# Matter & Measurement

# Chemistry

Chemistry: science that deals with materials (matters) and their changes.

**Central Science**

All sciences are connected to chemistry.

# Changes

Chemical change (chemical reaction):

substance(s) are used up (disappear) → other forms

**burning a paper or a leaf changes color in the fall**

Physical change: identities of the substances do not change.
(change of state)

**evaporation of water or melting**

# Physical and Chemical Properties

Physical Properties: a directly observable characteristic of a substance exhibited as long as no chemical change occurs.

Color, Odor, Volume, State, Density, Melting and boiling point.

Chemical Properties: Ability to chemical changes.
(forming a new substance(s))

Burning wood – rusting of the steel

# Scientific method

1. Fact: is a statement based on direct experience (observation).

State the problem and collect data.

Qualitative observation: Water is liquid.

Quantitative observation: water boils at 100°C.

$\downarrow$

Measurement

# Scientific method

2. Hypothesis: is statement that is proposed to explain the observation.

(without actual proof)

3. Experiment: performing some tests to prove hypothesis.

(find a proof)

# Scientific method

4. Theory (model): is a set of tested hypotheses.

We have a stronger belief in it because of more evidence supports it.

Law: is a summery of observed behavior.

We formulate this observations.

Serendipity: Observation by chance

# Theory verses Law

**Scientific theory :** Very general, "***Why it happens,"*** often includes many "Laws"

**Scientific law** : Very specific, "***What will happen"*** often expressed in mathematical equations.

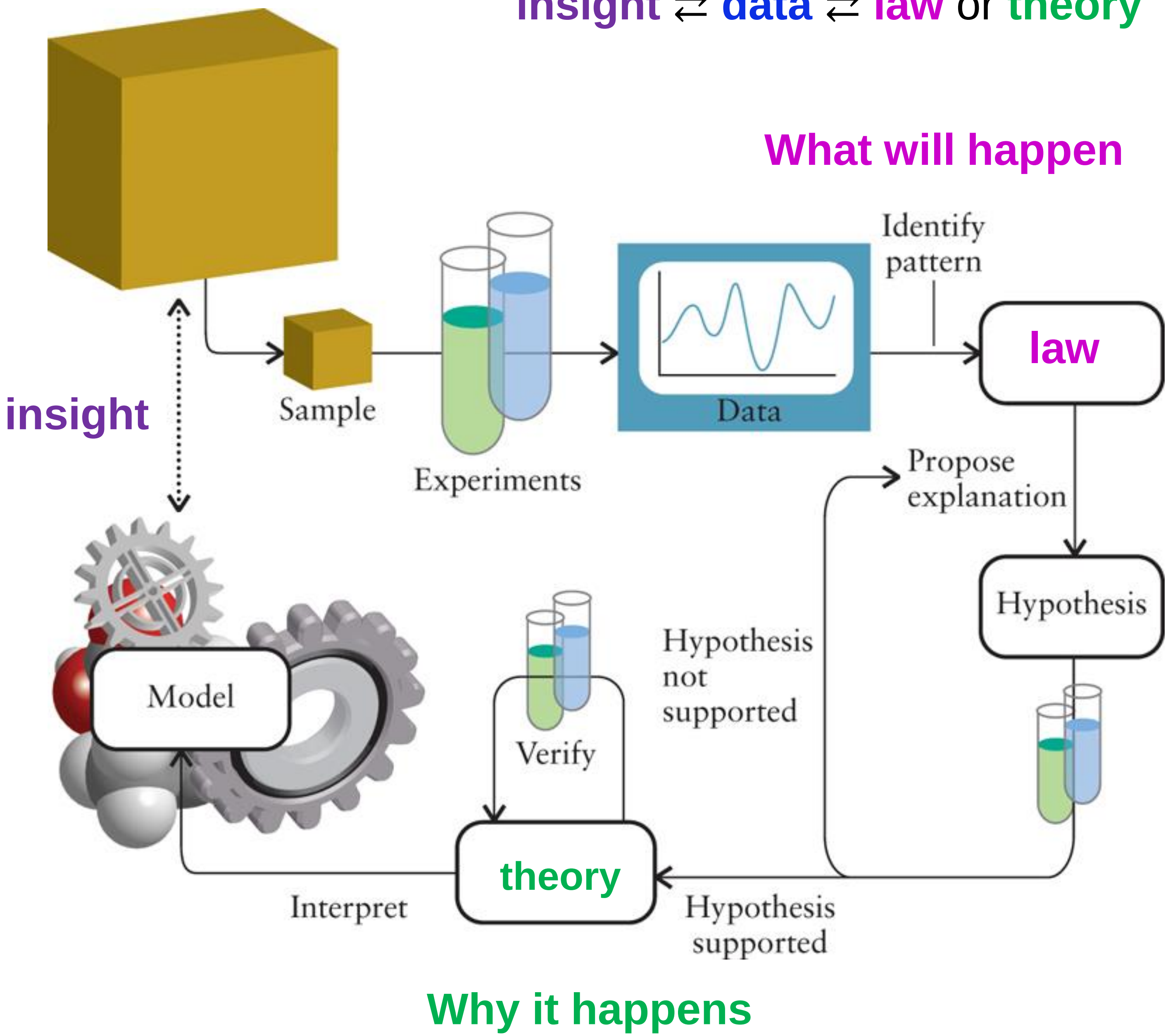
A **Heat** Theory and (many) **Heat** Law(s),

The **kinetic molecular theory of motion** is the idea that particles move faster ***because*** they are hotter.

**Charles Law** will tell us the consequences of the heating process, the exact, numerical value, what ***happens*** to pressure after heating an object.

How science progresses

insight ⇄ data ⇄ law or theory



# Matter

Matter: has mass and takes space.

# Matter & Energy

Are they related?

All objects **move**, except at the hypothetical temperature zero Kelvin.

A common **energy** formula: $E_{kinetic} = \frac{mv^2}{2}$ shows the relationship between **energy**, **mass** and **speed**.

We see in many energy equations this **mass** and **motion** relationship.

# Matter & Energy

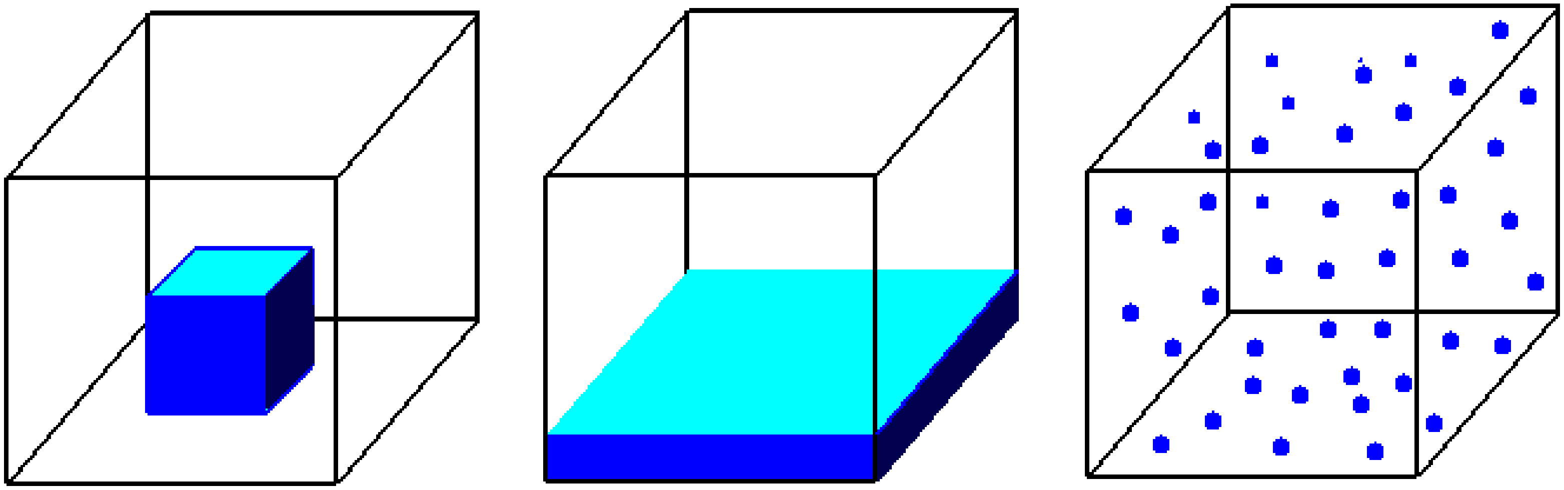
**Matter** has mass and volume.

3-states of matter

- a) **Solids**
- b) **Liquids**
- c) **Gases**

***Speed***

Atoms begin to move faster as temperature is increased.



Solid Liquid Gas

# Measurements

# Measurements

Measurement consists of two parts:

**Number - Unit**

2 pounds

Number

Unit

# Measurement and Units

Metric system or SI (International System of Units)

meter, liter, gram …

English system (use in the United States)

miles, gallons, pounds …

Advantages of SI: we have base unit for each kind of measurement

other units are related to the base unit by power of 10.

| Prefix (symbol) | Value |
|---|---|
| giga (G) | $10^9$ |
| mega (M) | $10^6$ |
| kilo (k) | $10^3$ |
| deci (d) | $10^{-1}$ |
| centi (c) | $10^{-2}$ |
| milli (m) | $10^{-3}$ |
| micro (μ) | $10^{-6}$ |
| nano (n) | $10^{-9}$ |

base unit of length: meter (m)

1 kilometer (km) = 1000 meter (m)
1 centimeter (cm) = 0.01 meter (m)
1 nanometer (nm) = $1 \times 10^{-9}$ meter (m)

base unit of mass: gram (g)

1 kilogram (kg) = 1000 gram (g)
1 milligram (mg) = 0.001 gram (g)

base unit of volume: liter (L)

1 milliliter (mL) = 0.001 liter (L)
1000 milliliter (mL) = 1 liter (L)

1 mL = 1 cc = $1 cm^3$

base unit of time: second (s)

60 seconds (s)= 1 minute (min)
60 minutes (min) = 1 hour (h)

# Tools (equipment) of measurement

Length: Meterstick or Ruler

Volume: Graduated cylinder, Pipette

Mass: Balance

# Temperature

english system $\longrightarrow$ Fahrenheit (°F)

metric system or SI $\longrightarrow$ Celsius or centigrade (°C)

$$°F = 1.8\ °C + 32$$

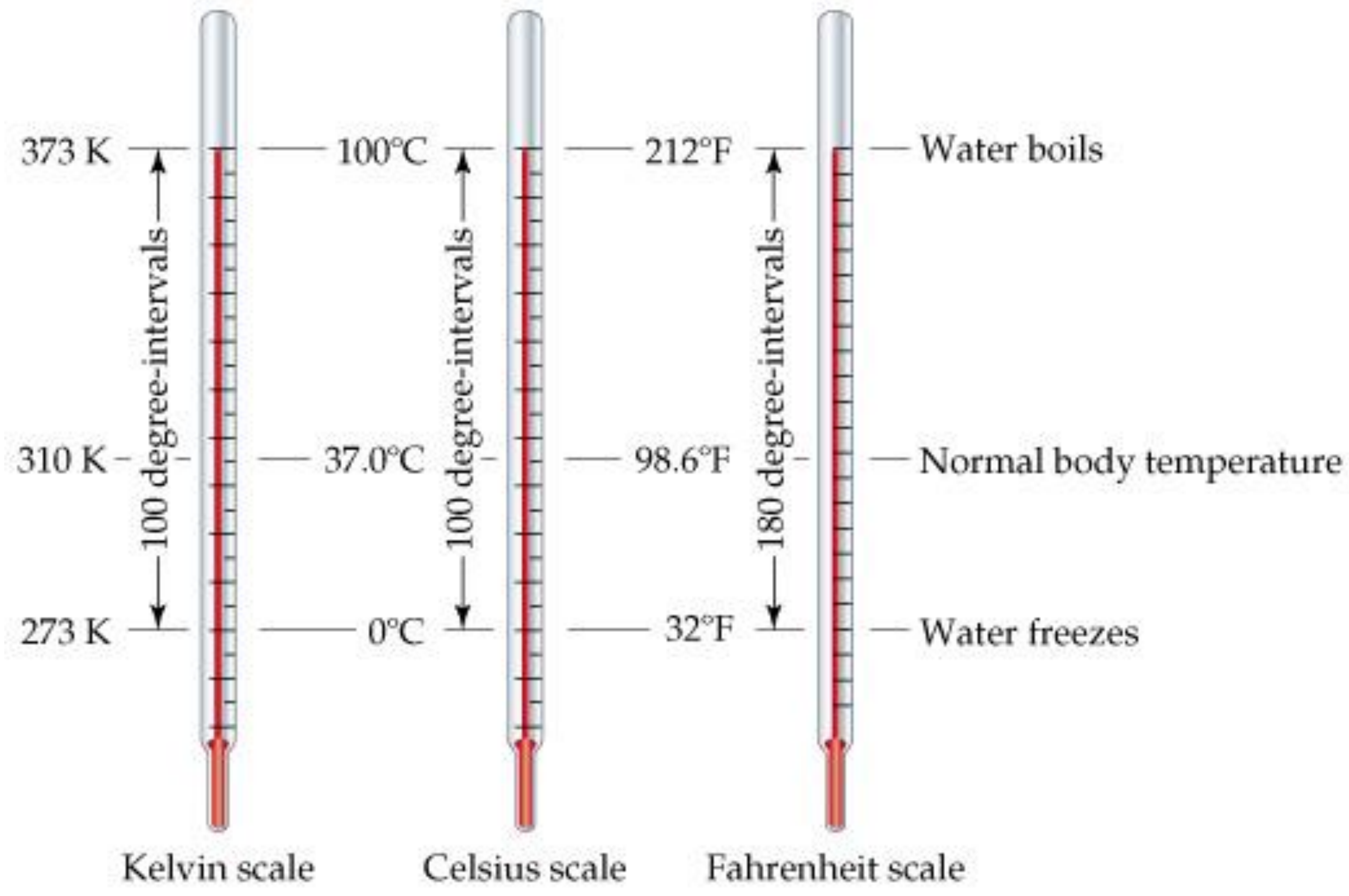
$$°C = \frac{°F - 32}{1.8}$$

Kelvin scale or absolute scale (K)

$$K = °C + 273$$

$$°C = K - 273$$

# Temperature



# Temperature

1. Size of degree is the same for Celsius and Kelvin scales.

2. Fahrenheit scale is smaller than others.

3. The zero points are different on all there scales.

# Scientific (exponential) notation

based on power of 10

$10000 = 1\times10^{4}$

$0.0001 = 1\times10^{-4}$

$4500000 = 4.5\times10^{6}$

$0.000078 = 7.8\times10^{-5}$

$94800 = 9.48\times10^{4}$

$0.0121 = 1.21\times10^{-2}$

Positive power: greater than 1

Negative power: Less than 1

# Scientific (exponential) notation

Power, Exponent

$$9.23 \times 10^{25}$$

Coefficient

Base

# Scientific (exponential) notation

$$(3.62 \times 10^{6})(7.43 \times 10^{3}) = 26.90 \times 10^{9} = 2.69 \times 10^{10}$$

$$\frac{(3.62 \times 10^{7})}{(1.35 \times 10^{5})} = 2.68 \times 10^{2}$$

Moving the decimal point to right $\longrightarrow$ Decreasing the power one point

Moving the decimal point to left $\longrightarrow$ Increasing the power one point

# Conversion of Units

Conversion Factor:

$$1 \text{ m} = 1000 \text{ mm}$$

Equivalence statement (Equality)

$$\frac{1 \text{ m}}{1000 \text{ mm}} \quad \text{or} \quad \frac{1000 \text{ mm}}{1 \text{ m}}$$

Conversion factor

Ratios of two parts of equality

# Conversion of Units

Factor-Label method (dimensional analysis):

$$36 \text{ m} = ? \text{ mm}$$

$$36 \text{ m} \times \text{conversion factor} = ? \text{ mm}$$

$$\frac{1 \text{ m}}{1000 \text{ mm}} \quad \text{or} \quad \frac{1000 \text{ mm}}{1 \text{ m}}$$

$$36 \cancel{\text{m}} \times \frac{1000 \text{ mm}}{1 \cancel{\text{m}}} = 36000 \text{ mm}$$

# Conversion of Units

## Factor-Label method

25kg = ? lb

$$25\cancel{kg} \times \frac{2.205\ lb}{1\ \cancel{kg}} = 55\ lb$$

78 mile = ? km

$$78\cancel{mi} \times \frac{1.609\ km}{1\ \cancel{mi}} = 130\ km$$

45 m/hr = ? in/min

$$45\ \frac{\cancel{m}}{\cancel{hr}} \times \frac{39.37\ in}{1\ \cancel{m}} \times \frac{1\ \cancel{hr}}{60min} = 30.\ in/min$$

# A note of good practice to students

$$\left(\frac{2}{x}\right)^2 = \frac{4}{x^2}$$

Which can also be written as $4x^{-2}$

$$\left(\frac{1\ cm}{10^{-2}m}\right)^{-3} = \left(\frac{10^{-2}m}{1\ cm}\right)^{3} = \frac{10^{-6}m^3}{1cm^3}$$

$$\frac{\frac{a}{b}}{\frac{c}{d}} = \frac{a}{b} \div \frac{c}{d} = \frac{a}{b} \times \frac{d}{c} = \frac{ad}{bc}$$

Example: Express an acceleration of 9.81 $m \cdot s^{-2}$ in kilometers per hour squared ($km \cdot hr^{-2}$).

1 km = 1000 m
1 hr = 3600 s

$$\frac{1\ \text{km}}{1000\ \text{m}} \times \frac{9.81\ \text{m}}{\text{s}^2} \times \left(\frac{3600\ \text{s}}{1\ \text{hr}}\right)^2 = \frac{1.27 \times 10^5\ \text{km}}{\text{hr}^2}$$

Our measured number with the least SF is 9.81 or 3SF

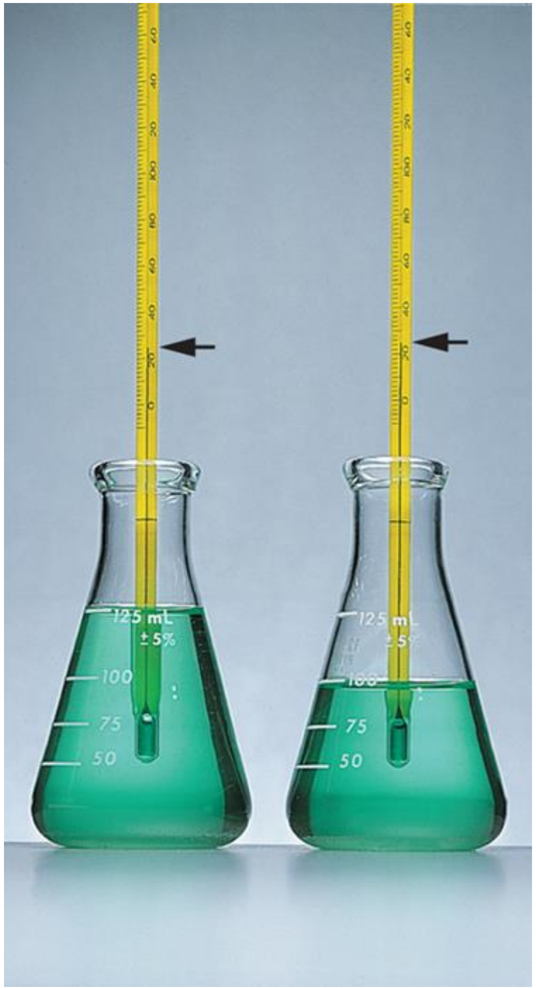
# Extensive & Intensive Properties

**Extensive properties** are physical properties that ***depend*** on the ***quantity*** of matter (*n* atoms):
*Volume* & *Mass*

Cutting coins in half will give you half the number of atoms

# Extensive & Intensive Properties

**Intensive properties** are ***independent*** of the ***quantity*** of matter: Density & Temperature




**Same temperature**
different volumes.

# Density and Specific gravity

density: amount of mass present in a given volume.

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

d: density (g/mL or g/cm$^3$)  m: mass  V: volume

The density of ice is less than the density of liquid water, so the ice floats on top of the water.

Salad oil is less dense than vinegar.

# Density and Specific gravity

**Table 2.8** Densities of Various Common Substances at 20 °C

| Substance | Physical State | Density ($g/cm^3$) |
|---|---|---|
| oxygen | gas | 0.00133* |
| hydrogen | gas | 0.000084* |
| ethanol | liquid | 0.785 |
| benzene | liquid | 0.880 |
| water | liquid | 1.000 |
| magnesium | solid | 1.74 |
| salt (sodium chloride) | solid | 2.16 |
| aluminum | solid | 2.70 |
| iron | solid | 7.87 |
| copper | solid | 8.96 |
| silver | solid | 10.5 |
| lead | solid | 11.34 |
| mercury | liquid | 13.6 |
| gold | solid | 19.32 |

*At 1 atmosphere pressure

Gas = low density

Liquids: close to 1 $g/cm^3$, 1 g/mL

Metals: various heavy densities.

# Density Examples

**Example 1.** A gas fills a volume of 1200. mL and has a mass of 1.60 g. What is the density of the gas?

$$d = \frac{m}{V} = \frac{1.60\ g}{1200.\ mL} = 0.00133\ g/mL$$

**Example 2.** A cube of pure silver measures 2.0 cm on each side. The density of silver is 10.5 g/cm$^3$. What is the mass of the cube?

$$V = L \times H \times W = 2.0\ cm \times 2.0\ cm \times 2.0\ cm = 8.0\ cm^3$$

$$m = d \times V = 8.0\ cm^3 \times 10.5\ g/cm^3 = 84.\ g$$

# Density Examples

**Example 3:** The density of air is $1.25 \times 10^{-3}$ g/cm$^3$. What is the mass of air in a room that is 5.00 meters long, 4.00 meters wide and 2.2 meters high?

$$V = L \times H \times W$$

$V = 5.0$ m x $4.0$ m x $2.2$ m = $44$ m$^3$ Hmm, not so helpful.

$V = 500.$ cm x $400.$ cm x $220$ cm = $44000000$ cm$^3$

$$d = \frac{m}{V} \qquad m = d \times V$$

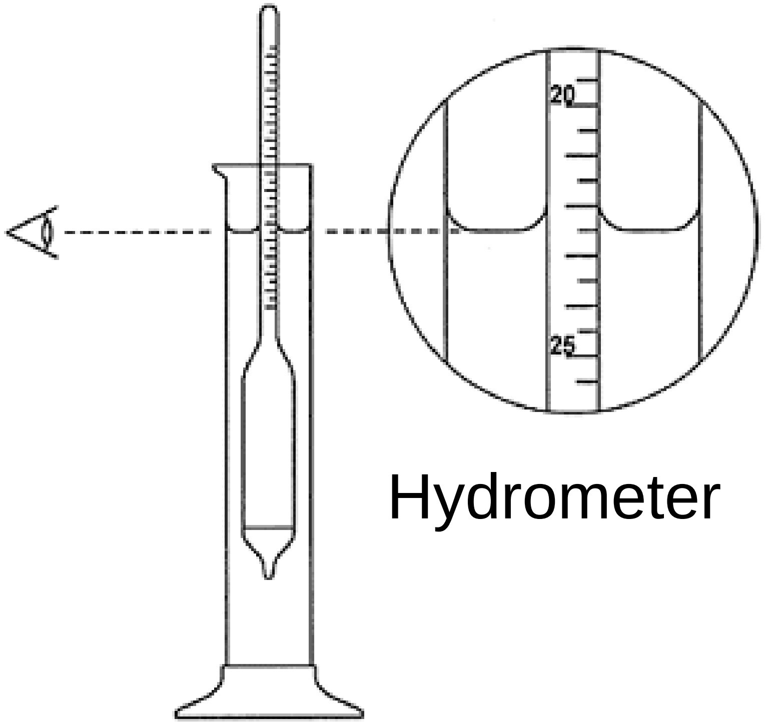
$m = (4.4 \times 10^7 \text{ cm}^3) \times (1.25 \times 10^{-3} \text{ g/cm}^3) = 55000$ g or $55$ kg

# Density and Specific gravity

## Specific gravity:

$$SG = \frac{d_{substance}}{d_{water}}$$

No units (dimensionless)




Hydrometer

# Significant Figures

Exact numbers: we do not use  a measuring devise.

(Counting numbers)

Number of students in class, 1m = 100cm

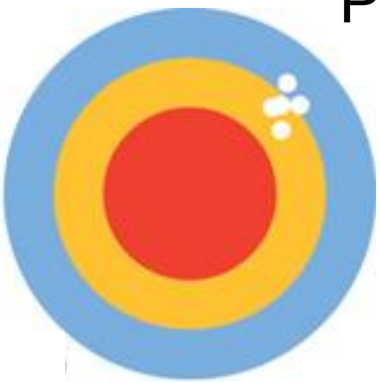
Inexact numbers: we use  a measuring devise.

(measuring numbers)

Temperature of room, mass of table

# Precise verses Accurate

Precise &
Accurate



Precise

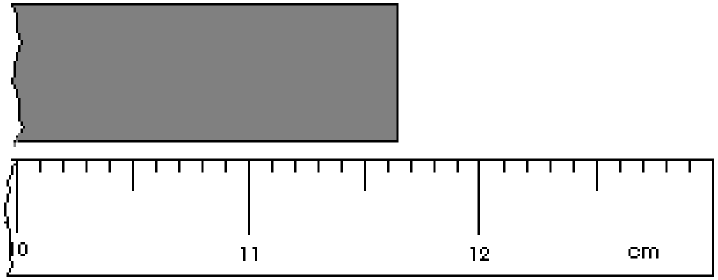
Accurate

Neither
precise nor
accurate

# Significant Figures

We always have errors in measurement: Personal and instrumental errors.

All measurements need an estimate.




between 11.6 and 11.7 $\longrightarrow$ 11.62 or 11.63 or 11.67 or …

# Significant Figures

Certain numbers: 11.6

Uncertain number: 11.66

(estimated digit - only the last digit)

Significant Figures: all numbers recorded in a measurement.

(certain and uncertain)

We only write the **correct number** of digits in our reports, these are called **significant figures**.

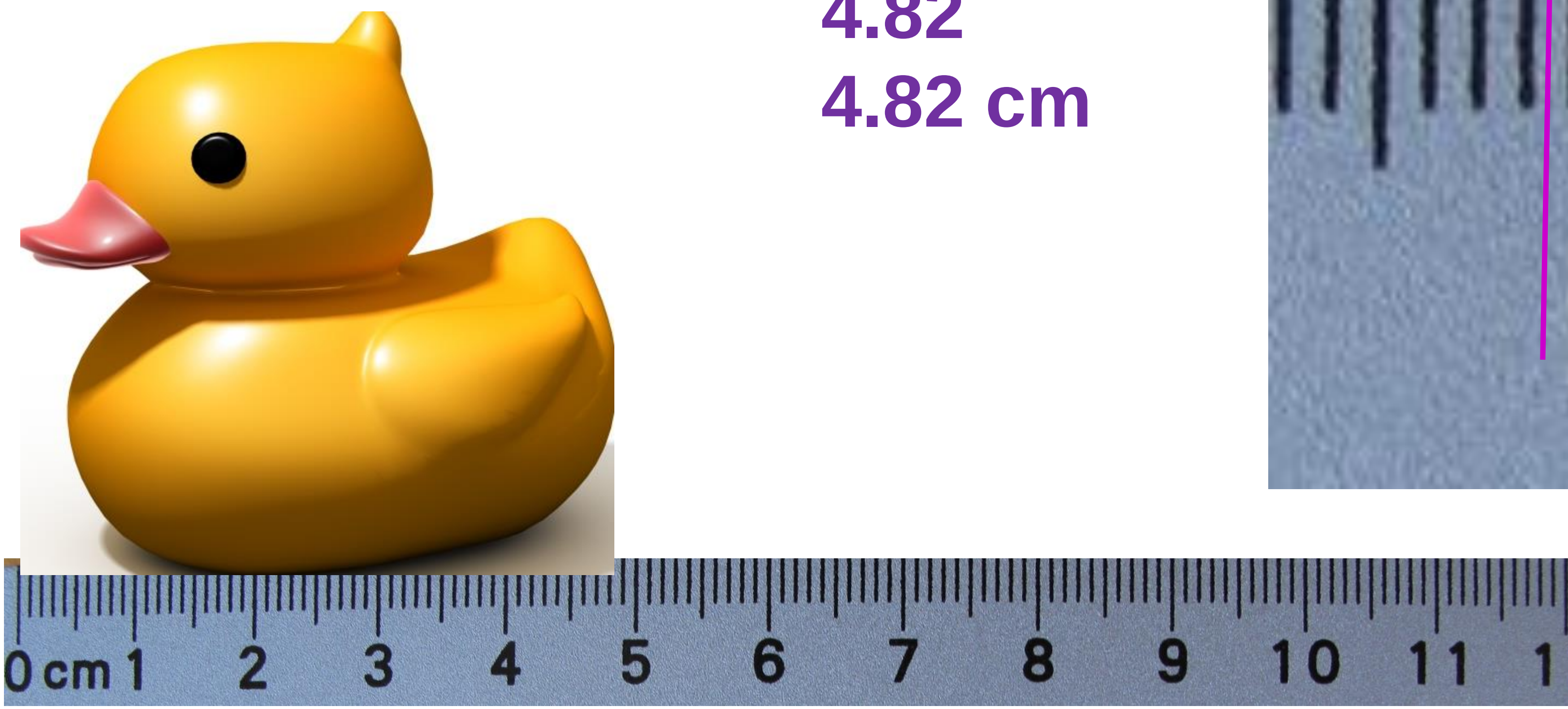
When we report, we show uncertainty with ± 11.66 ± 0.01

# Recording Measurements

When using a measuring tool

1. Write all the digits you see
2. Make one guess
3. Add units

**4.8**
**4.82**
**4.82 cm**

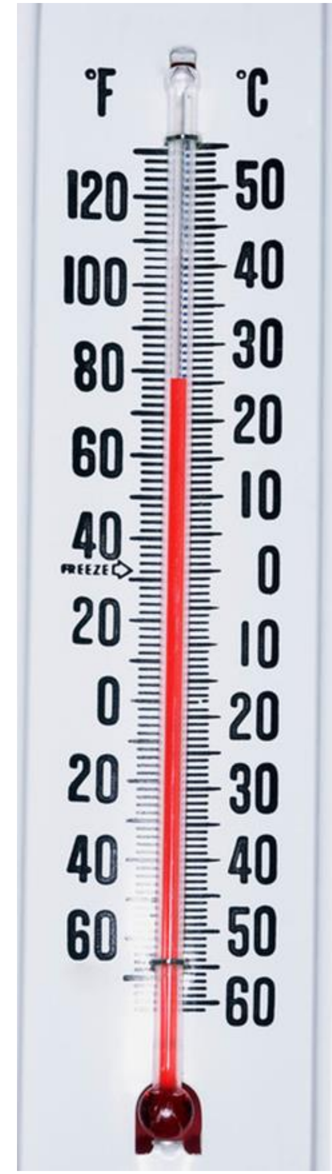



# Recording Measurements

When using a measuring tool

1. Write all the digits you see
2. Make one guess
3. Add units


**25**
**25.7**
**25.7 ℃**



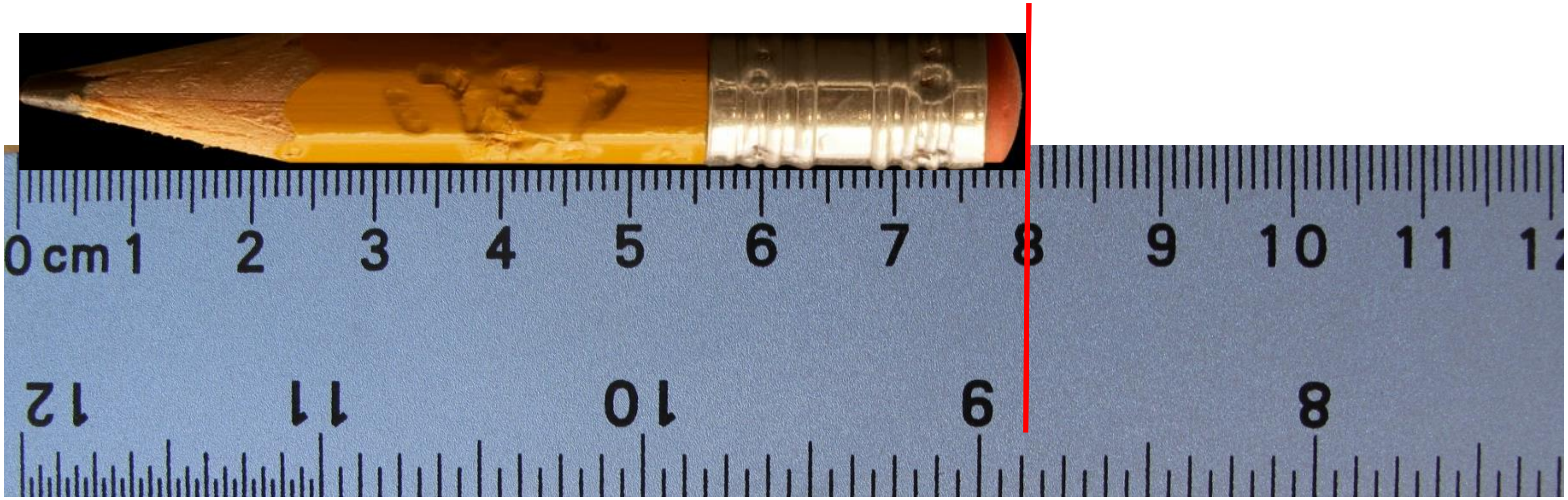
# Recording Measurements

When using a measuring tool

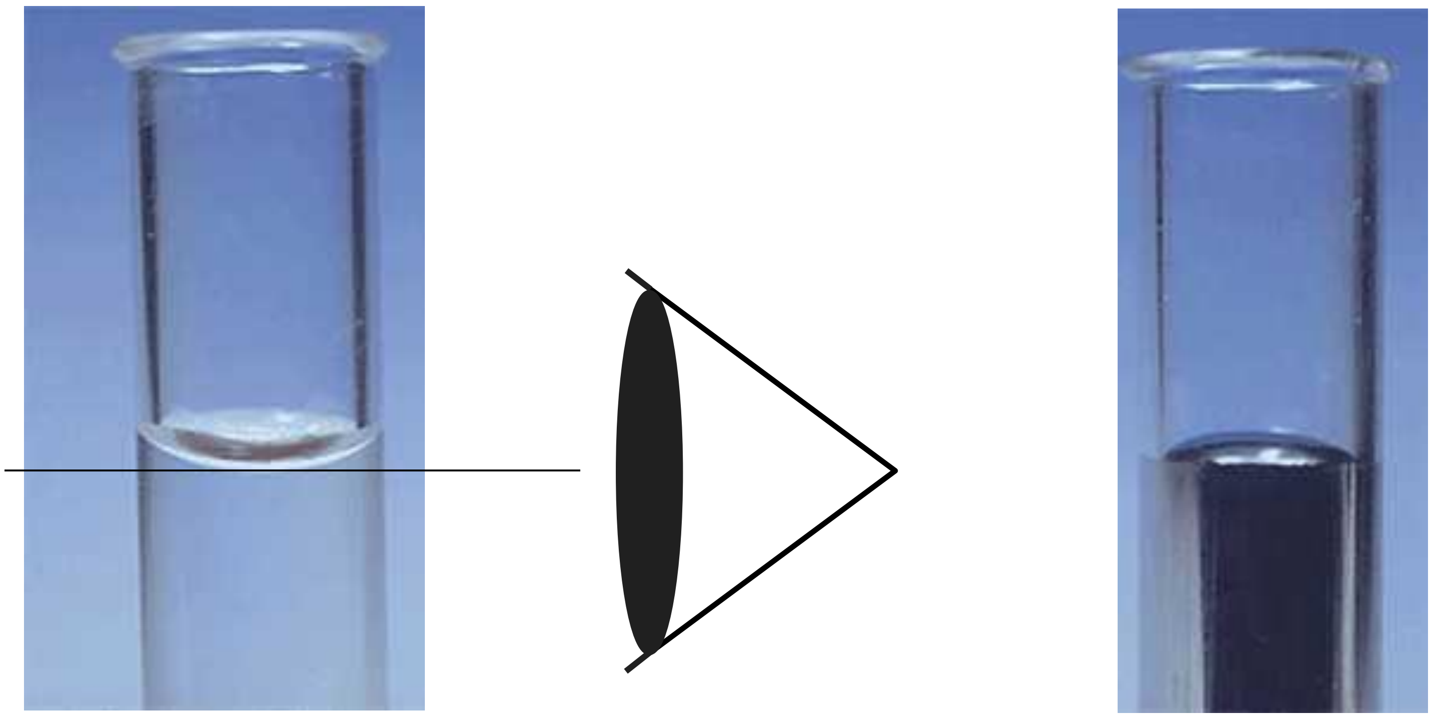
1. Write all the digits you see
2. Make one guess
3. Add units

**8.0**
**8.00**
**8.00 cm**

This number has 3 digits



**Reading liquid**: get eye level with the Bottom of the meniscus



Adhesion

Cohesion

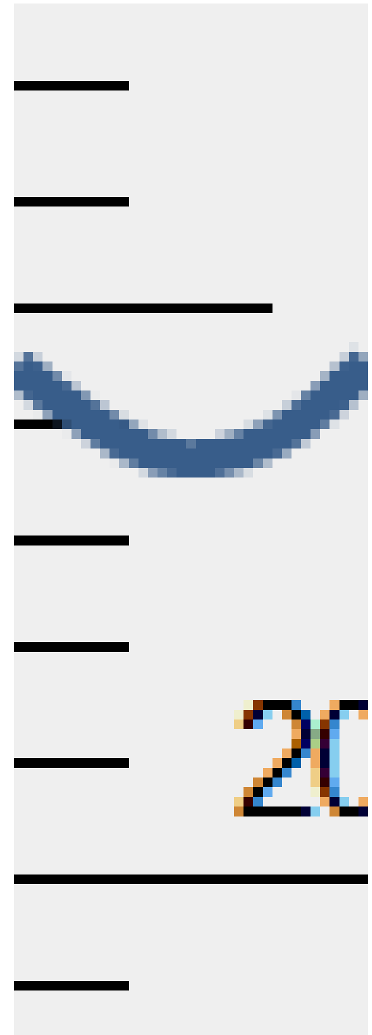
**Reading liquid**: get eye level with the Bottom of the meniscus

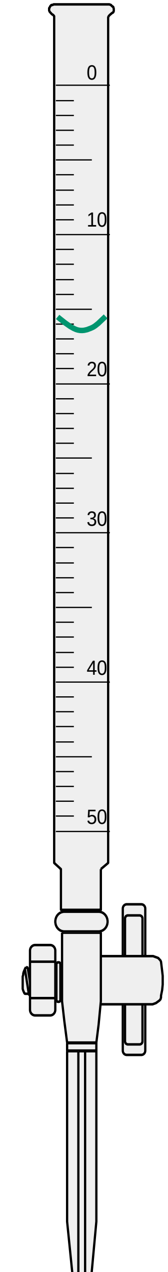
What would the volume reading be?

What do you see: 16

What is your guess: 16.4

Final: 16.4 mL





# Significant Figures rules

1. Nonzero digits count as significant figures. 297.32 ⟶ 5 S.F.

2. Zeros:

a) Zeros at the beginning of numbers do not count as S.F. (Leading zeros).

0.0031 ⟶ 2 S.F.

b) Zeros between two nonzero digits count as S.F. (Captive zeros).

600067 ⟶ 6 S.F.

c) Zeros at the end of numbers (Trailing zeros):

If there is a decimal point, count as S.F. 2.800 ⟶ 4 S.F.

If there is not a decimal point, do not count as S.F.

2800 ⟶ 2 S.F.

# Rounding off

1. If the digit to be removed:

a) is less than 5, the preceding digit stays the same.

5.343 ⟶ 5.34 (2 decimal places)

5.343 ⟶ 5.3 (1 d.p.)

b) is equal to or greater than 5, the preceding digit is increased by 1.

6.456 ⟶ 6.46 (2 decimal places)

6.456 ⟶ 6.5 (1 d.p.)

2. We round off at the end of calculation.

# Significant Figures in calculation

1. **Multiplication or division:**

Number of significant figures in result = Smallest number of significant figures.

$$\underbrace{4.000}_{4\text{ S.F.}} \times \underbrace{560}_{2\text{ S.F.}} \times \underbrace{7001}_{4\text{ S.F.}} \times \underbrace{0.003}_{1\text{ S.F.}} = 47046.72 = \underbrace{50000}_{1\text{ S.F.}}$$

4 S.F. ← 8.600

2 S.F. ← 44000

$$\frac{8.600}{44000} = 0.000195454 = \underbrace{0.00020}_{2\text{ S.F.}}$$

# Significant Figures in calculation

## 2. **Addition or subtraction:**

Number of decimal places in result = Smallest number of decimal places.

$$57.93 + 0.05 - 0.230 + 4600 = \cancel{4657.75} = 4658$$

2 d.p. 2 d.p. 3 d.p. 0 d.p. 0 d.p.

$$710.0 - 0.0063 - 4098.1 + 4.63 = \cancel{-3383.4763} = -3383.5$$

1 d.p. 4 d.p. 1 d.p. 2 d.p. 1 d.p.

# Significant Figures in calculation

Stop at the least SF digit

More practice:

| | | |
|---|---|---|
| 1200 | 500. | 100 |
| +250 | + 0.00021 | - 98.5 |
| 1450 | 500.00021 | 001.5 |

Recorded Answers

| | | |
|---|---|---|
| 1500 | 500. | 0 |

The zero does not count since there is no decimal.

## Significant Figures in calculation

More practice:

| | | | |
|---|---|---|---|
| 6600 | 10.0 | 10 | 10 |
| +25.0 | - 9.85 | - 9.85 | -4.85 |
| 6625.0 | .15 | 00.15 | 5.15 |

Recorded Answers

| | | | |
|---|---|---|---|
| 6600 | .2 | 0 | 10 |

# Significant Figures in calculation

- Significant Figures in Mixed operations

$(1.7 \times 10^6 \div 2.63 \times 10^5) + 7.33 = ???$

**Step 1:** Divide the numbers in the parenthesis. How many sig figs?

$(\underline{6.4}63878327\ldots) + 7.33$

**Step 2:** Add the numbers. How many decimal places to keep?

$$\begin{array}{r|l} \underline{6.4} & 63878327\ldots \\ +\ 7.3 & 3 \\ \hline 13.7 & 938 \end{array}$$

**Step 3:** Round answer to the appropriate decimal place.

**13.8 or $1.38 \times 10^1$**