Chapter 1

Fundamental Concepts and Units of Measurement

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1.1 Chemistry is important…

- Matter – anything that takes up space and has mass
- Chemistry – the science that deals with matter (the structure and properties of matter and the transformations from one form of matter to another)

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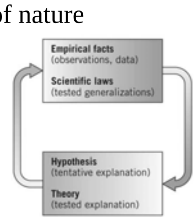
1.2 The Scientific Method

- Helps us build models of nature
  - Observations
  - Facts (data)
  - Hypothesis
  - Testing

  …

  …
  - Theory – a tested explanation of the behavior of nature



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The Scientific Method- Evaluating The Data

A ***theory*** is an explanation (based on well-tested, internally consistent experimental results) about why the phenomenon may occur

- It should explain currently available data
- It should be as simple as possible
- It should clearly show underlying connections
- It should accurately predict future behaviors

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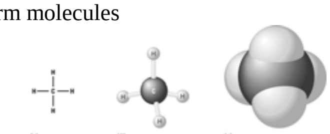
Atomic Theory – Model of Nature

- Most significant theoretical model of nature ever formulated
  - All chemical substances are composed of atoms.
  - Individual atoms combine in diverse ways to form molecules



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1.3 Matter is composed of elements, compounds, and mixtures

Elements (and atoms)

- Substances that cannot be decomposed into simpler materials by chemical reactions
- Identified by chemical symbol
- 90 naturally occurring, 27 man-made
- Check out capitalization

TABLE 1.2 Elements That Have Symbols Derived from Their Latin Names

Element	Symbol	Latin Name	Element	Symbol	Latin Name
Sodium	Na	Natrium	Gold	Au	Aurum
Potassium	K	Kalium	Mercury	Hg	Hydrargyrum
Iron	Fe	Ferrum	Antimony	Sb	Stibium
Copper	Cu	Cuprum	Tin	Sn	Stannum
Silver	Ag	Argentum	Lead	Pb	Plumbum

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Compounds

A substance formed from two or more different elements.

- Always combine in fixed proportions by mass
- For example, water (H_2O). Ratio of the masses of O:H is always 8:1
- Properties of compounds will usually differ from the properties of the elements they are formed from.

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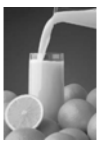

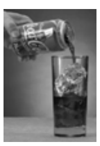

Mixtures

- Pure substances: elements and compounds
- A mixture contains two or more chemical substances.
  - Homogeneous – same properties throughout
  - Heterogeneous – when a mixture has two or more regions that differ in properties

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Classifications of Matter

	Sand	Ice (H_2O)	Flour	Table Salt (NaCl)
Pure		✓		✓
Element				
Compound		✓		✓
Molecule		✓		
Heterogeneous Mix	✓		✓	
Homogeneous Mix				

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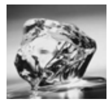
Physical and Chemical Changes

- Physical Property – observed without changing the chemical makeup of a substance, but can lead to a physical change.
- Chemical Property – also called a chemical reaction, a chemical change a substance undergoes
  - Evidence? Formation of a new solid, new liquid, new gas, temperature change, or an unexpected color change



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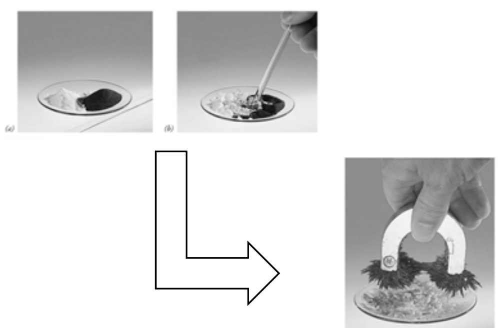

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Chemical Or Physical Change?

	Chemical	Physical
Magnesium burns when heated in a flame	✓	
Magnesium metal tarnishes in air	✓	
Magnesium metal melts at 922K		✓
Grape Kool-aid lightens when water is added		✓

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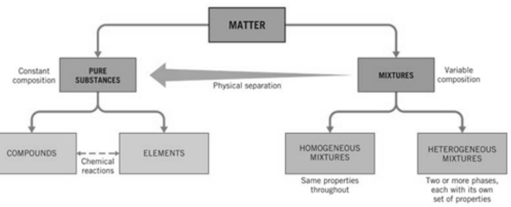
Matter Flow Chart



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1.4 Properties of Matter

- States of matter
  - Solids – have definite shapes and definite volume, not compressible
  - Liquids - have no definite shape, but do have a defined volume that remains constant from container to container, able to flow
  - Gases – have no definite shape or volume, molecules are very far apart, are highly compressible

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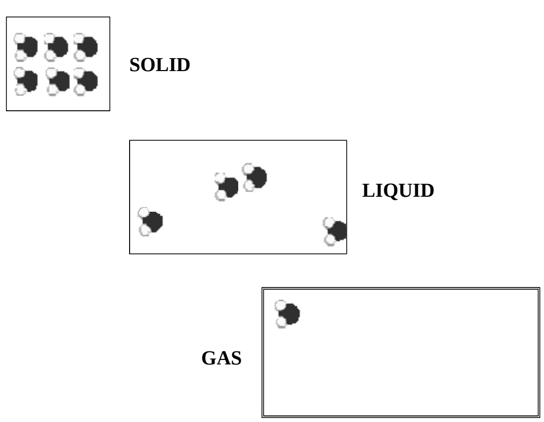

Other Classifications

- Extensive properties – a property that depends sample size
  - Volume, mass
- Intensive properties – a property that is independent of sample size
  - Color, electrical conductivity, melting points, etc…
  - Useful for identifying substances

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1.5 Measurements are essential

- Types of observations:
  - Qualitative = nonnumeric in nature
    - What is present?
  - Quantitative observations = numeric in nature, also called measurements
    - How much is present?

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Units

- A measurement requires two things
  - (1) A number – if less than 1 always include leading zero
    - 0.054 g ~~.054 g~~
  - (2) Units – don't forget the units!
- All measurements have some uncertainty.

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SI Units

- Systeme International (SI) expresses each fundamental physical quantity in decimally related units

TABLE 1.3 The SI Base Units

Measurement	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

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Magnitude (aka decimal multipliers)

Prefix	Symbol	Multiplication Factor	Absolute Factor
mega	M	10^6	10^6
kilo	k	10^3	10^3
centi	c	10^{-2}	10^2
milli	m	10^{-3}	10^3
micro	μ	10^{-6}	10^6
nano	n	10^{-9}	10^9

- You need to memorize these.
- We will use them often.
- I recommend learning the absolute factor --- less memorizing involved.

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Common Measurements in Chemistry

- Size – defined by length, area and volume
- Mass (m) – the certain quantity of matter
- Time (t) – to determine how long it takes for a chemical transformation to take place
- Temperature (T) – determines the direction of heat flow

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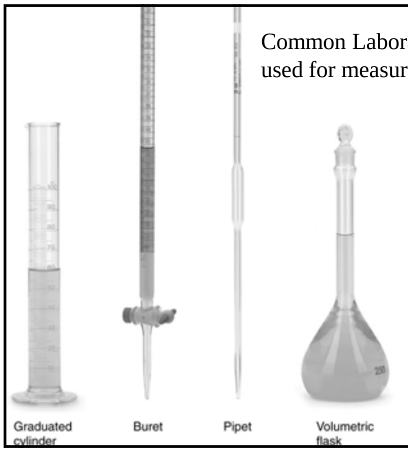
Common Laboratory glassware used for measuring volumes.



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Temperature

Three scales commonly used:

1. Fahrenheit (°F)
   – The boiling point of water is set at 212°F and the freezing point at 32°F.
2. Centigrade or Celsius (°C)
   – The boiling point of water is set at 100°C and the freezing point at 0°C.
3. Kelvin (K) (also called Absolute scale)
   – Has the same divisions as the Celsius scale, but a different zero point -273.15°C = 0 K
   – -273.15°C and 0 K are referred to as absolute zero…WHY?

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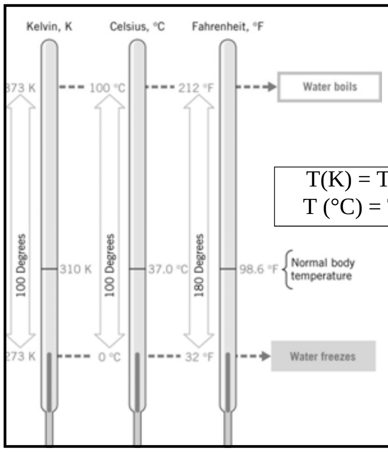

$$T(K) = T(°C) + 273.15$$
$$T(°C) = T(K) - 273.15$$

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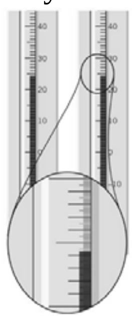
1.6 Measurements always contain some uncertainty

- All measurements have some uncertainty
- Error – the difference between the true value and the measurement.
- Error arises when a person estimates readings from a scale.
- The error is dependent on the measuring device and of course, the "measurer".



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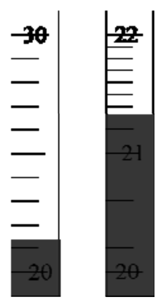
Significant Figures (sig. figs.)

- When making a measurement, all digits in a measurement **up to and including the first estimated digit** are recorded.
- These digits are called significant figures.



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Significant Figures

This digit has some uncertainty.

21.3ºC

These two digits are known for sure.

This digit has some uncertainty.

21.32ºC

These three digits are known for sure.

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Uncertainty comes from…

- Errors – inherent error due to the equipment or procedure
  - Changing volume due to thermal expansion or contraction (temperature changes)
  - Improperly calibrated equipment
  - procedural design allows variable measurements
- Mistakes – blunders that you know that you have made. Do not use these data
  - Spillage
  - Incomplete procedures
  - Reading scales incorrectly
  - Using the measuring device incorrectly

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Accuracy and Precision

- An accurate measurement is close to the true or correct value.
- A precise measurement is close to the average of a series of repeated measurements.
- When calibrated instruments are used properly, the greater the number of significant figures, the greater is the degree of precision for a given measurement.

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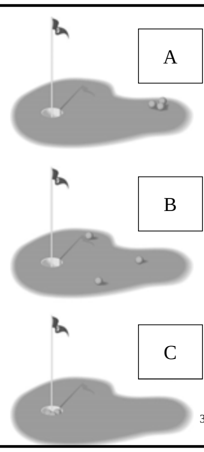
Accuracy and Precision

A =

B =

C =



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Measurements & Significant Figures

- What is a significant figure?
  - Digits that result from measurement such that only the digit farthest to the right is not known with certainty
- The greater the number of significant figures, the greater is the degree of precision.
- All digits (1 – 9) are significant.
- When are zeros significant?

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Zeros

1. Zeros between significant digits are significant. (3005)
2. Zeros to the left of the first nonzero digit are never counted as significant. (0.045 → 4.5×10^{-2})
3. Zeros to the right of a decimal point are always counted as significant. (4.500)
4. Zeros on the end of a number **WITHOUT** a decimal point are assumed not to be significant. (12000 → 1.2×10^{4})

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More on zeros

Indicate the uncertain digit and the number of significant units in the following measurements

1. 4.005 m
2. 4.500 m
3. 0.0023 m
4. 45,000 m

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SF and Multiplication/Division

- The number of significant figures in the answer should equal the number of significant figures in the **least precise** measurement, meaning the smallest number of SF.

$$\frac{3.14 \times 2.751}{0.64} = 13$$

$$\frac{(3\,\text{sig. figs.}) \times (4\,\text{sig. figs.})}{(2\,\text{sig. figs.})} = (2\,\text{sig. figs.})$$

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SF and Addition/Subtraction

- The answer should have the same number of decimal places as the quantity with the fewest number of decimal places

$$\begin{array}{r} 3.247 \\ 41.36 \\ +125.2 \\ \hline 169.8 \end{array}$$

125.2 ← This number has only 1 decimal place.

169.8 ← The answer has been rounded to 1 decimal place.

Let's look at the significant figure handout.

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1.7 Units can be converted

- Also called factor-label method or dimensional analysis
- We frequently need to convert a measurement from one unit to another.
- When multiplying numbers, we also multiply units and when dividing numbers, we also divide units
- Never forget units!
- Learn this now or you will regret it in Chapters 3 and 4.

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Some useful conversions

- The most useful are the multipliers: mega, kilo, centi, milli, micro, nano and pico
- Others can be found on the back cover of your book

TABLE 1.6 Some Useful Conversions

Measurement	English to Metric	Metric to English
Length	1 in. = 2.54 cm 1 yd = 0.9144 m 1 mi = 1.609 km	1 m = 39.37 in. 1 km = 0.6215 mi
Mass	1 lb = 453.6 g 1 oz = 28.35 g	1 kg = 2.205 lb
Volume	1 gal = 3.785 L 1 qt = 946.4 mL 1 oz (fluid) = 29.6 mL	1 L = 1.057 qt

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How to complete MOST chemistry calculations for this class

1. If there is a reaction going on, write down the balanced equation
2. Write down what you know with units.
3. Write down what you are looking for with the desired units.
4. Start with the measurement that only has one unit. Never start at the question marks!
   – 15 cm or 25 m/s
5. Use what you know and any unit conversions to get to the correct answer with the desired units.

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Learn by practice

- Convert 3.25 m to millimeters (mm).
- Convert 2500 m to feet.
- Convert 3.00 yd to inches.

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- Convert 1.25 kg to centigrams.
- Convert 20.2 miles/gallon to km/L.
- Convert 55 miles/hr to m/s.

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1.8 Density is a useful intensive property

- <u>Density</u> (d) – the mass per unit volume of a substance

$$d = \frac{m}{V} \qquad \text{m = mass, V = volume}$$

- Common Units – g/cm^3, g/mL and g/L

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- The density of seawater is 1.03 g/mL. What mass of seawater would fill a vessel to a volume of 225 mL?
- An ocean – dwelling dinosaur has an estimated body volume of 1.38×10^6 cm^3. The animal's live mass was estimated at 1.24×10^6g. What is its density?

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Additional notes on density

- The density of a solid or liquid is constant at a given temperature.
- Almost always, density decreases with increasing temperature.
- Water is the most common exception. (density increases from 0°C to 4 °C)
- The density of water is 1.00 g/mL

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