Lecture PowerPoint

General Chemistry

Chapter 1 •

Introduction

Introduction Chapter 1 :

1.1 What Is Chemistry

1.5 Measurement in Scientific Study

1.6 Uncertainty in Measurement: Significant Figures

CHEMISTRY

Is the study of matter,

its properties,

the changes that matter undergoes,

and

the energy associated with these changes.

Definitions

Matter anything that has mass and volume -the “stuff” of the universe: books, planets, trees, professors, students

Composition the types and amounts of simpler substances that make up a sample of matter

Properties the characteristics that give each substance a unique identity

Physical Properties

those which the substance shows by itself without interacting with another substance such as color, melting point, boiling point, density

Chemical Properties

those which the substance shows as it interacts with, or transforms into, other substances such as flammability, corrosiveness

Some Characteristic Properties of Copper

Table 1.1 Some Characteristic Properties of Copper

Physical Properties

reddish brown, metallic luster

easily shaped into sheets (malleable) and wires (ductile)

good conductor of heat and electricity

can be melted and mixed with zinc to form brass

density = 8.95 g/cm^3

melting point = 1083˚C

boiling point = 2570˚C

Chemical Properties

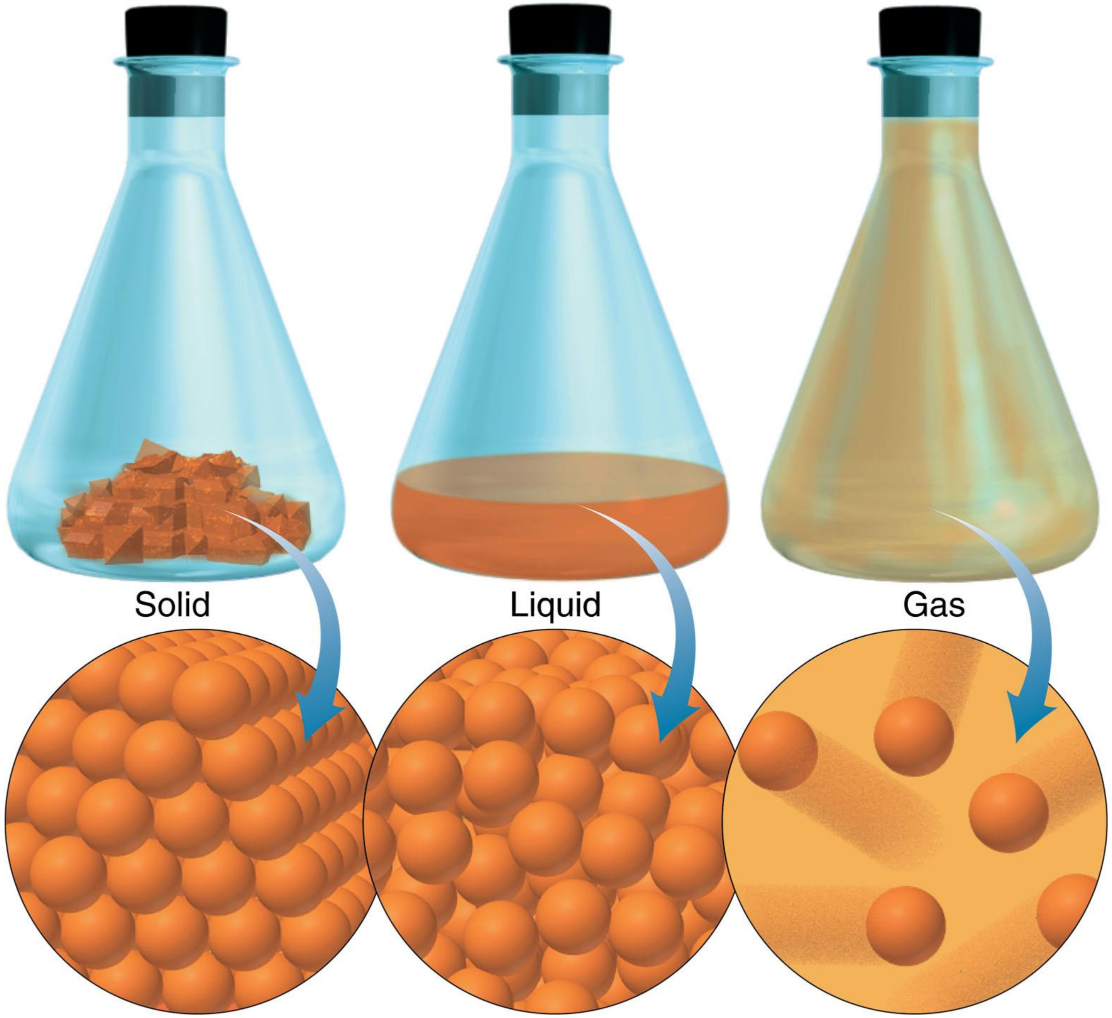
slowly forms a basic blue-green sulfate in moist air

reacts with nitric acid and sulfuric acid

slowly forms a deep-blue solution in aqueous ammonia

The Physical States of Matter

Figure 1.2 **The physical states of matter.**



Sample Problem 1.2 Distinguishing Between Physical and Chemical Change

PROBLEM: Decide whether each of the following processes is primarily a physical or a chemical change, and explain briefly:

(a) Frost forms as the temperature drops on a humid winter night.

(b) A cornstalk grows from a seed that is watered and fertilized.

(c) A match ignites to form ash and a mixture of gases.

(d) Perspiration evaporates when yo relax after jogging.

PLAN: "Does the substance change composition or just change form?"

SOLUTION:

(a) physical change (b) chemical change (c) chemical change

(d) physical change

International System Of Units (SI)

SI Base Units (French System)

Physical Quantity (Dimension)	*Unit Name*	*Unit 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

Working With Larger and Smaller Units

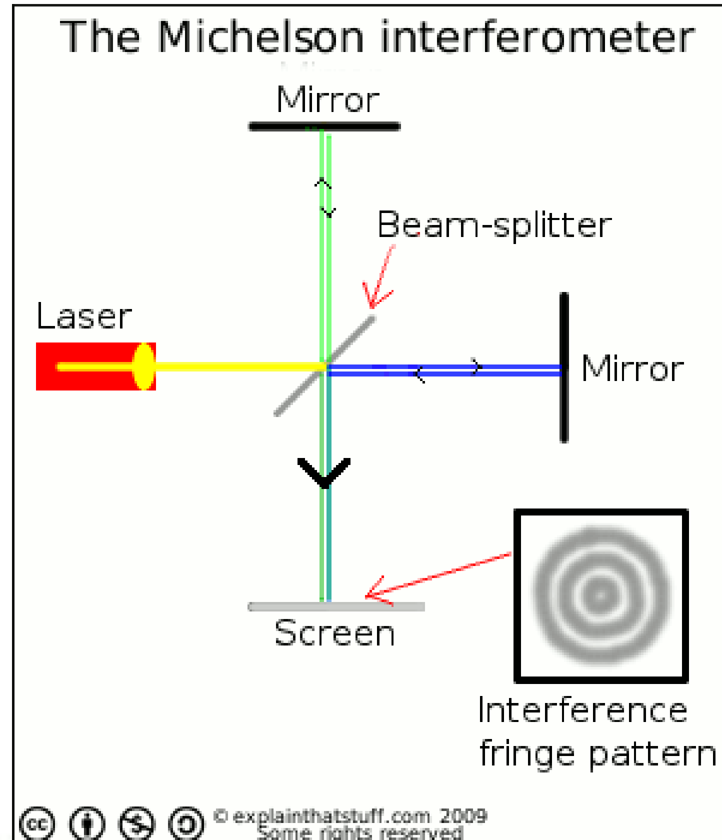
Table 1.3 **Common Decimal Prefixes Used with SI Units**

Prefix	*Prefix Symbol*	*Word*	*Conventional Notation*	*Exponential Notation*
tera	T	trillion	1,000,000,000,000	1×10^{12}
giga	G	billion	1,000,000,000	1×10^{9}
mega	M	million	1,000,000	1×10^{6}
kilo	k	thousand	1,000	1×10^{3}
hecto	h	hundred	100	1×10^{2}
deka	da	ten	10	1×10^{1}
-----	----	one	1	1×10^{0}
deci	d	tenth	0.1	1×10^{-1}
centi	c	hundredth	0.01	1×10^{-2}
milli	m	thousandth	0.001	1×10^{-3}
micro	μ	millionth	0.000001	1×10^{-6}
nano	n	billionth	0.000000001	1×10^{-9}
pico	p	trillionth	0.000000000001	1×10^{-12}
femto	f	quadrillionth	0.000000000000001	1×10^{-15}

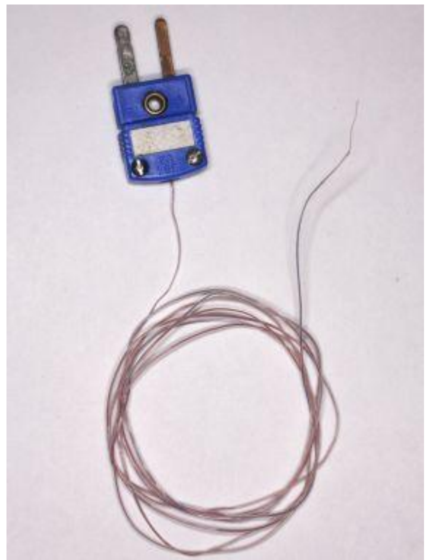
Systems for length measurement

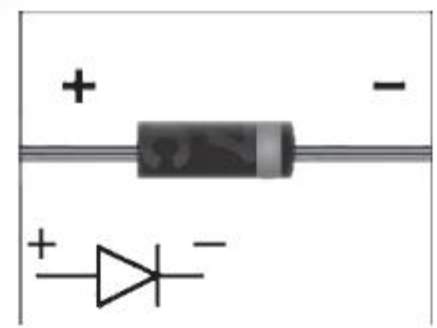
Ruler, Calliper/Vernier, Micrometer

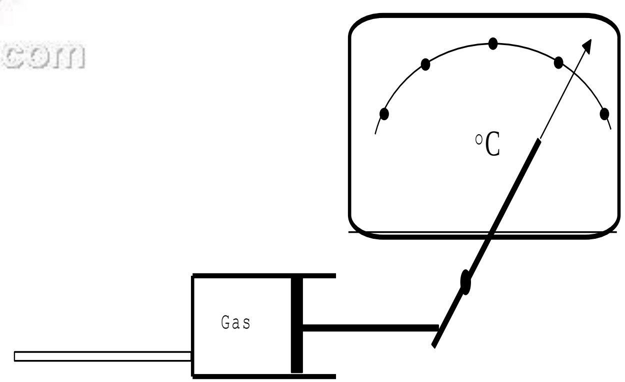
Michelson interferometer



Temperature Measurement



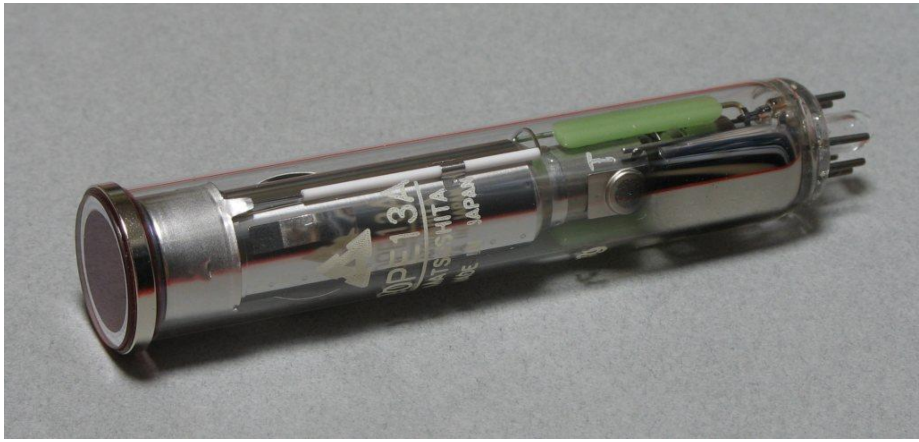


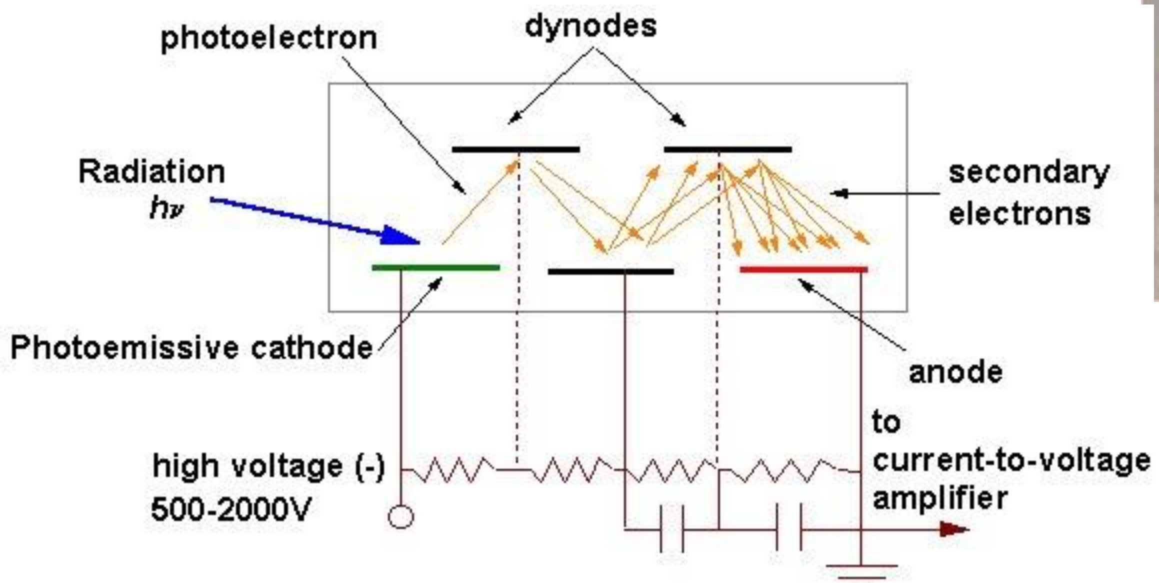


Instruments for heat measurement Bomb Calorimeter

Constant pressure calorimeter

Systems for light intensity measurement





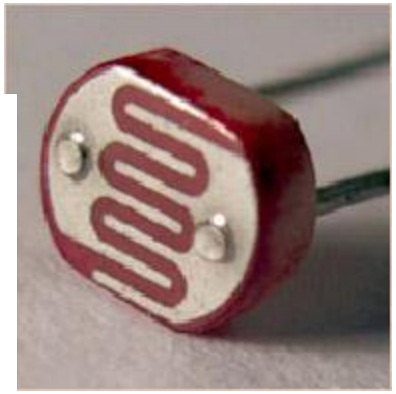


Table 1.4 Common SI-English Equivalent Quantities

Quantity	SI to English Equivalent	English to SI Equivalent
Length	1 km = 0.6214 mile	1 mi = 1.609 km
	1 m = 1.094 yard	1 yd = 0.9144 m
	1 m = 39.37 inches	1 ft = 0.3048 m
	1 cm = 0.3937 inch	1 in = 2.54 cm
Volume	1 cubic meter (m^3) = 35.31 ft^3	1 ft^3 = 0.02832 m^3
	1 dm^3 = 0.2642 gal	1 gal = 3.785 dm^3
	1 dm^3 = 1.057 qt	1 qt = 0.9464 dm^3
	1 cm^3 = 0.03381 fluid ounce	1 qt = 946.4 cm^3
		1 fluid ounce = 29.57 cm^3
Mass	1 kg = 2.205 lb	1 lb = 0.4536 kg
	1 g = 0.03527 ounce (oz)	1 oz = 28.35 g

Conversions Among SI Unites

PROBLEM: A desk is found to be 1437 mm wide. What is this width expressed in meters?

SOLUTION:

According to table 1.3 Each mm is equivalent to part of thousandth of meter

1mm → 1×10^{-3} m

1437 mm → ?

Cross Multiplication: ? = 1437 x 1×10^{-3} m

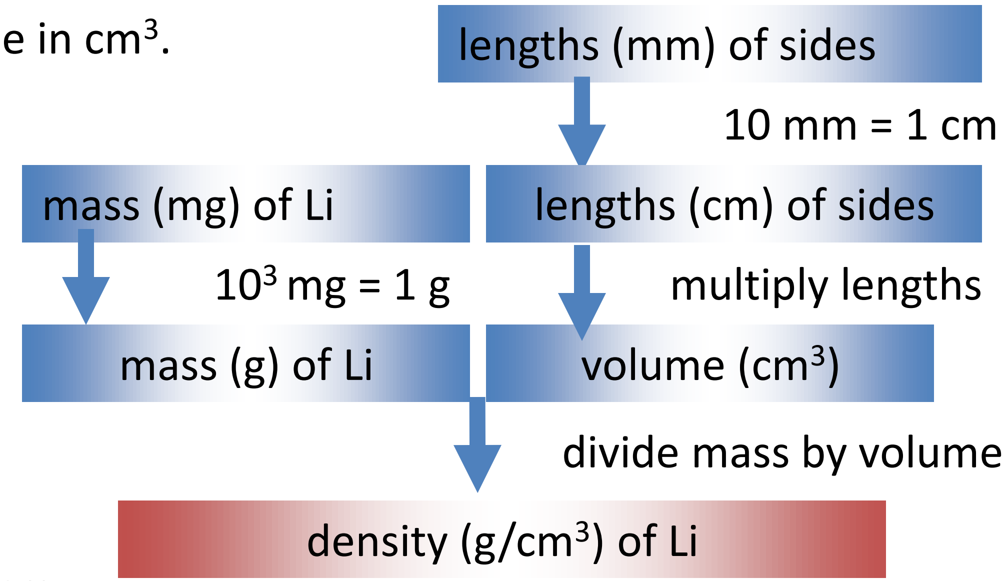
1

= 1.437 m

Sample Problem — Calculating Density from Mass and Length

PROBLEM: If a rectangular slab of Lithium (Li) weighs 1.49×10^3 mg and has sides that measure 20.9 mm by 11.1 mm by 11.9 mm, what is the density of Li in g/cm^3 ?

PLAN: Density is expressed in g/cm^3 so we need the mass in grams and the volume in cm^3.



SOLUTION:

$$1.49\times10^3 \text{ mg} \times \frac{1 \text{ g}}{1000 \text{ mg}} = 1.49 \text{ g}$$

$$20.9 \text{ mm} \times \frac{1 \text{ cm}}{10 \text{ mm}} = 2.09 \text{ cm}$$

Similarly the other sides will be 1.11 cm and 1.19 cm, respectively.

$$2.09 \times 1.11 \times 1.19 = 2.76 \text{ cm}^3$$

$$\text{density of Li} = \frac{1.49 \text{ g}}{2.76 \text{ cm}^3} = 0.540 \text{ g/cm}^3$$

Homework # 1:

Checkpoint 1.3 (Page: 13)

Questions: 1.3.2

1.3.4

Checkpoint 1.4 (Page 16)

Questions: 1.4.1

1.4.2

<u>1.5 (PAGE: 17) Uncertainty in Measurements and Significant Figures:</u>

In Chemistry there are tow types of numbers:

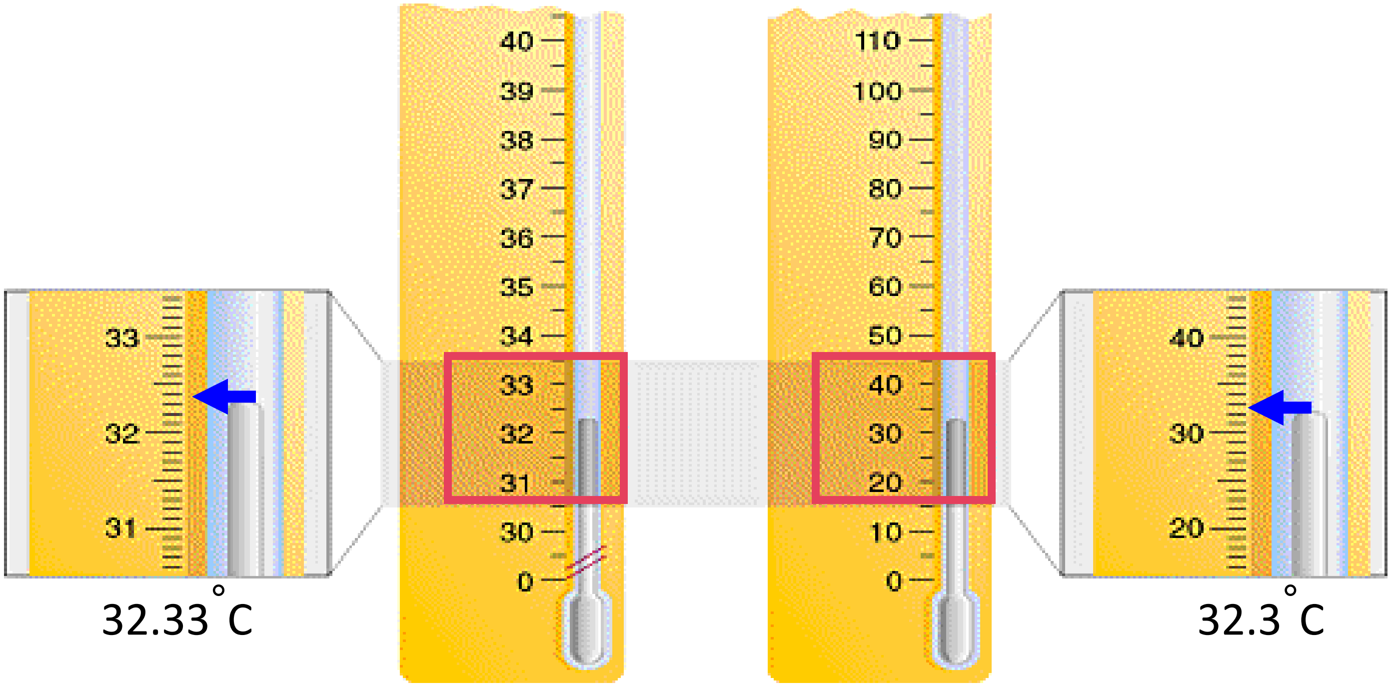
-**EXACT NUMBERS:** Numbers with defined values such as: Dozen=12, 1 kg=1000 g, 1 inch = 2.54 cm,……..etc. **SUCH EXACT NUMBERS DO NOT CONTAIN SIGNIFICANT FIGURES.**

-INEXACT NUMBERS: Measured numbers

All measured number are inexact because of the uncertainty in the measuring devices and the individual who use them. **TO INDICATE THE UNCERTAINETY IN SUCH MEASURED NUMBERS, THESE NUMBERS MUST BE REPORTED USING SIGNIFICANT FIGURES.**

Figure 1.14A

The number of significant figures in a measurement depends upon the measuring device.



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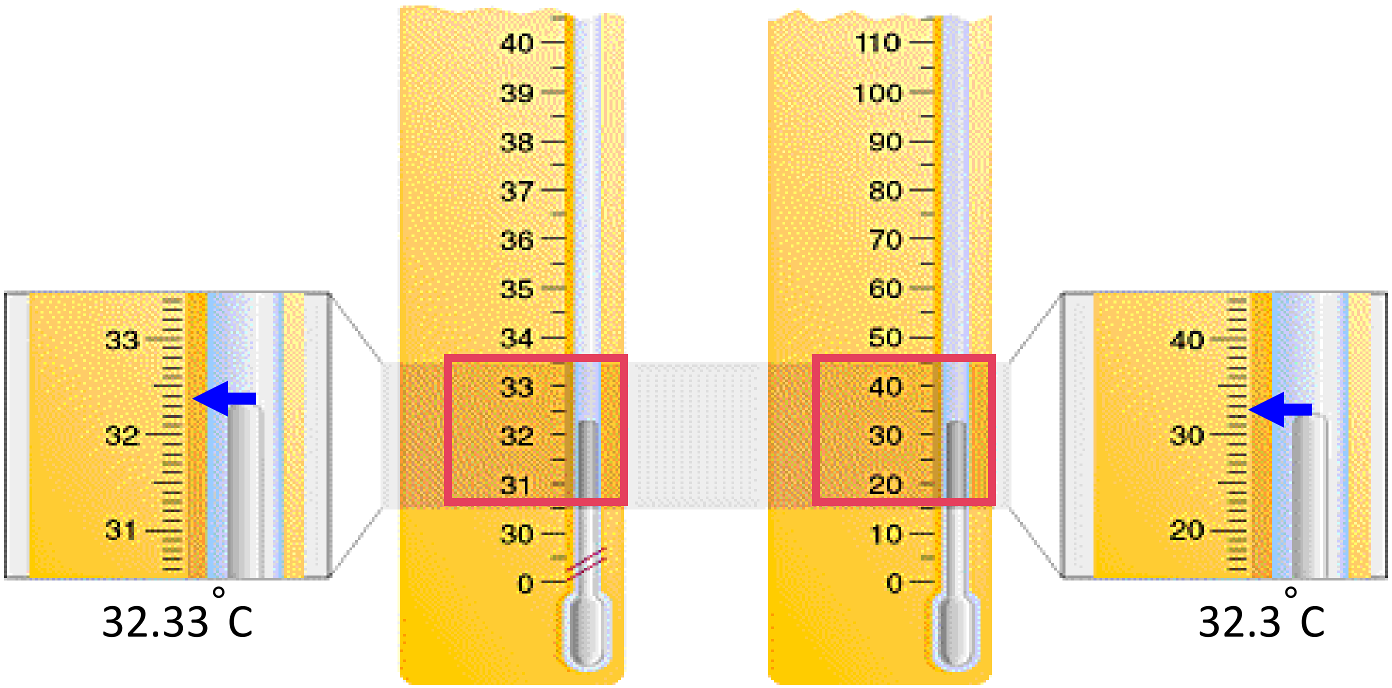


Rules for Determining Which Digits are Significant

- Any digit that is not zero **is significant** (112.333 has 6 sig. fig.)

- Zeros located between nonzero digits **are sig.** (3004 has 4 sig fig)

- Zeros to the left of the first nonzero digit **are not sig.** (0000.000000001 has 1 sf)

-Zeros to the right of the last nonzero digit **are sig**. **IF THE NUMBER CONTAINS A DECIMAL**. (100.0 has 4sf, 1.200 has 4 sf)

- Zeros to the right of the last nonzero digit in a number that **does not contain decimal** **may or may not be sig. You have to express these numbers using *<u>SEIENTIFIC NOTATION.</u>***

(IN 100 THERE ARE THREE METHODS TO EXPRESS THIS NUMBER USING SCIENTIFIC NOTATION: $1X10^2$ **has 1 sf**

1.0×10^2 **has 2 sf**

1.00×10^2 **has 3 sf.**

Determining the Number of Significant Figures

PROBLEM: •For each of the following quantities, underline the zeros that are significant figures (sf), and determine the number of significant figures in each quantity. For **(d)** to **(f),** express each in exponential notation first.

(a) 0.0030 L

(b) 0.1044 g

(c) 53,069 mL

(d) 0.00004715 m

(e) 57,600. s

(f) 0.0000007160 cm^3

SOLUTION:

(a) 0.0030 L 2sf

(b) 0.1044 g 4sf

(c) 53.069 mL 5sf

(d) 0.00004715 m
4.715×10^{-5} m 4sf

(e) 57,600. s
5.7600×10^{4} s 5sf

(f) 0.0000007160 cm^3
7.160×10^{-7} cm^3 4sf

Rules for Significant Figures in Calculations (Page 18)

1. *For multiplication and division*. The answer contains the same number of significant figures as there are in the measurement with the fewest significant figures.

Multiply the following numbers:

9.2 cm x 6.8 cm x 0.3744 cm = 23.4225 cm^3 = 23 cm^3

Rules for Significant Figures in Calculations

2. *For addition and subtraction*. The answer has the *same number of decimal places as there are in the measurement with the fewest decimal places.*

Example: adding two volumes

$$\begin{array}{r} 83.5 \text{ mL} \\ +\ 23.28 \text{ mL} \\ \hline 106.78 \text{ mL} \end{array} = \mathbf{106.8} \text{ mL}$$

Example: subtracting two volumes

$$\begin{array}{r} 865.9 \text{ mL} \\ -\ 2.8121 \text{ mL} \\ \hline 863.0879 \text{ mL} \end{array} = \mathbf{863.1} \text{ mL}$$

3. Exact numbers do not limit the number of sf in the answer.

3 x 2.4 g = 7.2 g not 7

Rules for Rounding Off Numbers

1. If the digit removed is 5 or *more than 5,* the preceding number increases by 1.
5.379 rounds to 5.38 if three significant figures are retained and to 5.4 if two significant figures are retained.

2. If the digit removed is *less than 5,* the preceding number is unchanged.
0.2413 rounds to 0.241 if three significant figures are retained and to 0.24 if two significant figures are retained.

3. Exact numbers **do not limit** the number of sf in the calculated result.
Example: 2x 2.11 cm x 1.0 cm = 4.2 NOT 4

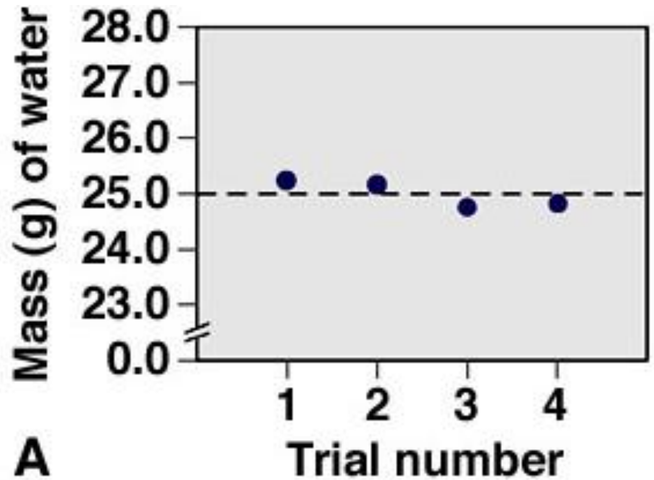
4. Be sure to carry two or more additional significant figures through a multistep calculation and round off only the *final* answer only.
EXAMPLE: A X B= C
C X D= E
If A= 3.66, B= 8.45, D= 2.11
AXB= 3.66 X 8.45=30.927 **NOT 30.9**
CXD= 30.927X 2.11= 65.3 **NOT 65.2**
If D= 2.11000 THEN YOU HAVE TO ROUND THE FIRST STEP
C X D = 30.9 X 2.11000 = 65.2

Precision and Accuracy
Errors in Scientific Measurements (page 20-21)

Precision -
Refers to *reproducibility* or how close the measurements are <u>to each other</u>.

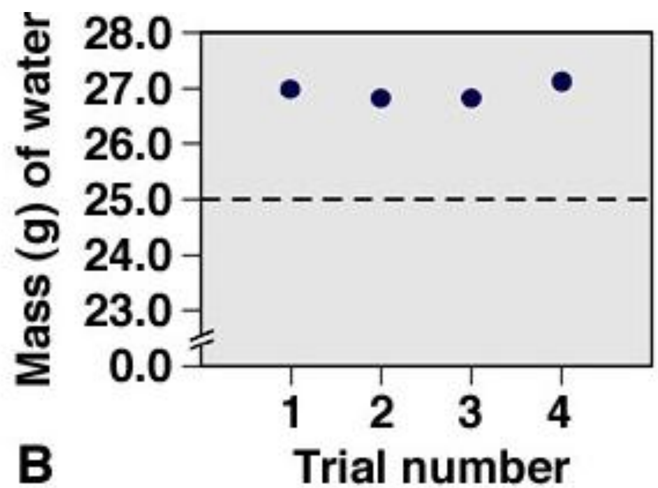
Accuracy -
Refers to how close a measurement is to the <u>real value</u>.

Precision and accuracy in the laboratory.



precise and accurate

precise but not accurate



Home work # 2

Checkpoint 1.5 (P 21)
Questions: 1.5.2, 1.5.3, 1.5.4

Checkpoint 1.6 (P 24)
Question 1.6.1

QUESTIONS AND PROBLEMS (P27-32)
Questions: 1.36, 1.50, 1.72, 1.65