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Chapter 1
Chemistry: The central science

- The study of chemistry
- Classification of matter
- Scientific measurement
- The properties of matter
- Uncertainty in measurement
- Using units and solving problems

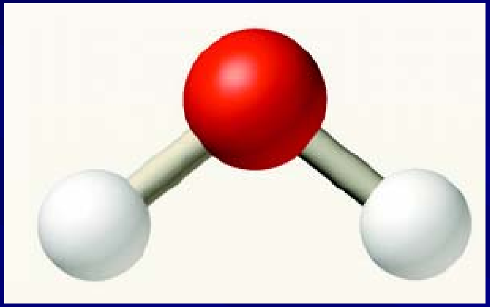
1.1 The Study of Chemistry

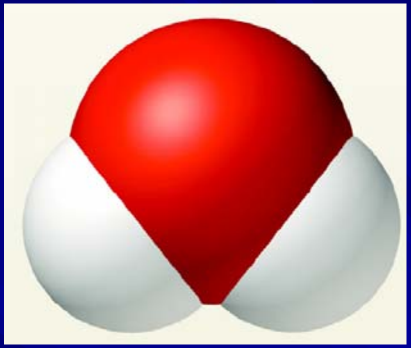
- *Chemistry*
  - the study of *matter* and the *changes* that matter undergoes
- *Matter*
  - anything that has mass and occupies space
- Matter is composed of atoms
- Atoms are found as individuals or molecules
- Atoms and molecules are connected by electrons
- Matter is composed of various types of atoms or molecules.
- Water is composed of O and H; H_2O
- An electric spark causes a mixture of O_2 and H_2 to explode forming H_2O.

- Chemistry you may already know
  - Familiar terms: molecules, atoms, and chemical reactions
- Familiar chemical formula: H_2O

- Molecules can be represented several different ways including *molecular formulas* and *molecular models.*

H_2O





  – Molecular models can be “ball-and-stick” or “space-fill.” Each element is represented by a particular color

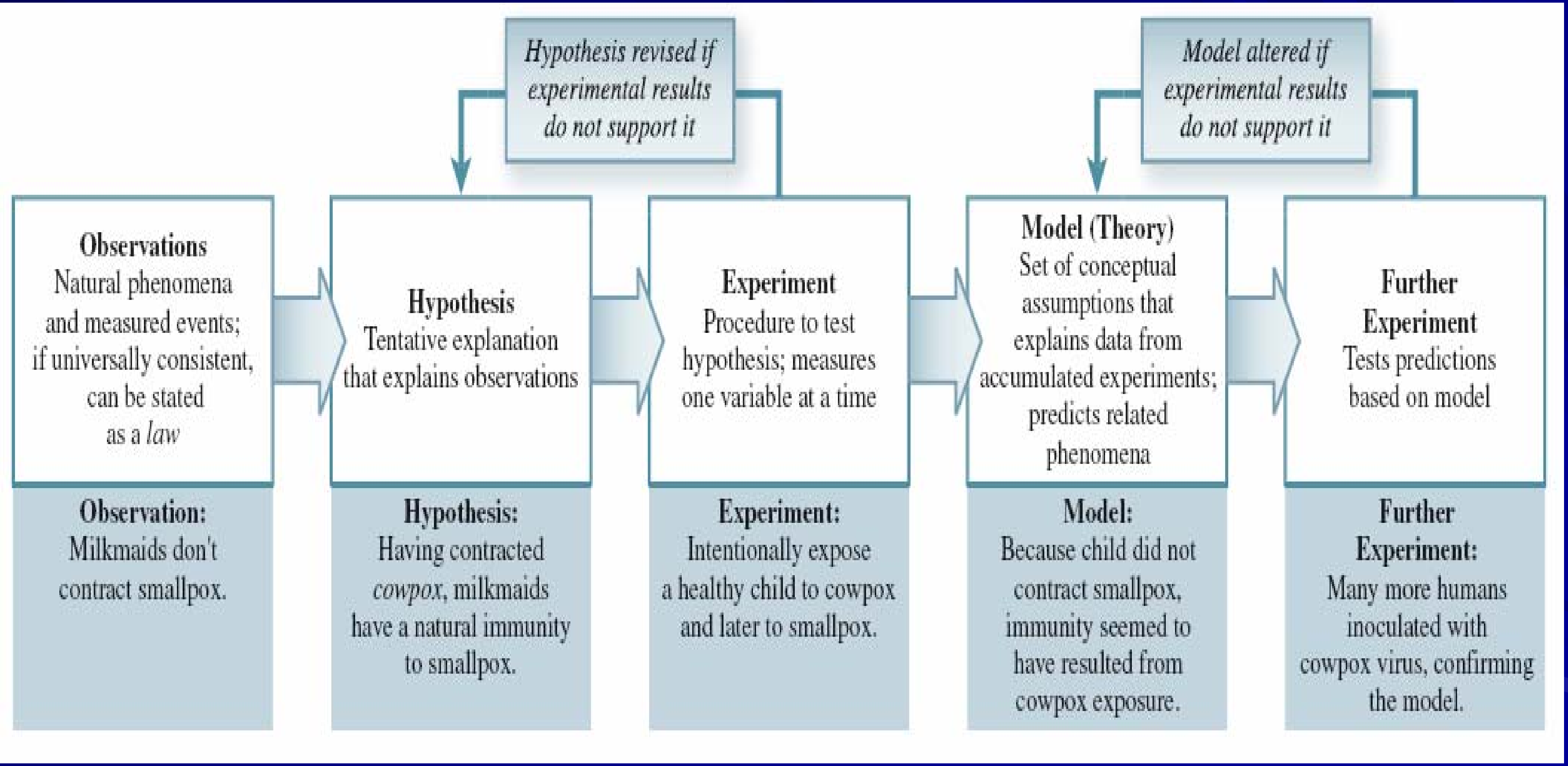
1.2 The scientific method

- It is a way of solving problems
- It consists of the following steps:
  - <u>Observation</u>- what is seen or measured
  - <u>Hypothesis</u>- guess of why things behave the way they do. (possible explanation for an observation)
  - Experiment- designed to test hypothesis
- These steps would lead to new observations, and the cycle goes on
- Once a set of hypotheses agree with observations, they are grouped into a <u>theory</u>

Scientific method

- Thery is a set of tested hypothesis that gives an overall explanation for a natural phenomenon
- Laws are summaries of observations
    - Often mathematical relationship

1.2 The Scientific Method



1.3 Classification of Matter

- **Matter is either classified as a *substance* or a *mixture* of substances.**
- ***Substance***
  - **Can be either an *element* or a *compound***
  - **Has a definite (constant) composition and distinct properties**
  - **Examples: sodium chloride, water, oxygen**

States of Matter

- **Three States of Matter:**
- **Solid:** rigid - fixed volume and shape
- **Liquid:** definite volume but assumes the shape of its container
- **Gas:** no fixed volume or shape - assumes the shape of its container

- States of Matter
  - Solid
    - particles close together in orderly fashion
    - little freedom of motion
    - a solid sample does not conform to the shape of its container
  - Liquid
    - particles close together but not held rigidly in position
    - particles are free to move past one another
    - a liquid sample conforms to the shape of the part of the container it fills

- Gas
  - particles randomly spread apart
  - particles have complete freedom of movement
  - a gas sample assumes both shape and *volume* of container.
- States of matter can be inter-converted without changing chemical composition
  - solid → liquid → gas (add heat)
  - gas → liquid → solid (remove heat)

Substances

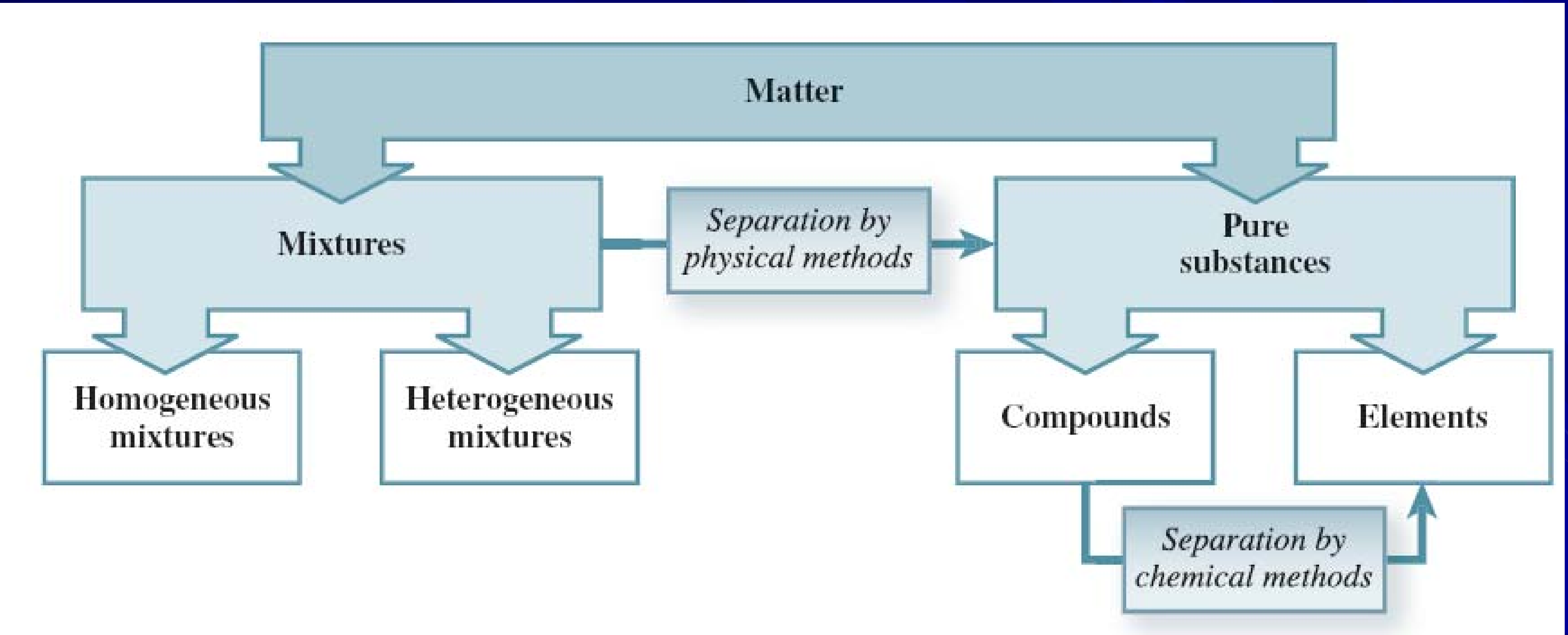
- ***Element***: **cannot be separated into simpler substances by chemical means.**
  - **Examples: iron, mercury, oxygen, and hydrogen**
- ***Compounds***: **two or more elements chemically combined in *definite ratios***
  - **Cannot be separated by physical means**
  - **Examples: salt, water and carbon dioxide**

Mixtures

- ***Mixture:*** **physical combination of two or more substances**
  - **Substances retain distinct identities**
  - **No universal constant composition**
  - **Can be separated by physical means**
    - **Examples: sugar/iron; sugar/water**

- **Types of Mixtures**
  - ***Homogeneous*: composition of the mixture is *uniform* throughout**
    - **Example: sugar dissolved in water**
  - ***Heterogeneous*: composition is *not* uniform throughout**
    - **Example: uniformample: sugar mixed with iron filings**

Classification of Matter



1.3 Scientific Measurement

- Used to measure quantitative properties of matter
- SI base units

SI system (le Systeme International in French) based on the metric system

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

SI Prefixes

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

- ***Mass*: measure of the *amount of matter in an object***
  - **(*weight* refers to gravitational pull)**
  - **Mass cannot change weight can**
- **Temperature:**
  - **Celsius**
    - **Represented by °C**
    - **Based on freezing point of water as 0°C and boiling point of water as 100°C**
  - **Kelvin**
    - **Represented by K (no degree sign)**
    - **The *absolute* scale**
    - **Units of Celsius and Kelvin are equal in magnitude**
  - **Fahrenheit (the English system) (°F)**

Electronic Analytical Balance

Units of Temperature between Boiling and Freezing

	Fahrenheit	Celsius	Kelvin
Water boils	212°F	100°C	373 K
	180°	100°C	100K
Water freezes	32°F	0°C	273 K

Equations for Temperature Conversions

$$^{o}C = (^{o}F - 32) \times \frac{5}{9}$$

$$K = {}^{o}C + 273.15$$

$$^{o}F = \frac{9}{5} \times {}^{o}C + 32$$

Practice

Convert the temperature reading on the local bank (28°C) into the corresponding Fahrenheit temperature.

$$^{o}F = \frac{9}{5} \times {}^{o}C + 32$$

$$^{o}F = \frac{9}{5} \times 28\ {}^{o}C + 32 = 82\ {}^{o}F$$

Units of measurements

- Every measurement has two parts
- Number
- Scale (called a unit)
- SI system (le Systeme International in French) based on the metric system
- Examples:

Prefix

20 grams

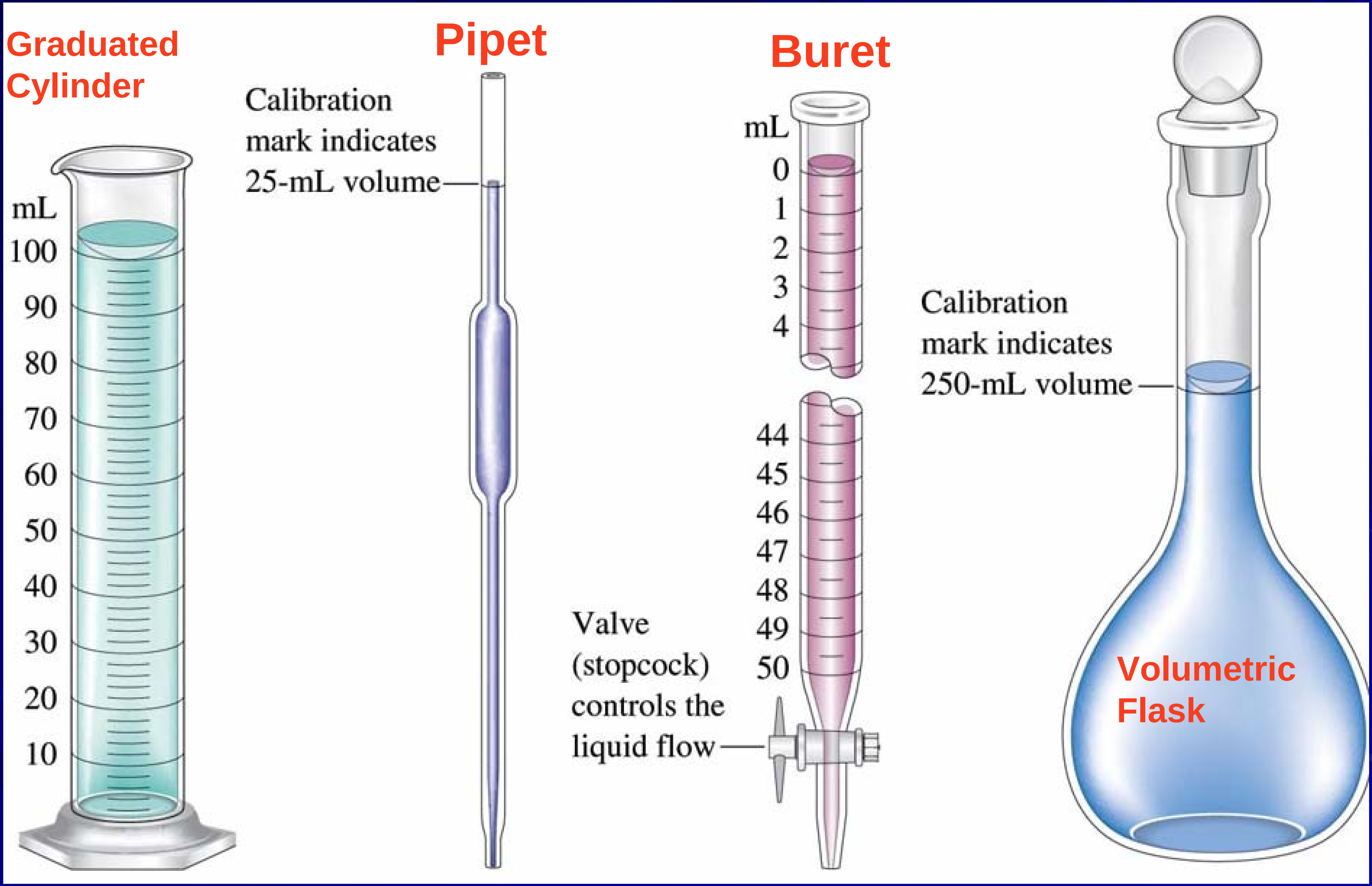
20 k g = 20×10^{3} g

20 m g = 20×10^{-3} g

6.63×10^{-34} Joule seconds

Volume measurement: Liter

- Liter is defined as the volume of 1 dm^3
  - 1 dm^3 =
  - $(10cm)^3$ =
  - 1000 cm^3 =
  - 1000mL



Graduated Cylinder
mL
100
90
80
70
60
50
40
30
20
10
Pipet
Calibration mark indicates 25-mL volume
Buret
mL
0
1
2
3
4
44
45
46
47
48
49
50
Valve (stopcock) controls the liquid flow
Calibration mark indicates 250-mL volume
Volumetric Flask

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

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

Densities of Various Common Substances* at 20° C

TABLE 1.5 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	0.789
Benzene	Liquid	0.880
Water	Liquid	0.9982
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

Density: Ratio of mass to volume•

– Formula:

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

– d = density (g/mL)

– m = mass (g)

– V = volume (mL or cm^3)

(*gas densities are usually expressed in g/L)

Practice

The density of a piece of copper wire is 8.96 g/cm^3. Calculate the volume in cm^3 of a piece of copper with a mass of 4.28 g.

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

$$V = \frac{m}{d} = \frac{4.28 \text{ g}}{8.96 \dfrac{\text{g}}{\text{cm}^3}} = 0.478 \text{ cm}^3$$

Density Problem

- An empty container weighs 121.3 g. When filled with a liquid (density 1.53 g/cm^3) the container weighs 283.2 g. What is the volume of the container?

$$\text{Mass of the liquid} = 161.9\ \text{g}$$

$$V = 161.9 g$$

$$V = 161.9 g \times \frac{1 cm^3}{1.53 g} = 106 cm^3$$

1.4 Properties of Matter

- ***Quantitative***: expressed using ***numbers***
- ***Qualitative***: no precise measurements are needed
- ***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
  – Examples: changes of state (melting, freezing)

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

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

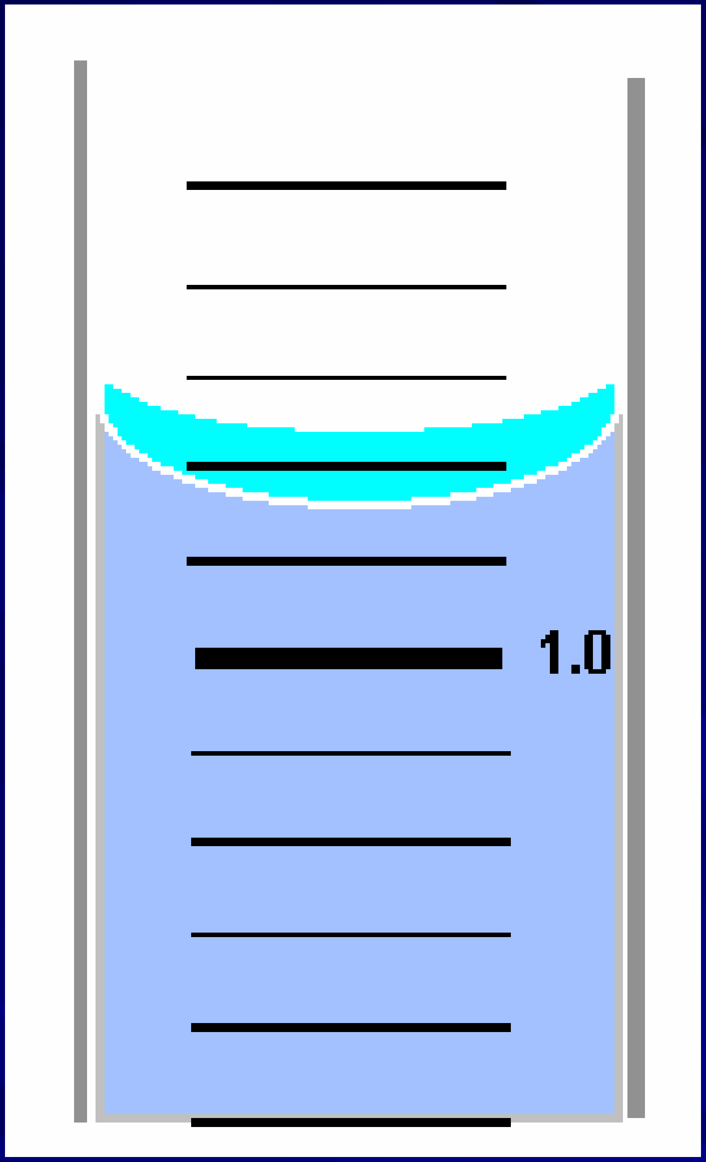
1.5 Uncertainty in Measurement

- **Exact: numbers with defined values**
  - Examples: counting numbers, conversion factors based on definitions
- **Inexact: numbers obtained by any method other than counting**
  - Examples: measured values in the laboratory

Uncertainty in Measurement

- A measurement always has some degree of uncertainty.
- Uncertainty has to be indicated in any measurement.
- Any measurement has **certain digits and one uncertain digit.**
- A digit that must be estimated is called **uncertain.**

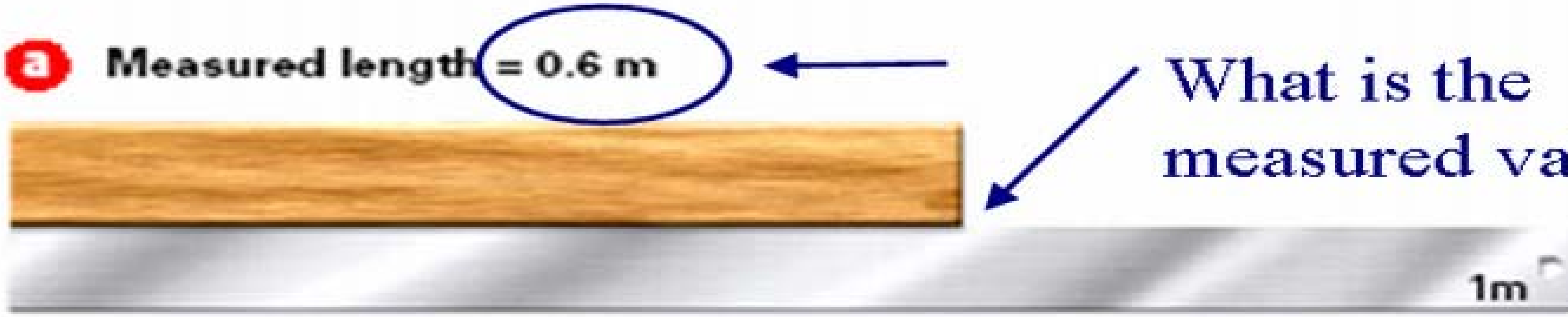
Uncertainty in Measurements



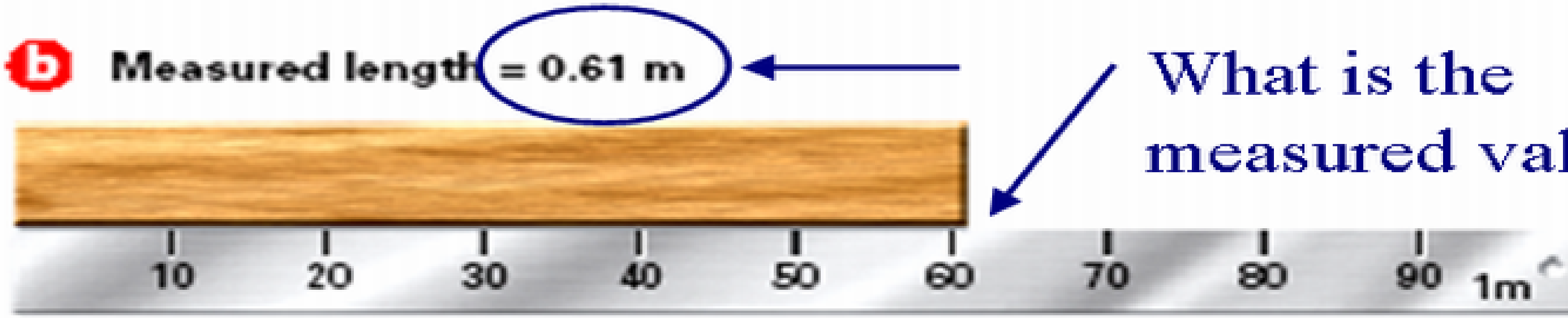
1.14 mL? 1.15 mL? 1.16 mL?

1.15 $\pm$ 0.01 mL

uncertain digit
(1/10 the smallest scale division)



What is the measured value?



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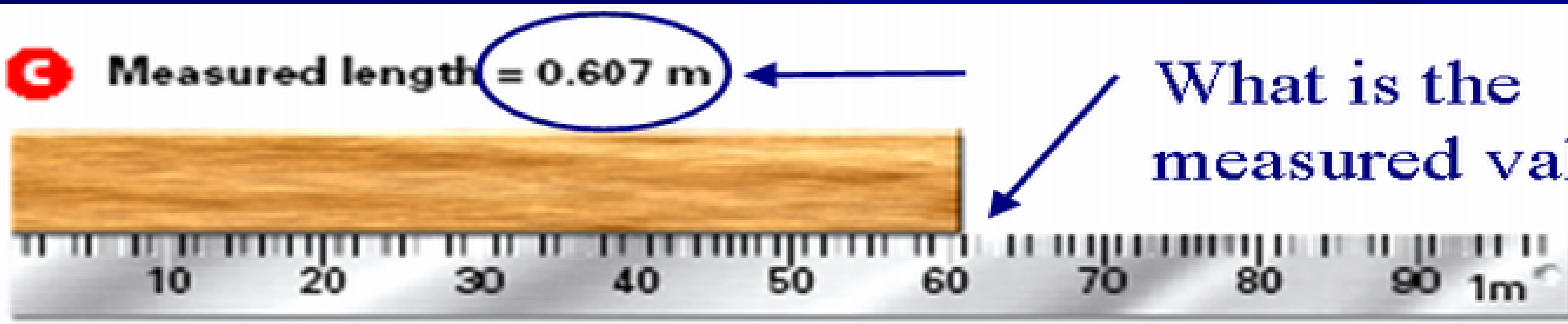
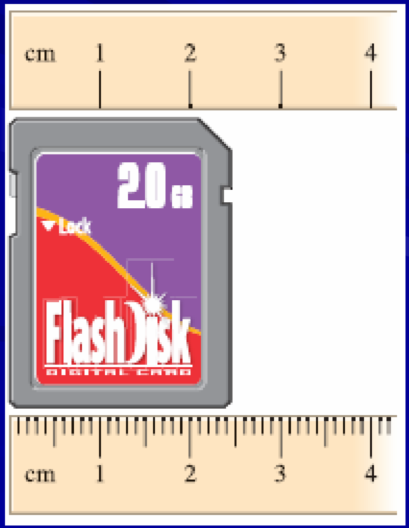


What is the measured value?

Significant Figures

- Used to express the uncertainty of inexact numbers obtained by measurement
- The last digit in a measured value is an uncertain digit - an estimate

- <u>The number of certain digits + the uncertain digit is called number of significant figures.</u>



- Guidelines for significant figures
  - Any non-zero digit is significant
  - Zeros between non-zero digits are significant
  - Zeros to the left of the first non-zero digit are not significant
  - Zeros to the right of the last non-zero digit are significant if decimal is present
  - Zeros to the right of the last non-zero digit are not significant if decimal is not present

Practice

Determine the number of significant figures in each of the following.

345.5 cm

4 significant figures

0.0058 g

2 significant figures

1205 m

4 significant figures

250 mL

2 significant figures

250.00 mL

5 significant figures

- Calculations with measured numbers
  - Addition and subtraction
    - Answer cannot have more digits to the right of the decimal than any of original numbers
    - Example:

102.50 two digits after decimal point

+ 0.231 three digits after decimal point

102.731 round to 102.73

Multiple computations

$$\frac{2.54 \times 0.0028}{0.0105 \times 0.060} =$$

1) 11.3 2) 11 3) 0.041

Continuous calculator operation =

2.54 x 0.0028 ÷ 0.0105 ÷ 0.060 = 11

$$\frac{6.404 \times 2.91}{18.7 - 17.1}$$

- **Here, the mathematical operation requires that we apply the addition/ subtraction rule first, then apply the multiplication/division rule.**

$$\frac{6.404 \times 2.91}{18.7 - 17.1} = \frac{6.404 \times 2.91}{1.\underline{6}}$$

= 12

- Exact numbers
  - Do not limit answer because exact numbers have an infinite number of significant figures
  - Example:

    A coin of 25 halals has a mass of 2.5 g. If we have three such coines, the total mass is

$$3 \times 2.5\ g = 7.5\ g$$

  - In this case, 3 is an exact number and does not limit the number of significant figures in the result.

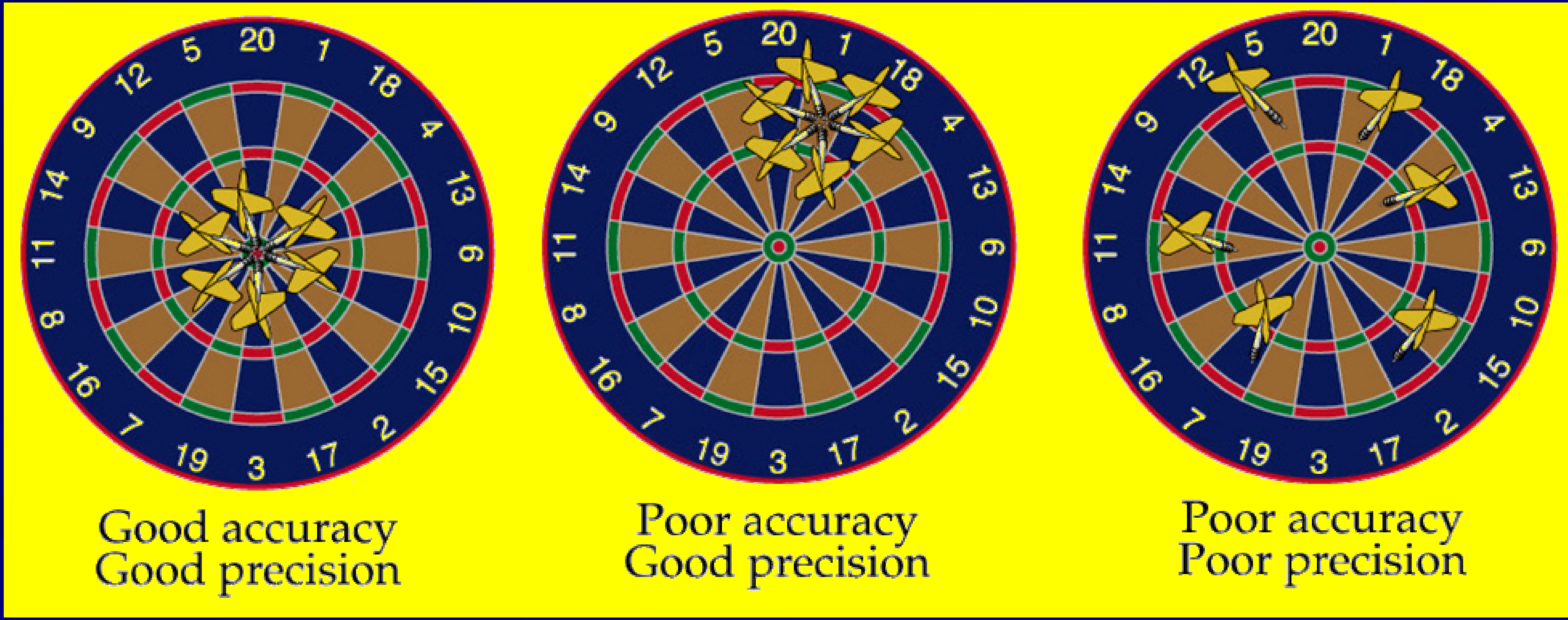
- **Rounding rules**
  - If the number is less than 5 round “down”.
  - If the number is 5 or greater round “up”.

Practice

- **105.5 L + 10.65 L =**
  - **Calculator answer: 116.15 L**
  - **Round to: 116.2 L ⟵ Answer to the tenth position**
- **1.0267 cm x 2.508 cm x 12.599 cm**
  - **Calculator answer: 32.4419664 cm^3**
  - **Round to: 32.44 cm^3 ⟵ round to the smallest number of significant figures**

- **Accuracy and precision**
  - Two ways to gauge the quality of a set of measured numbers
  - ***Accuracy***: how close a measurement is to the true or accepted value
  - ***Precision***: how closely measurements of the same thing are to one another

Precision and Accuracy

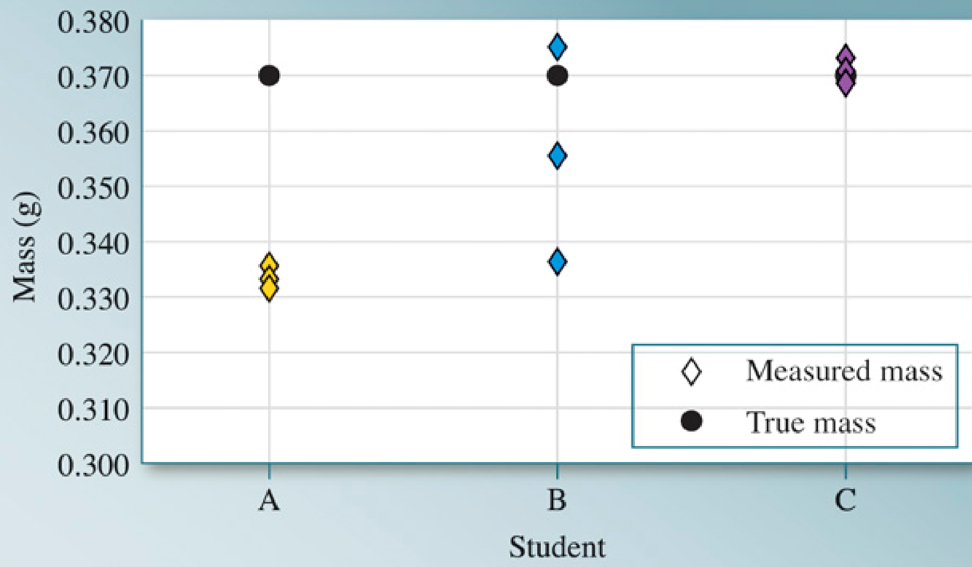


Describe accuracy and precision for each set

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 is 0.370 grams

	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



Student A's results are precise but not accurate.

Student B's results are neither precise nor accurate.

Student C's results are both precise and accurate.

1.6 Using Units and Solving Problems

- ***Conversion factor*: a fraction in which the same quantity is expressed one way in the numerator and another way in the denominator**
  - **Example: by definition, 1 inch = 2.54 cm**

$$\frac{1\,\text{in}}{2.54\,\text{cm}} \qquad \frac{2.54\,\text{cm}}{1\,\text{in}}$$

- ***Dimensional analysis*: a problem solving method employing conversion factors to change one measure to another often called the "factor-label method"**
  - **Example: Convert 12.00 inches to meters**
    - **Conversion factors needed:**

2.54 cm = 1 in and 100 cm = 1 meter

$$12.00\,\cancel{\text{in}} \times \frac{2.54\,\cancel{\text{cm}}}{1\,\cancel{\text{in}}} \times \frac{1\,\text{m}}{100\,\cancel{\text{cm}}} = 0.3048\,\text{m}$$

***Note that neither conversion factor limited the number of significant figures in the result because they both consist of exact numbers.**

Practice

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?

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

How many seconds are in 1.4 days?

Unit plan: days → hr → min → seconds

$$1.4\ \cancel{\text{days}} \times \frac{24\ \text{hr}}{1\ \cancel{\text{day}}} \times \text{??}$$

Exact numbers

$$1.4\ \cancel{\text{day}} \times \frac{24\ \cancel{\text{hr}}}{1\ \cancel{\text{day}}} \times \frac{60\ \cancel{\text{min}}}{1\ \cancel{\text{hr}}} \times \frac{60\ \text{sec}}{1\ \cancel{\text{min}}}$$

$$= 1.2 \times 10^5\ \text{sec}$$

How many minutes are in 2.5 hours?

Initial unit

2.5 hr

Conversion factor ***Final unit***

$$2.5\ \cancel{hr} \times \frac{60\ min}{1\ \cancel{hr}} = 150\ min$$

Multiple units

- The speed limit is 65 mi/hr. What is this in m/s?
  - 1 mile = 1760 yds
  - 1 meter = 1.094 yds

$$\frac{65\ \text{mi}}{\text{hr}} \cdot \frac{1760\ \text{yd}}{1\ \text{mi}} \cdot \frac{1\ \text{m}}{1.094\ \text{yd}} \cdot \frac{1\ \text{hr}}{60\ \text{min}} \cdot \frac{1\ \text{min}}{60\ \text{s}}$$

If you are running at a speed of 65 meters per minute, how many seconds will it take for you to walk a distance of 8450 feet?

Initial

$$8450\ \cancel{ft} \times \frac{12\ \cancel{in.}}{1\ \cancel{ft}} \times \frac{2.54\ \cancel{cm}}{1\ \cancel{in.}} \times \frac{1\ \cancel{m}}{100\ \cancel{cm}}$$

$$\times \frac{1\ \cancel{min}}{65\ \cancel{m}} \times \frac{\cancel{60}\ sec}{1\ min} = 2400\ sec$$

Key Points

- Scientific method
- Classifying matter
- SI conversions
- Density
- Temperature conversions
- Physical vs chemical properties and changes
- Precision vs accuracy
- Dimensional analysis