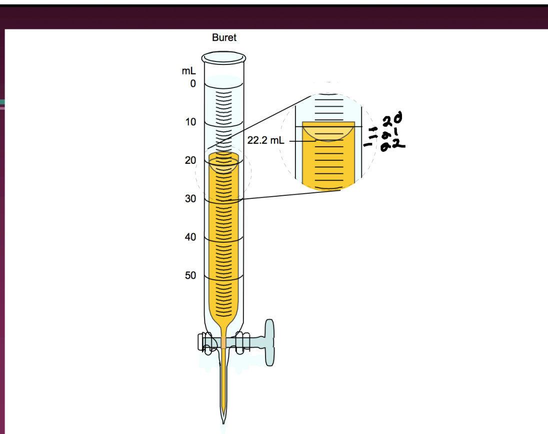
1 m = 100 cm

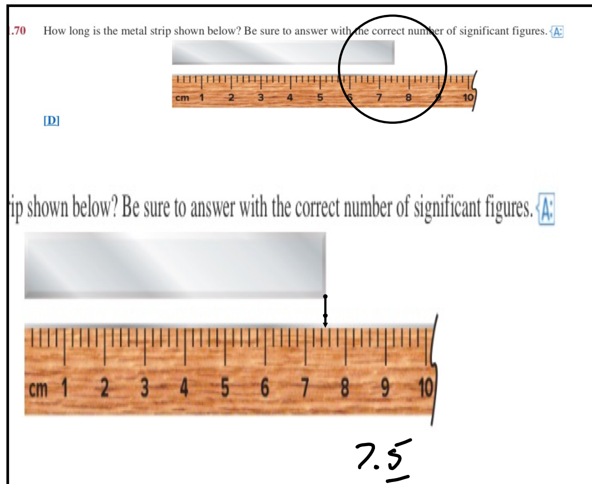
$$1\ \text{km} \times \frac{1000\ \text{m}}{1\ \text{km}} \times \frac{1000\ \text{mm}}{1\ \text{m}} = 10^6\ \text{mm}$$

Uncertainty in Measurement

A digit that must be estimated is called uncertain. A measurement always has some degree of uncertainty.



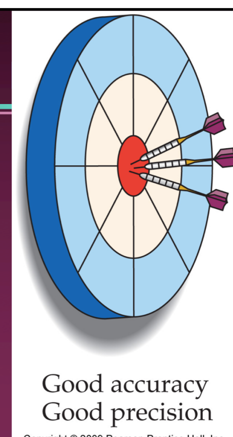


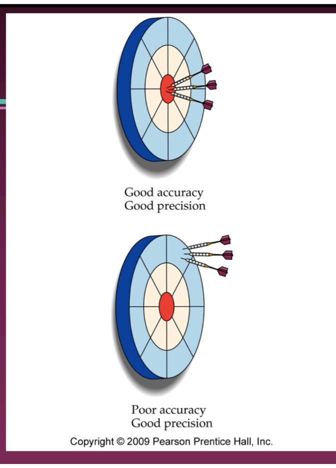


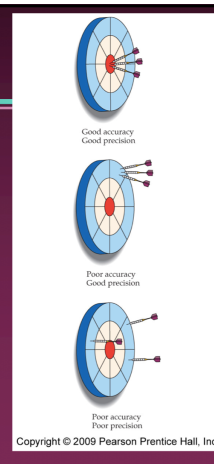
Precision and Accuracy

Accuracy refers to the agreement of a particular value with the true value.

Precision refers to the degree of agreement among several elements of the same quantity.







In three trials, a student measures a sample's mass to be to be 0.100 g, 0.600 g, and 0.300 g. The accepted value is 0.340 g. The student's data have

a. good accuracy and good precision.
b. good accuracy, but poor precision.
c. poor accuracy, but good precision.
d. poor accuracy and poor precision.

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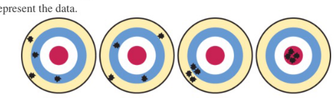
Types of Error

Random Error (Indeterminate Error) - measurement has an equal probability of being high or low.

Systematic Error (Determinate Error) - Occurs in the same direction each time (high or low), often resulting from poor technique.

1.73 The following dartboards illustrate the types of errors often seen in measurements. The bull's-eye represents the actual value, and the darts represent the data.



Exp. I Exp. II Exp. III Exp. IV

[D]

(a) Which experiments yield the same average result? I & III and II & IV

(b) Which experiment(s) display(s) high precision?

(c) Which experiment(s) display(s) high accuracy? IV & II

(d) Which experiment(s) show(s) a systematic error? III

Rules for Counting Significant Figures - Overview

1. Nonzero integers
2. Zeros
   - leading zeros
   - captive zeros
   - trailing zeros
3. Exact numbers

Rules for Counting Significant Figures - Details

Nonzero integers always count as significant figures.

3456 has
4 sig figs.

Rules for Counting Significant Figures - Details

Zeros
- Leading zeros do not count as significant figures.

0.0486 has
3 sig figs.

4.86×10^{-2}

Rules for Counting Significant Figures - Details

Zeros
- Captive zeros always count as significant figures.

16.07 has
4 sig figs.

Rules for Counting Significant Figures - Details

Zeros
- Trailing zeros are significant only if the number contains a decimal point.

9.300 has
4 sig figs.

8200 2 sig figs

8.2×10^{3}

Rules for Counting Significant Figures - Details

Exact numbers have an infinite number of significant figures.

1 inch = 2.54 cm, exactly

The measured quantity 0.0860 g contains ______ significant figures.

a. three
b. four
c. five
d. six

The measured quantity 0.0860 g contains ______ significant figures.

a. three
b. four
c. five
d. six

1.52 Underline the significant zeros in the following numbers:

(a) 0.41; (b) 0.041; (c) 0.0410; (d) 4.0100×10^4.

1.53 Underline the significant zeros in the following numbers:

(a) 5.08; (b) 508; (c) 5.080×10^3; (d) 0.05080.

1.55 Round off each number to the indicated number of significant figures (sf): (a) 231.554 (to 4 sf); (b) 0.00845 (to 2 sf); (c) 144,000 (to 2 sf).

a.) 231.554

231.6

b. 0.00845

0.0085

8.5×10^{-3}

Rules for Significant Figures in Mathematical Operations

Multiplication and Division: # sig figs in the result equals the number in the least precise measurement used in the calculation.

$$6.38 \times 2.0 =$$

$$12.76 \rightarrow 13 \text{ (2 sig figs)}$$

Rules for Significant Figures in Mathematical Operations

Addition and Subtraction: # decimal places in the result equals the number of decimal places in the least precise measurement.

$$6.8 + 11.934 =$$

$$18.734 \rightarrow 18.7 \text{ (3 sig figs)}$$

4.1 g + 7.08 g = _____ g

4.1

a. 11.180
b. 11.18
c. 11.2
d. 11

4.1 + 7.08 =

4.1 g + 7.08 g = _____ g

a. 11.180
b. 11.18
c. 11.2
d. 11

6.9 g – 5.07 g = _____ g

a. 1.830
b. 1.83
c. 1.8
d. 1

6.9 g – 5.07 g = _____ g

a. 1.830
b. 1.83
c. 1.8
d. 1

5.2 cm × 7.01 cm = ____ cm^2

a. 36
b. 36.4
c. 36.45
d. 36.452

5.2 cm × 7.01 cm = ____ cm^2

a. 36
b. 36.4
c. 36.45
d. 36.452

6.03 g ÷ 7.1 mL = _____ g/mL

a. 0.8
b. 0.85
c. 0.849
d. 0.849257

6.03 g ÷ 7.1 mL = _____ g/mL

a. 0.8
b. 0.85
c. 0.849
d. 0.849257

1.58 Carry out the following calculations, making sure that your answer has the correct number of significant figures:

(a) $\frac{2.795\ \text{m} \times 3.10\ \text{m}}{6.48\ \text{m}}$

(b) $V = \frac{4}{3}\pi r^3$, where $r = 17.282$ mm

(c) 1.110 cm + 17.3 cm + 108.2 cm + 316 cm

a. 1.337114 → 1.34 m

b. $V = \frac{4}{3} \times \pi \times (17.282)^3 =$ 21620.74286

21601 mm^3

c. = 442.610 → 443 cm

Dimensional Analysis

Proper use of "unit factors" leads to proper units in your answer.

1 in = 2.54 cm

1 pound = 453.59 g

1 gallon = 3.7854 L

Figure 01.25

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

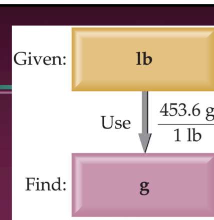
Desired unit; Given unit

2.54 cm = 1 in

$27.8\ \text{cm} \times \frac{1\ \text{in}}{2.54\ \text{cm}} = 10.95$ in   10.9448

10.9

$142\ \text{ft} \times \frac{12\ \text{in}}{1\ \text{ft}} \times \frac{2.54\ \text{cm}}{1\ \text{in}} = 4330$ cm

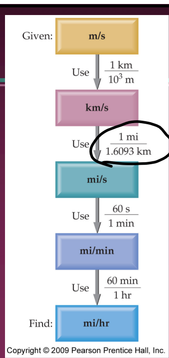


$160.5\ \text{lb} \times \frac{453.59\ \text{g}}{1\ \text{lb}} \times \frac{1\ \text{kg}}{1000\ \text{g}}$

72.8028

= kg

September 8, 2017



$$\frac{60.\ \text{mi}}{\text{hr}} \times \frac{1.6093\ \text{km}}{1\ \text{mi}} \times \frac{1000\ \text{m}}{1\ \text{km}} \qquad \frac{\text{m}}{\text{s}}$$

$$\times \frac{1\ \text{hr}}{60\ \text{min}} \times \frac{1\ \text{min}}{60\ \text{s}} = 26.8\ \frac{\text{m}}{\text{s}} \qquad 27\ \frac{\text{m}}{\text{s}}$$

1.28 What is the length in inches (in) of a 100.-m soccer field?

2.54 cm = 1 in

$$100.\ \text{m} \times \frac{100\ \text{cm}}{1\ \text{m}} \times \frac{1\ \text{in}}{2.54\ \text{cm}} = 3940\ \text{in}$$

$$3.94 \times 10^3\ \text{in}$$

1.34 The average density of Earth is 5.52 g/cm³. What is its density in (a) kg/m³; (b) lb/ft³?

2.54 cm = 1 in

1 kg = 2.205 lb

~~1 cm = 2.54 in~~

a.) $$\frac{5.52\ \text{g}}{\text{cm}^3} \times \frac{1\ \text{kg}}{1000\ \text{g}} \times \left(\frac{100\ \text{cm}}{1\ \text{m}}\right)^3$$

$$\frac{5.52\ \text{g}}{\text{cm}^3} \times \frac{1\ \text{kg}}{1000\ \text{g}} \times \frac{(100)^3\ \text{cm}^3}{1^3\ \text{m}^3} \qquad \frac{\text{kg}}{\text{m}^3}$$

$$= 5520\ \text{kg/m}^3$$

b. $$\frac{5.52\ \text{g}}{\text{cm}^3} \times \left(\frac{2.54\ \text{cm}}{1\ \text{in}}\right)^3 \times \left(\frac{12\ \text{in}}{1\ \text{ft}}\right)^3 \rightarrow \frac{\text{lb}}{\text{ft}^3}$$

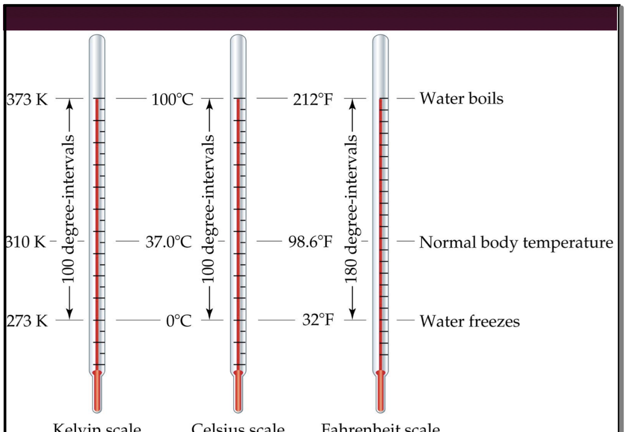
$$\times \frac{1\ \text{kg}}{1000\ \text{g}} \times \frac{2.205\ \text{lb}}{1\ \text{kg}} = 345 \qquad \frac{\text{lb}}{\text{ft}^3}$$

Temperature

Celsius scale =°C

Kelvin scale = K

Fahrenheit scale =°F



Is the following statement true or false? The “size” of a degree on the Celsius scale is the same as the “size” of a degree on the Kelvin scale.

a. True

b. False

Is the following statement true or false? The "size" of a degree on the Celsius scale is the same as the "size" of a degree on the Kelvin scale.

a. True

b. False

Temperature

$$T_K = T_C + 273.15$$

$$T_F = T_C \times \frac{9°F}{5°C} + 32°F$$

1.42 Perform the following conversions: [A]

(a) 68°F (a pleasant spring day) to °C and K

(b) −164°C (the boiling point of methane, the main component of natural gas) to K and °F

(c) 0 K (absolute zero, theoretically the coldest possible temperature) to °C and °F

a.) $T_F = T_C \frac{9}{5} + 32$

$T_C = \frac{(T_F - 32)5}{9} = \frac{(68-32)5}{9} = 20.°C$

$T_K = T_C + 273.15 = 293\ K$

b. $= ~~-327~~ °F$ -263 $= 109\ K$

Density

Density is the mass of substance per unit volume of the substance:

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

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

1.38 An empty vial weighs 55.32 g. (a) If the vial weighs 185.56 g when filled with liquid mercury (d = 13.53 g/cm³), what volume of mercury is in the vial? (b) How much would the vial weigh if it were filled with the same volume of water (d = 0.997 g/cm³ at 25°C)? [A]

a.) $d = \frac{m}{V}$ $V = \frac{m}{d}$

$m_{sub.} = 185.56_g - 55.32_g = 130.24_g$

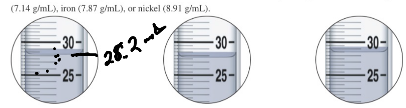
$V = \frac{130.24_g}{13.53 \frac{g}{cm^3}} = 9.626\ cm^3$

$1\ cm^3 = 1\ mL$

b. $m_w = d_w V_w = \quad g$

$m_v = m_w + m_{EV} = 64.92\ g$

1.44 A 25.0-g sample of each of three unknown metals is added to 25.0 mL of water in graduated cylinders A, B, and C, and the final volumes are depicted in the circles below. Given their densities, identify the metal in each cylinder: zinc (7.14 g/mL), iron (7.87 g/mL), or nickel (8.91 g/mL).



A B C

[D]

a. $d = \frac{m}{v} = \frac{25.0_g}{~~28.2~~ mL} = ~~0.8865~~ \frac{g}{mL}$

$28.2 - 25.0 = 3.2\ mL$

$\frac{25.0_g}{3.2\ mL} = 7.8 \frac{g}{mL}$

b.) $d = \frac{m}{v} = \frac{25.0\ g}{2.9\ mL} = 8.62 \frac{g}{mL}$

c.) $d = \frac{25.0_g}{3.5\ mL} = 7.14\ g/mL$

Table 01.06

TABLE 1.6 ■ Densities of Some Selected Substances at 25 °C	
Substance	**Density (g/cm^3)**
Air	0.001
Balsa wood	0.16
Ethanol	0.79
Water	1.00
Ethylene glycol	1.09
Table sugar	1.59
Table salt	2.16
Iron	7.9
Gold	19.32