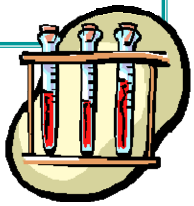
Chapter 1

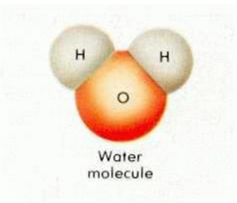
Chemistry: The Central Science

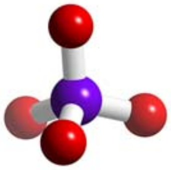


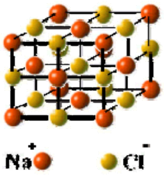
Chapter 1 Section 1

Why Chemistry?

- Everything in this universe is made out of approximately 100 different kinds of atoms.
  - Sand (Silicon, Oxygen)
  - Table Salt (Sodium, Chloride)
  - Water (Oxygen, Hydrogen)
- They are as letters in an alphabet.

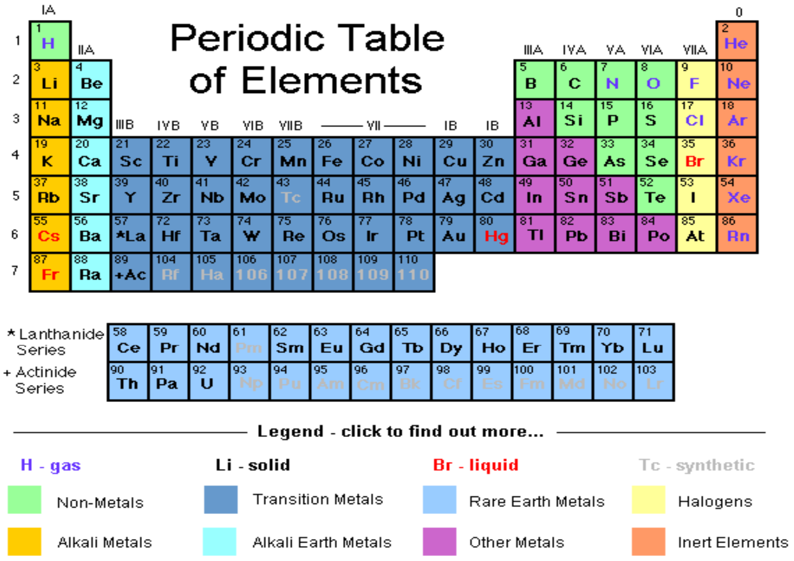






Chapter 1 Section 1

Why Chemistry?



Periodic Table of Elements

	IA	IIA	IIIB	IVB	VB	VIB	VIIB	VII			IB	IB	IIIA	IVA	VA	VIA	VIIA	0
1	1 H																	2 He
2	3 Li	4 Be											5 B	6 C	7 N	8 O	9 F	10 Ne
3	11 Na	12 Mg											13 Al	14 Si	15 P	16 S	17 Cl	18 Ar
4	19 K	20 Ca	21 Sc	22 Ti	23 V	24 Cr	25 Mn	26 Fe	27 Co	28 Ni	29 Cu	30 Zn	31 Ga	32 Ge	33 As	34 Se	35 Br	36 Kr
5	37 Rb	38 Sr	39 Y	40 Zr	41 Nb	42 Mo	43 Tc	44 Ru	45 Rh	46 Pd	47 Ag	48 Cd	49 In	50 Sn	51 Sb	52 Te	53 I	54 Xe
6	55 Cs	56 Ba	57 *La	72 Hf	73 Ta	74 W	75 Re	76 Os	77 Ir	78 Pt	79 Au	80 Hg	81 Tl	82 Pb	83 Bi	84 Po	85 At	86 Rn
7	87 Fr	88 Ra	89 +Ac	104 Rf	105 Ha	106 106	107 107	108 108	109 109	110 110								

* Lanthanide Series	58 Ce	59 Pr	60 Nd	61 Pm	62 Sm	63 Eu	64 Gd	65 Tb	66 Dy	67 Ho	68 Er	69 Tm	70 Yb	71 Lu
+ Actinide Series	90 Th	91 Pa	92 U	93 Np	94 Pu	95 Am	96 Cm	97 Bk	98 Cf	99 Es	100 Fm	101 Md	102 No	103 Lr

Legend - click to find out more...

H - gas
Li - solid
Br - liquid
Tc - synthetic

🟩 Non-Metals
🟦 Transition Metals
🟦 Rare Earth Metals
🟨 Halogens
🟧 Alkali Metals
🟦 Alkali Earth Metals
🟪 Other Metals
🟧 Inert Elements

Chapter 1 Section 1

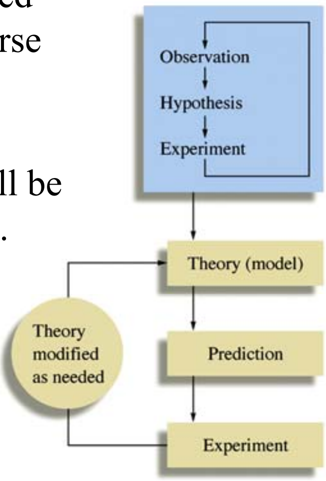
Why Chemistry?

- Chemical reactions are very common in this life and are important for our survival on the Earth.
- They also cause many problems for humanity!
- Can you think of "good" and "bad" chemical reactions?

Chapter 1 Section 1

Scientific Method

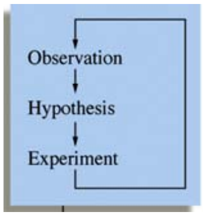
- The **scientific method** is the way used by scientists to understand the universe and its changes.
- The more creative you're at solving problems, the more effective you will be in your career and your personal life.
- Chemistry helps to develop solving problems capabilities.



Chapter 1 Section 1

Scientific Method

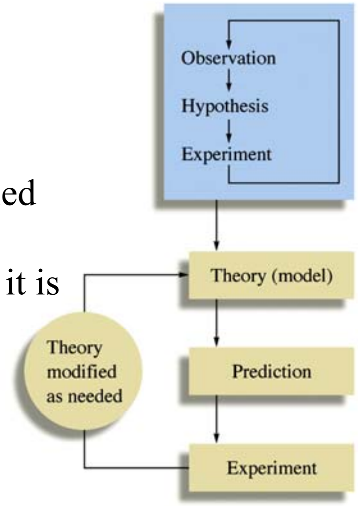
- To understand the universe and its changes.
- Steps of scientific methods:
  - Observation.
    - Can be qualitative or quantitative.
  - Prediction (Hypothesis).
    - Trying to explain the observation.
  - Experiment.
    - Performed to test the validity of the hypothesis.
    - Experiments always produce new information.



Chapter 1 Section 1

Scientific Method

- A set of valid hypotheses is assembled into a **theory** (model).
- This **theory** is confirmed, modified or may be discarded as more **observations** are recorded. Thus, it is a continuous process.



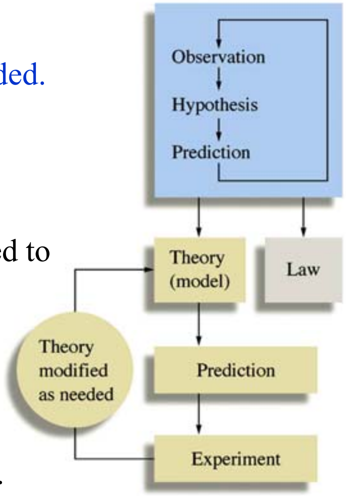
Chapter 1 Section 1

Theory vs. Law

- Theory changes over time as more observations from experiments are recorded.
- Some observations are found to apply to many different systems.
- Law (what happens)

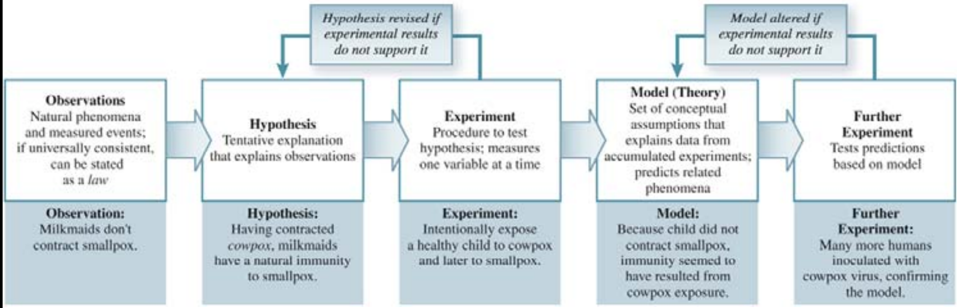
  A summary of observed behaviors applied to different systems.
  - Law of conservation of mass.
  - Law of conservation of energy.
- Theory (why it happens)

  An attempt to explain these observations.



Chapter 1 Section 1

Scientific Method



Hypothesis revised if experimental results do not support it

Model altered if experimental results do not support it

Observations Natural phenomena and measured events; if universally consistent, can be stated as a *law*

Hypothesis Tentative explanation that explains observations

Experiment Procedure to test hypothesis; measures one variable at a time

Model (Theory) Set of conceptual assumptions that explains data from accumulated experiments; predicts related phenomena

Further Experiment Tests predictions based on model

Observation: Milkmaids don't contract smallpox.

Hypothesis: Having contracted *cowpox*, milkmaids have a natural immunity to smallpox.

Experiment: Intentionally expose a healthy child to cowpox and later to smallpox.

Model: Because child did not contract smallpox, immunity seemed to have resulted from cowpox exposure.

Further Experiment: Many more humans inoculated with cowpox virus, confirming the model.

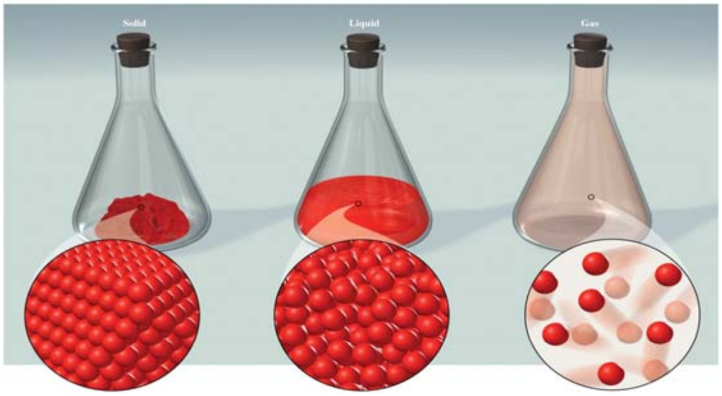
Chapter 1 Section 2

Classification of Matter

- ***Matter*** is anything occupying the space and having a mass.
- Matter can be:
  - A ***substance*** is a form of matter that has a defined composition and unique properties.
  - A ***mixture*** is a combination of two or more substances that retain their distinct identities.

Chapter 1 Section 2

Classification of Matter

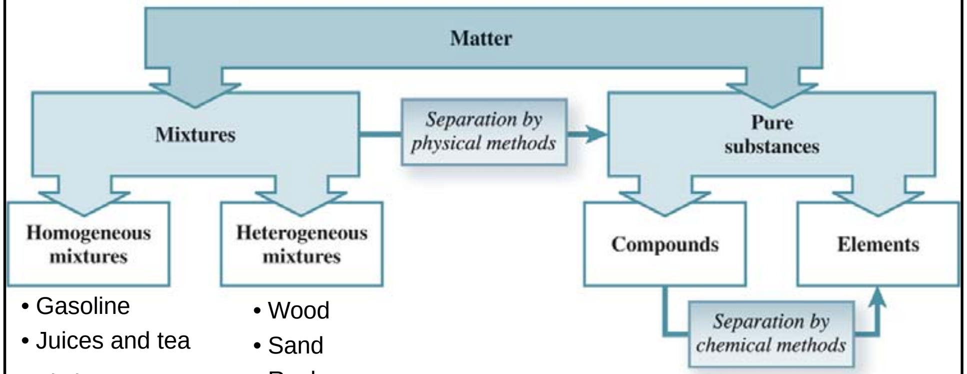


States of matter:

- <u>Solid</u>: rigid, has fixed volume and shape.
- <u>Liquid</u>: has definite volume but not fixed shape.
- <u>Gas</u>: has no fixed volume or shape. and is compressible.

Chapter 1 Section 2

Classification of Matter



Matter

Mixtures

Separation by physical methods

Pure substances

Homogeneous mixtures

Heterogeneous mixtures

Compounds

Elements

Separation by chemical methods

- Gasoline
- Juices and tea
- Air (N_2, O_2, H_2, CO_2, etc.)

- Wood
- Sand
- Rocks

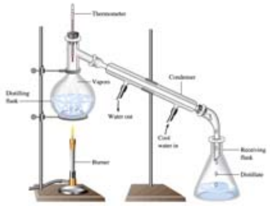
- H_2O
- NaCl
- H_2SO_4

- C (carbon)
- Hg (mercury)
- H or H_2 (hydrogen)

Chapter 1 Section 2

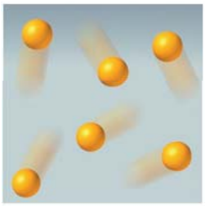
Separation (Physical) Methods

- Based on the physical properties of the substances (boiling point, adsorption, solubility, etc.)
- Separation methods discussed in the text:
  - Distillation.
  - Filtration.
  - Chromatography.

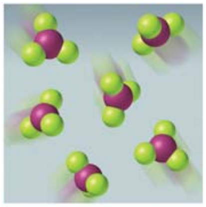


Chapter 1 Section 2

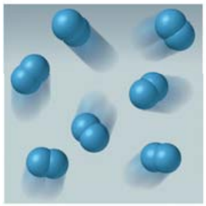
Elements and Compounds



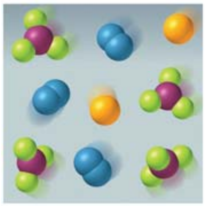
Atoms of an element



Molecules of a compound



Molecules of an element



Mixture of elements and a compound

Elements are substances that can't be decomposed into simpler substances by physical or chemical means

Compounds are substances with constant composition that can be broken down into elements by chemical processes

Chemical changes (decomposition of compounds or recombining elements), such as electrolysis, heating, and photolysis.

Classification of Matter

- Classify the following as mixtures, elements, compounds, etc.

Aluminum foil: substance, element

Table salt: substance, compound

Milk: mixture, homogeneous

Sea water: mixture, homogeneous

Copper wire: substance, element

The Properties of Matter

- ***Quantitative*: expressed using *numbers*.**
- ***Qualitative*: expressed using *properties*.**

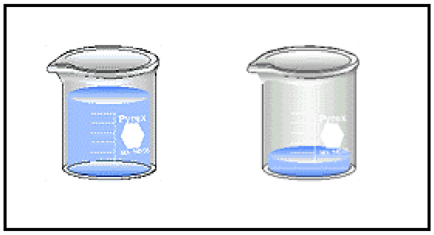
- ***Physical properties*:** can be observed and measured without changing the substance.
  - Examples: color, melting point, states of matter.
- ***Physical changes*:** the identity of the substance stays the same and only its state changes.
  - Examples: changes of state (melting, freezing).

The Properties of Matter

- ***Chemical properties*****:** must be determined by the chemical changes that are observed.
  - Examples: flammability, acidity, corrosiveness, reactivity.
- ***Chemical changes*****:** after a chemical change, the original substance no longer exists
  - Examples: combustion, digestion, corrosion.

The Properties of Matter

- ***Extensive property*****:** depends on the amount of matter.
  - Examples: mass, length.
- ***Intensive property*****:** does not depend on the amount of matter.
  - Examples: density, temperature, color.



Mass of water	100.0 g	10.0 g
Volume of water	0.100 L	0.010 L
Temperature of water	25 °C	25 °C
Density of water	1.00 g/mL	1.00 g/mL

Chapter 1 Section 3

Scientific Measurement

- Making observations can be done quantitatively or qualitatively.
- A ***quantitative*** observation is called a measurement. It must include two important pieces of information:

Number

Unit

There are two major systems of measurements:

(1) Metric Units

International System (SI Units)

Used in Science

(2) English Units

Chapter 1 Section 3

The Fundamental SI Units

TABLE 1.2 Base SI Units

Base Quantity	Name of Unit	Symbol
Length	meter	m
Mass	kilogram	kg
Time	second	s
Electric current	ampere	A
Temperature	kelvin	K
Amount of substance	mole	mol
Luminous intensity	candela	cd

All other units of measurement can be derived from the above seven fundamental SI units .

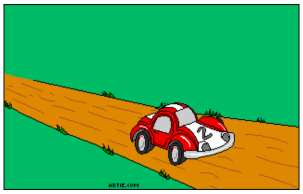
Units of Measurement

There are two major systems of measurements:

Number	Unit	
10	kilometers	(1) Metric Units International System (SI Units) Used in Science
1	gram	
5	kelvin	
20	miles	(2) English Units
1	pound	
60	Fahrenheit	

Using Prefixes in the SI System

- The distance between Dammam and Jubail is 90,000 meters.
  - 90×10^3 meters.
  - 90 kilo-meters (km)



- The capacity of this computer is 80,000,000,000 bites.
  - 80×10^9 bites
  - 80 giga-bites (GB)

Chapter 1 Section 3

Table of Prefixes in the SI System

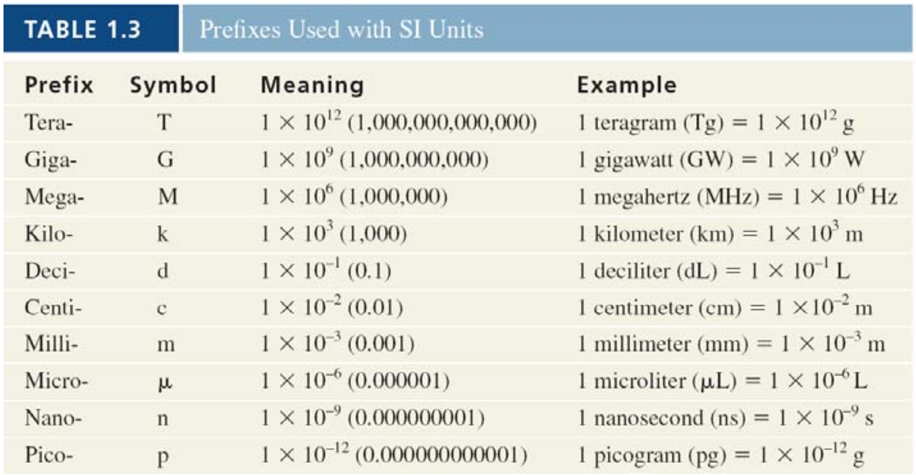


TABLE 1.3 Prefixes Used with SI Units

Prefix	Symbol	Meaning	Example
Tera-	T	1×10^{12} (1,000,000,000,000)	1 teragram (Tg) = 1×10^{12} g
Giga-	G	1×10^{9} (1,000,000,000)	1 gigawatt (GW) = 1×10^{9} W
Mega-	M	1×10^{6} (1,000,000)	1 megahertz (MHz) = 1×10^{6} Hz
Kilo-	k	1×10^{3} (1,000)	1 kilometer (km) = 1×10^{3} m
Deci-	d	1×10^{-1} (0.1)	1 deciliter (dL) = 1×10^{-1} L
Centi-	c	1×10^{-2} (0.01)	1 centimeter (cm) = 1×10^{-2} m
Milli-	m	1×10^{-3} (0.001)	1 millimeter (mm) = 1×10^{-3} m
Micro-	μ	1×10^{-6} (0.000001)	1 microliter (μL) = 1×10^{-6} L
Nano-	n	1×10^{-9} (0.000000001)	1 nanosecond (ns) = 1×10^{-9} s
Pico-	p	1×10^{-12} (0.000000000001)	1 picogram (pg) = 1×10^{-12} g

Must be memorized!

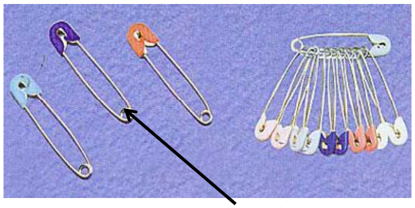
Chapter 1 Section 3

Scientific Notation

- 0.0001 kg
  - $1 \times \boxed{10^{-4}}$ kg
  - 0.1 g

Scientific Notation

- 1 with 35 zeros kg
  - $1 \times \boxed{10^{35}}$ kg
- Mass vs. Weight



How much does that pin weigh?

How much does the earth weigh?

Scientific Notations

- 123.1 = 1.231×10^2 = 1.231 × 100
- 0.00013 = 1.3×10^{-4} = 1.3 / 10000
  = 0.13×10^{-3}
- Avogadro's Number*:
  602,214,000,000,000,000,000,000

*The number of atoms contained in 12 g of carbon and is equal to 1 mole.

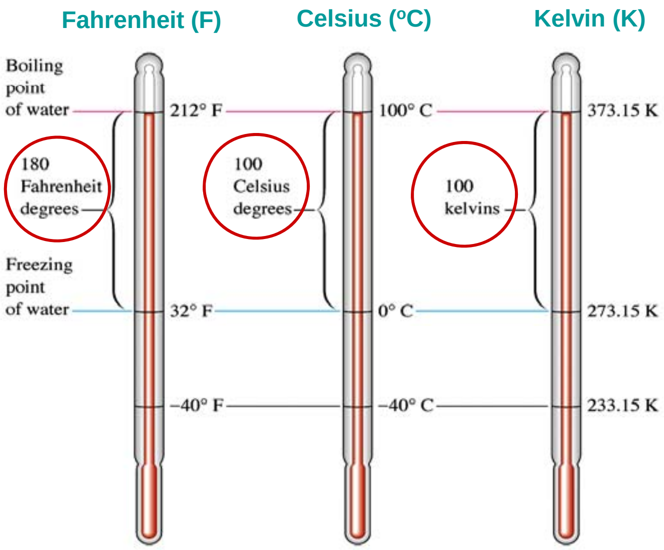
6.022×10^{23}

Scientific notation is a very convenient way to express the number of atoms in chemistry problems.

Temperature

- Three systems are used to measure temperatures:
  - Celsius scale (°C)
  - Kelvin scale (K)
  - Fahrenheit (°F)
- You have to be able to convert from one scale to another.

The Three Major Temperature Scales

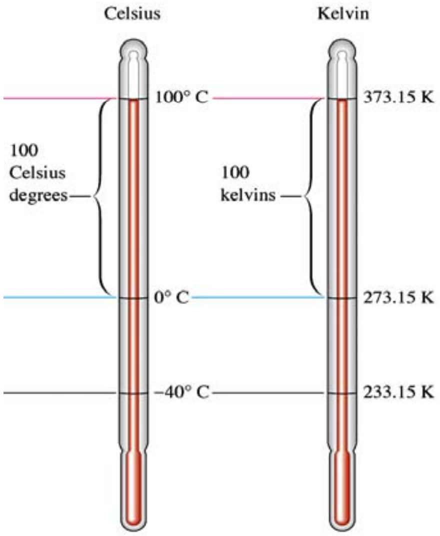


Celsius Scale vs. Kelvin Scale

- Temperature scales for °C and K are identical, but their zeros are different.

$$T_K = T_C + 273.15$$

$$T_C = T_K - 273.15$$



Chapter 1 Section 3

Fahrenheit Scale vs. Celsius Scale

- Both unit temperature size and zero locations are different.

Since:

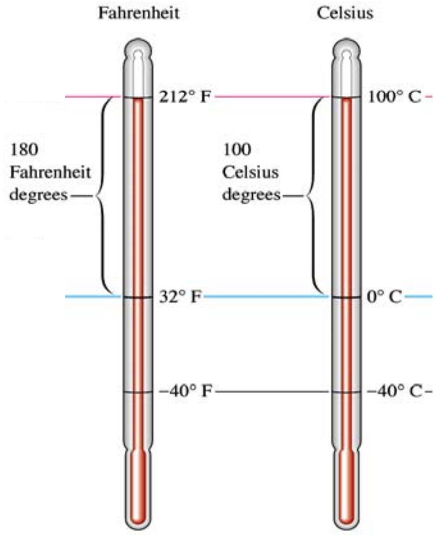
180°F = 100°C => 9°F = 5°C

and:

32°F = 0°C

Then to convert from °F to °C:

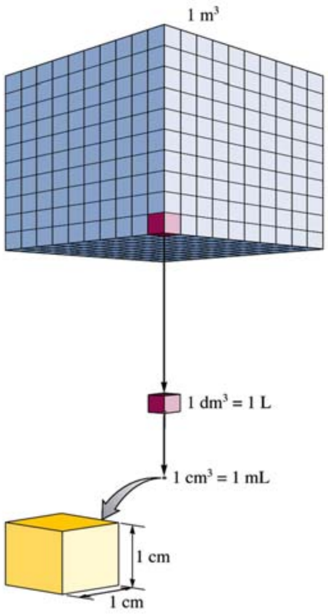
$$\boxed{[T_f\,(°\mathrm{F}) - 32\,(°\mathrm{F})]\,\frac{5°\mathrm{C}}{9°\mathrm{F}} = T_c\,(°\mathrm{C})}$$



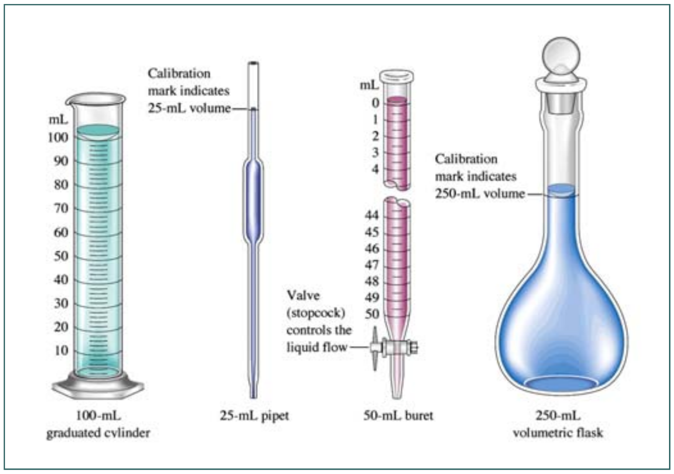
Chapter 1 Section 3

Volume

- Volume is not an SI unit, but it is extremely important in chemical measurements.
- Volume = 1m × 1m × 1m = 1m^3
  - 1m = 10dm
    $(1\text{m})^3 = (10\text{dm})^3$
    $1\text{m}^3 = 1000\text{dm}^3$
  - $1\text{dm}^3 = 1\text{L}$
    $1\text{L} = 1000\text{mL} = 1000\text{cm}^3$
    $1\text{mL} = 1\text{cm}^3$
  - (milli)Liter milli = 10^{-3}
    (centi)Meter centi = 10^{-2}



Measurement of Volume



Common types of laboratory equipment used to measure liquid volume.

Density

- It is the mass of substance per unit volume.

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

Substance	Physical State	Density (g/cm^3)
Oxygen	Gas	0.00133
Hydrogen	Gas	0.000084
Ethanol	Liquid	0.789
Water	Liquid	0.998
Aluminum	Solid	1.47
Iron	Solid	7.87
Mercury	Liquid	13.6

Chapter 1 Section 5

Uncertainty in Measurement

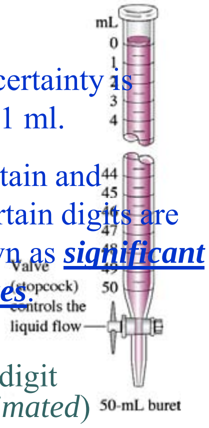
- A measurement always has some degree of uncertainty.

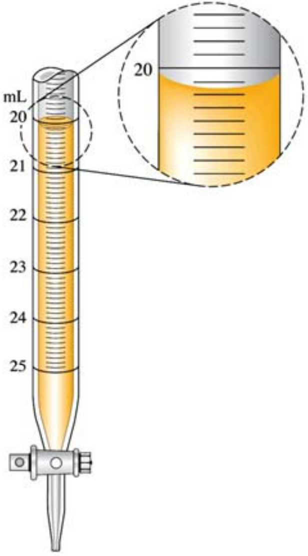
20.16 ml
20.17 ml
20.15 ml
20.18 ml
20.16 ml

certain digit

uncertain digit (must be *estimated*)

- Uncertainty is ± 0.01 ml.
- Certain and uncertain digits are known as ***significant figures***.





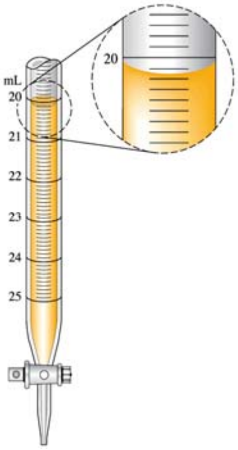
Chapter 1 Section 5

Uncertainty in Measurement

- Now you should be able to tell how many digits you need to include in your reading (measurement).

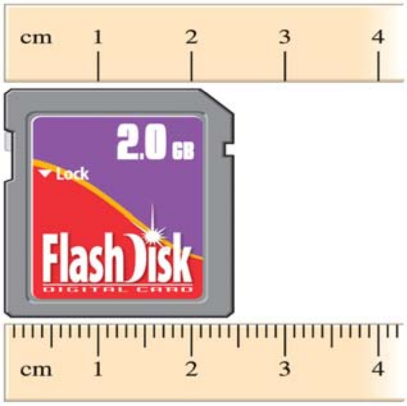
- 20 ml
  20.1 ml
  20.16 ml
  20.160 ml
  20.1600 ml

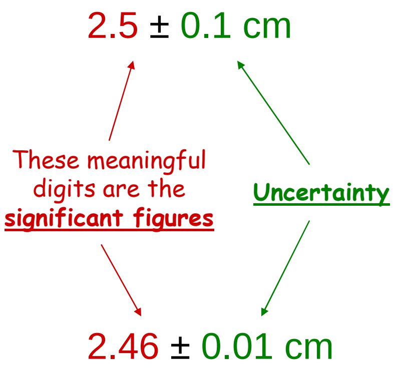
Which one??



These meaningful digits are the **significant figures**

Uncertainty in Measurement





2.5 ± 0.1 cm

These meaningful digits are the <u>significant figures</u>

<u>Uncertainty</u>

2.46 ± 0.01 cm

Different equipments have different uncertainties in their measurements.

Significant Figures

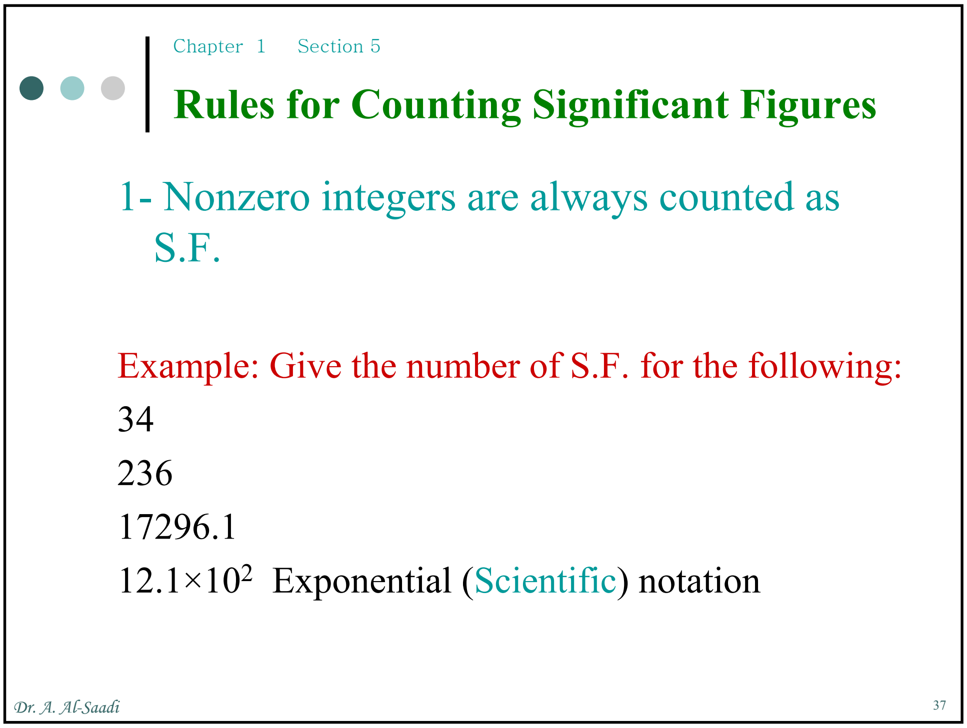
- In many cases, important physical quantities are obtained from measured values.

  Volume = *l w h*

  Density = mass / volume

Mathematical operations

- Calculations need to be done on the basis of ***Significant Figure (S.F.) Rules***
  - Rules for counting S.F.
  - Rules of mathematical operations on S.F.
- Implication of the word "Significant". It is to have the correct degree of uncertainty in the resultant physical quantities.



Chapter 1 Section 5

Rules for Counting Significant Figures

1- Nonzero integers are always counted as S.F.

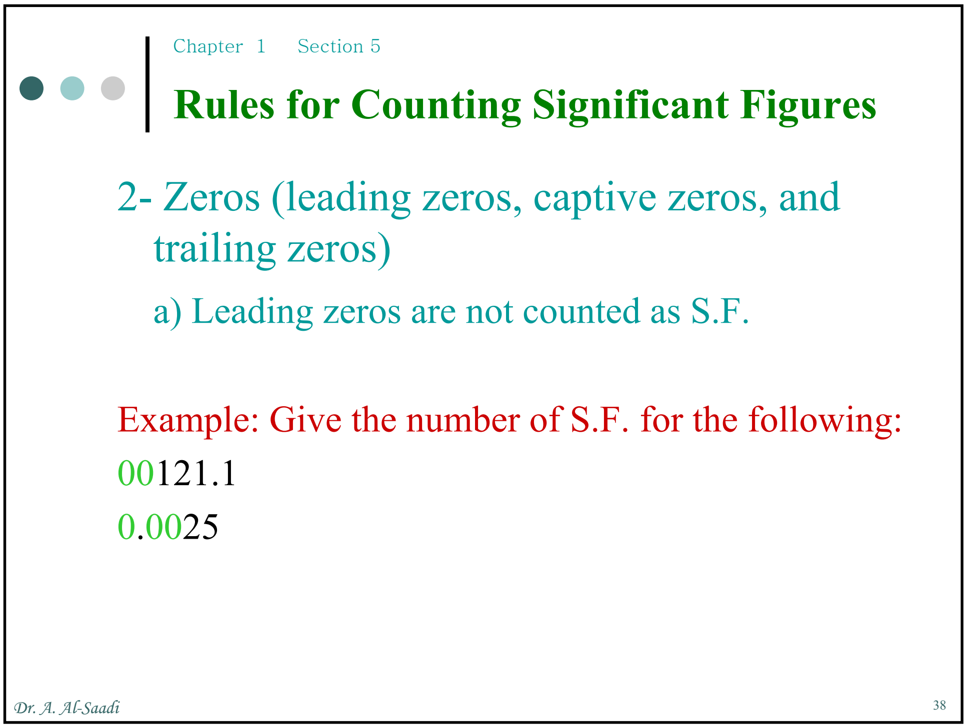
Example: Give the number of S.F. for the following:

34

236

17296.1

12.1×10^{2} Exponential (Scientific) notation



Chapter 1 Section 5

Rules for Counting Significant Figures

2- Zeros (leading zeros, captive zeros, and trailing zeros)

a) Leading zeros are not counted as S.F.

Example: Give the number of S.F. for the following:

00121.1

0.0025

Chapter 1 Section 5

Rules for Counting Significant Figures

b) Captive zeros are always counted as S.F.

Example: Give the number of S.F. for the following:

1.008

701.1×10^{-4}

3.000000008

0.0901

Chapter 1 Section 5

Rules for Counting Significant Figures

c) Trailing zeros are counted as S.F. only if the number contains a decimal point.

Example: Give the number of S.F. for the following:

1.000

320.00×10^{-1}

100 (can’t tell!) may be 1, 2 or 3.

100.

100.0

Rules for Counting Significant Figures

3- Exact numbers are assumed to have an infinite number of S.F.

Examples:

3 Apples is 3.00000000 (zeros are all the way to ∞)

2 in 2πr (the circumference of a cycle).

Mathematical relationships

1 km = 1000 m

1 in = 2.54 cm

Definitions

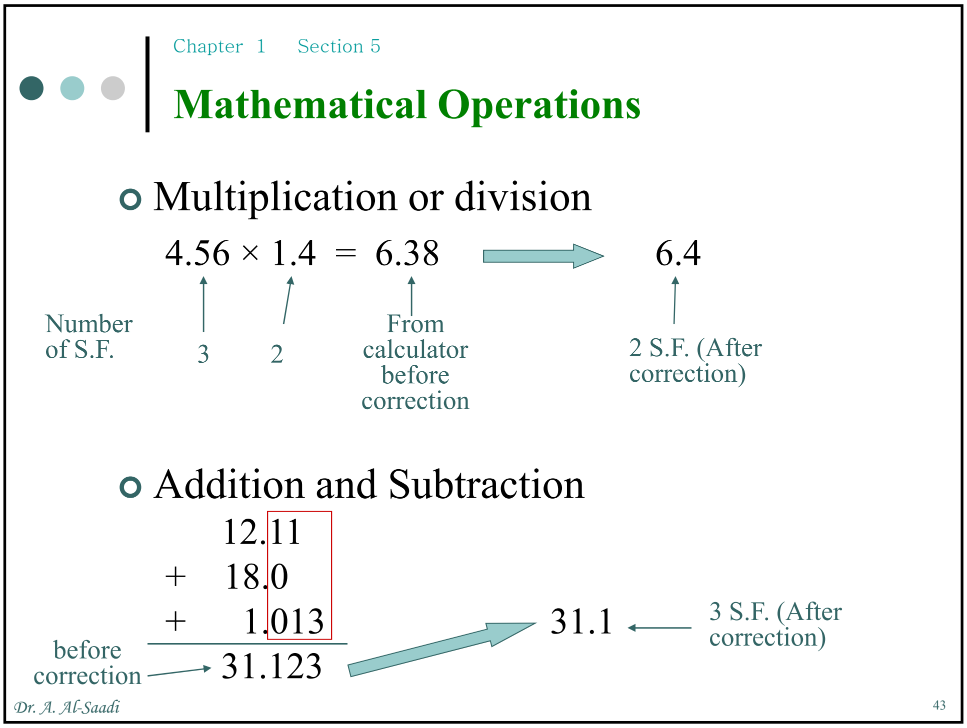
Exercise

- How many significant figures are there in the following numbers?

5.3004 m

0.120×10^{10} cm^3

250.0 kg



Mathematical Operations

- Multiplication or division

$$4.56 \times 1.4 = 6.38 \Rightarrow 6.4$$

Number of S.F.: 3, 2

6.38: From calculator before correction

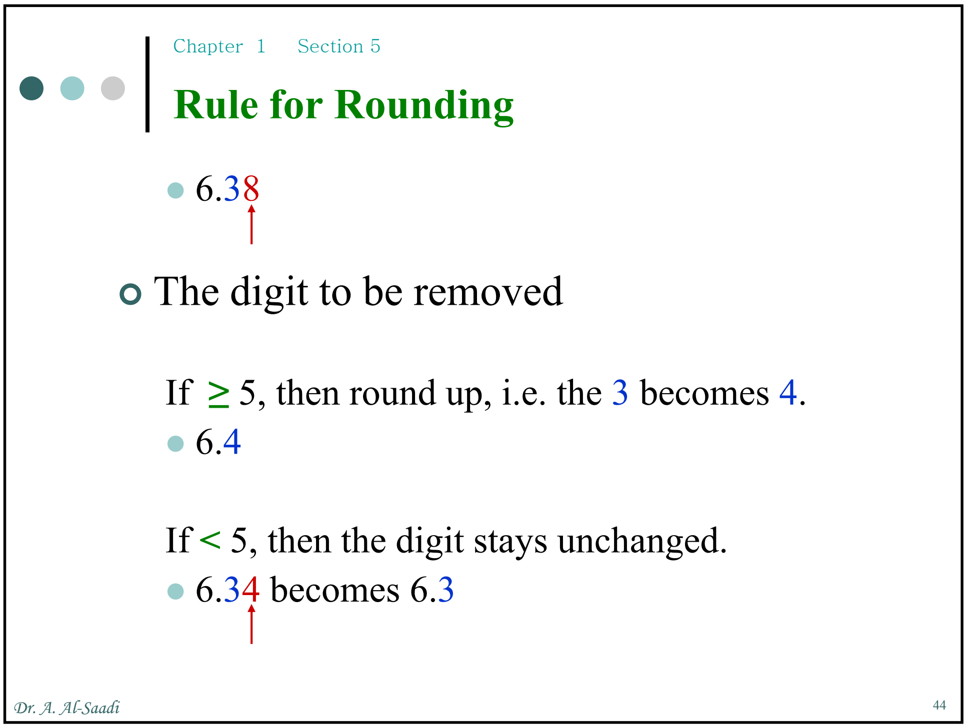
6.4: 2 S.F. (After correction)

- Addition and Subtraction

$$\begin{array}{rl} & 12.11 \\ + & 18.0 \\ + & 1.013 \\ \hline & 31.123 \end{array} \Rightarrow 31.1$$

31.123: before correction

31.1: 3 S.F. (After correction)



Rule for Rounding

- 6.38

- The digit to be removed

If ≥ 5, then round up, i.e. the 3 becomes 4.

- 6.4

If < 5, then the digit stays unchanged.

- 6.34 becomes 6.3

Chapter 1 Section 5

Exercises

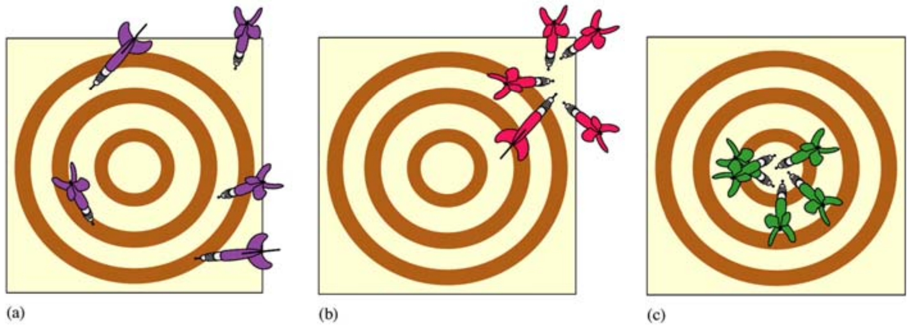
- Perform the following mathematical operation and express the result to the correct number of significant figures:

$$\frac{0.102\times 0.0821\times 273}{1.01}$$

Rounding off should be carried out for the final answer and NOT to the intermediate answers. However, you must keep track of the significant figures in the intermediate steps.

Chapter 1 Section 5

Precision and Accuracy



Random error

Neither precise nor accurate

Poor technique

Systematic error

Precise but not accurate *"reproducible"*

Good technique but needs calibration

Accurate and precise

Good technique

Precision and Accuracy

- Accuracy: Agreement of a particular value (measurement) with the true value.

- Precision: Agreement among several values (measurements), not necessarily agreeing with the true value.

Precision and Accuracy

Three students are asked to determine the mass of an aspirin tablet.

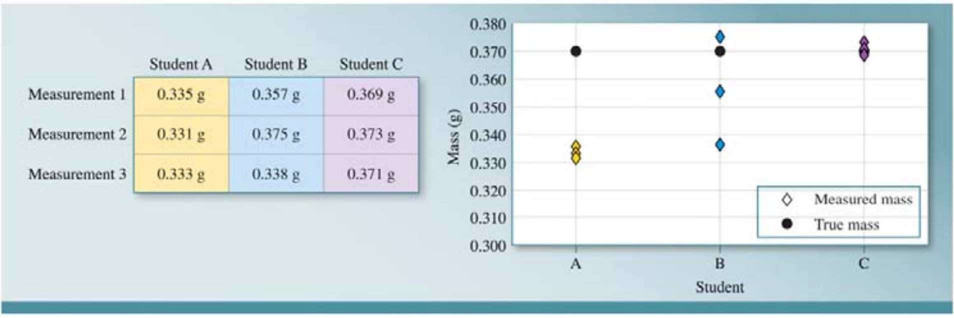
	Student A	Student B	Student C
	0.335 g	0.357 g	0.369 g
	0.331 g	0.375 g	0.373 g
	0.333 g	0.338 g	0.371 g
Average:	0.333 g	0.357 g	0.371 g

- True mass was 0.370 grams

Chapter 1 Section 5

Precision and Accuracy

Three students are asked to determine the mass of an aspirin tablet.



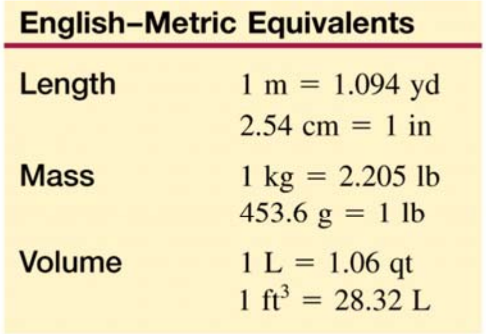
	Student A	Student B	Student C
Measurement 1	0.335 g	0.357 g	0.369 g
Measurement 2	0.331 g	0.375 g	0.373 g
Measurement 3	0.333 g	0.338 g	0.371 g

True mass was 0.370 grams

Chapter 1 Section 6

Conversion Factors

- Used to convert from one unit to another.



English–Metric Equivalents	
Length	1 m = 1.094 yd
	2.54 cm = 1 in
Mass	1 kg = 2.205 lb
	453.6 g = 1 lb
Volume	1 L = 1.06 qt
	1 ft^3 = 28.32 L

Example 1:

How many centimeters are in 25.5 inches (in)?

$$25.5\ \cancel{\text{in}} \times \frac{2.54\ \text{cm}}{1\ \cancel{\text{in}}} = 64.8\ \text{cm}$$

Example 2:

How many inches are in 25.5 centimeters?

$$25.5\ \cancel{\text{cm}} \times \frac{1\ \text{in}}{2.54\ \cancel{\text{cm}}} = 10.0\ \text{in}$$

Dimensional Analysis – Tracking Units

$1\text{L} = 1000\text{ ml}$

$1\text{ ml} = 0.001\text{ L}$

$= 1\times 10^{-3}\text{ L}$

Example 3:

How many ml are in 1.63 L?

Which direction you choose?

?

$$1.63\text{ L}\times\frac{1\text{ L}}{1000\text{ ml}} = 0.00163\frac{\text{L}^2}{\text{ml}}$$

$$1.63\text{ L}\times\frac{1000\text{ ml}}{1\text{ L}} = 1.63\times10^3\text{ ml}$$

Solving Problems Using Dimensional Analysis

- How many centimeters are in 0.25 megameters?

$$0.25\ \cancel{\text{megameters}}\times\frac{1\times10^6\text{ m}}{1\ \cancel{\text{megameters}}} = 0.25\times10^6\text{ m}$$

$$0.25\times10^6\ \cancel{\text{m}}\times\frac{100\text{ cm}}{1\ \cancel{\text{meters}}} = 0.25\times10^8\text{ cm}$$

$$= 2.5\times10^7\text{ cm}$$

$$= 25.\times10^6\text{ cm}$$

Chapter 1 Section 6

Solving Problems Using Dimensional Analysis

- The Food and Drug Administration (FDA) recommends that dietary sodium intake be no more than 2400 mg per day.
  What is this mass in pounds (lb), if 1 lb = 453.6 g?

$$2400\,\cancel{\text{mg}} \times \frac{1\,\cancel{\text{g}}}{1000\,\cancel{\text{mg}}} \times \frac{1\,\text{lb}}{453.6\,\cancel{\text{g}}} = 5.3 \times 10^{-3}\,\text{lb}$$

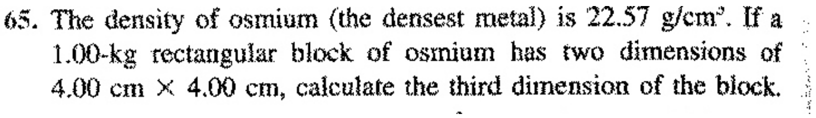
Exercise

- Perform the following mathematical operation and express the result to the correct number of significant figures:

$$\frac{2.526}{3.1} + \frac{0.470}{0.623} + \frac{80.705}{0.4326}$$

Rounding off should be carried out for the final answer and NOT to the intermediate answers. However, you must keep track of the significant figures in the intermediate steps.

Exercise



65. The density of osmium (the densest metal) is 22.57 g/cm³. If a 1.00-kg rectangular block of osmium has two dimensions of 4.00 cm × 4.00 cm, calculate the third dimension of the block.

Methodology to solve such problems:

- Start with the quantity given in the question.
- Use possible conversion factors to convert the unit of the given quantity in the question to the desired/needed unit.