Objectives

- **Describe** the purpose of the scientific method.
- **Distinguish** between qualitative and quantitative observations.
- **Describe** the differences between hypotheses, theories, and models.

Scientific Method

- The **scientific method** is a logical approach to solving problems by observing and collecting data, formulating hypotheses, testing hypotheses, and formulating theories that are supported by data.

Observing and Collecting Data

- *Observing* is the use of the senses to obtain information.
- data may be
  - *qualitative* (descriptive)
  - *quantitative* (numerical)
- A **system** is a specific portion of matter in a given region of space that has been selected for study during an experiment or observation.

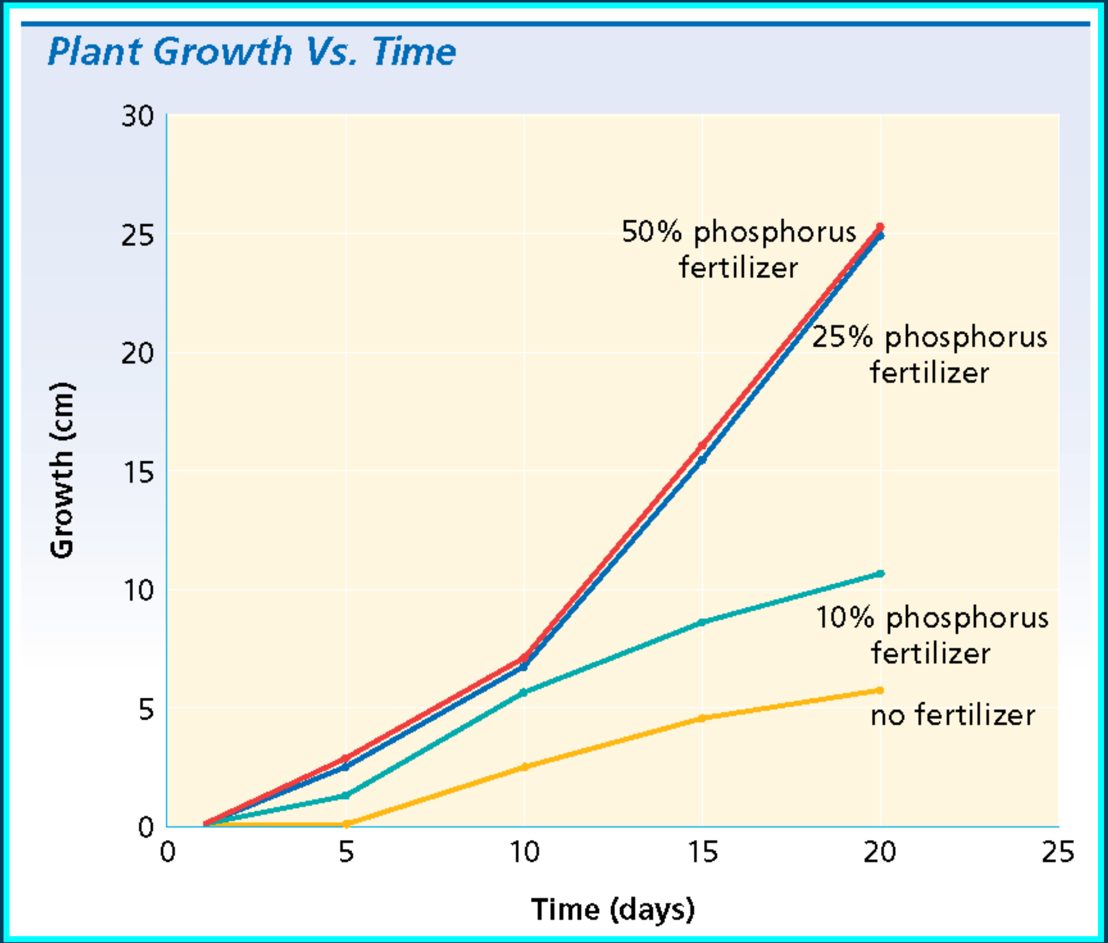
End Of Slide

Formulating Hypotheses

- Scientists make generalizations based on the data.
- Scientists use generalizations about the data to formulate a **hypothesis**, or testable statement.
  - Hypotheses are often "if-then" statements.

End Of Slide

Formulating Hypotheses



Testing Hypotheses

- Testing a hypothesis requires experimentation that provides data to support or refute a hypothesis or theory.
- *Controls* are the experimental conditions that remain constant.
- *Variables* are any experimental conditions that change.

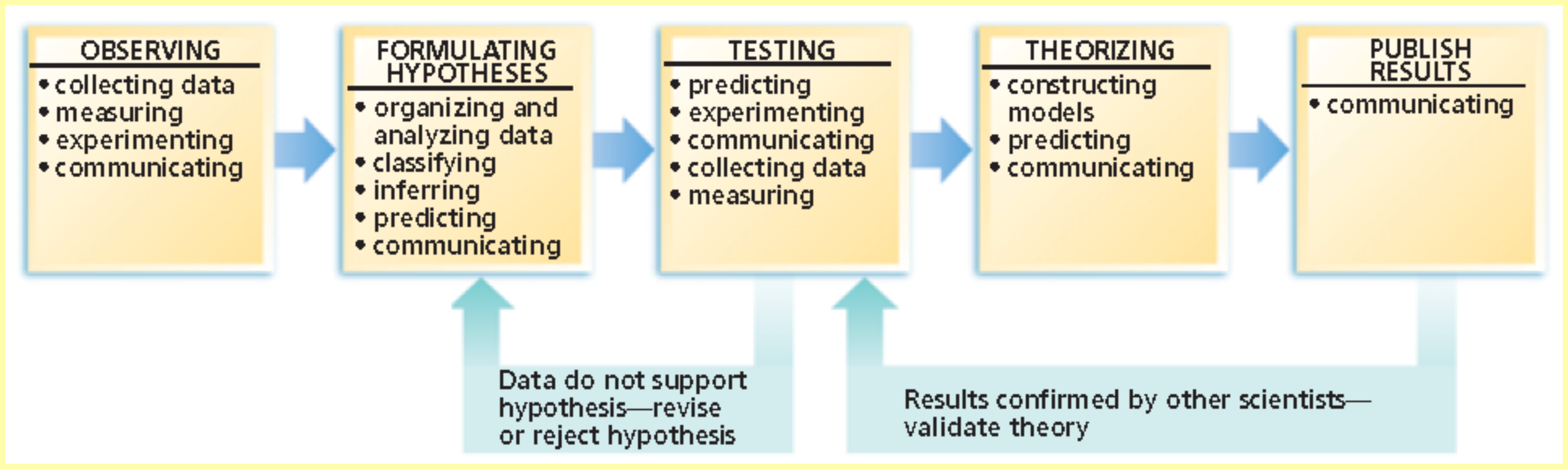
Theorizing

- A **model** in science is more than a physical object; it is often an explanation of how phenomena occur and how data or events are related.
  - visual, verbal, or mathematical
  - example: atomic model of matter

- A **theory** is a broad generalization that explains a body of facts or phenomena.
  - example: atomic theory

End Of Slide

Scientific Method



Lesson Starter

- Would you be breaking the speed limit in a 40 mi/h zone if you were traveling at 60 km/h?
  - one kilometer = 0.62 miles
  - 60 km/h = 37.2 mi/h
  - You would not be speeding!
- km/h and mi/h measure the same quantity using different units

Objectives

- **Distinguish** between a quantity, a unit, and a measurement standard.
- **Name** and **use** SI units for length, mass, time, volume, and density.
- **Distinguish** between mass and weight.
- **Perform** density calculations.
- **Transform** a statement of equality into a conversion factor.

Units of Measurement

- Measurements represent quantities.
- A **quantity** is something that has magnitude, size, or amount.
- measurement ≠ quantity
  - the teaspoon is a unit of measurement
  - volume is a quantity
- The choice of unit depends on the quantity being measured.

End Of Slide

SI Measurement

- Scientists all over the world have agreed on a single measurement system called Le Système International d' Unités, abbreviated **SI.**

- SI has seven base units
  - most other units are derived from these seven

End Of Slide

SI Base Units

Quantity	Quantity symbol	Unit name	Unit abbreviation	Defined standard
Length	l	meter	m	the length of the path traveled by light in a vacuum during a time interval of 1/299 792 458 of a second
Mass	m	kilogram	kg	the unit of mass equal to the mass of the international prototype of the kilogram
Time	t	second	s	the duration of 9 192 631 770 periods of the radiation corresponding to the transition between the two hyperfine levels of the ground state of the cesium-133 atom
Temperature	T	kelvin	K	the fraction 1/273.16 of the thermodynamic temperature of the triple point of water
Amount of substance	n	mole	mol	the amount of substance of a system which contains as many elementary entities as there are atoms in 0.012 kilogram of carbon-12
Electric current	I	ampere	A	the constant current which, if maintained in two straight parallel conductors of infinite length, of negligible circular cross section, and placed 1 meter apart in vacuum, would produce between these conductors a force equal to 2×10^{-7} newton per meter of length
Luminous intensity	I_v	candela	cd	the luminous intensity, in a given direction, of a source that emits monochromatic radiation of frequency 540×10^{12} hertz and that has a radiant intensity in that direction of 1/683 watt per steradian

SI Base Units

Mass

- Mass is a measure of the quantity of matter.
- The SI standard unit for mass is the kilogram.
- **Weight** is a measure of the gravitational pull on matter.
- Mass does not depend on gravity.

SI Base Units

Length

- Length is a measure of distance.
- The SI standard for length is the meter.
- The kilometer, km, is used to express longer distances
- The centimeter, cm, is used to express shorter distances

Derived SI Units

- Combinations of SI base units form **derived units.**
  - pressure is measured in $kg/m \cdot s^2$, or *pascals*

Quantity	Quantity symbol	Unit	Unit abbreviation	Derivation
Area	A	square meter	m^2	length × width
Volume	V	cubic meter	m^3	length × width × height
Density	D	kilograms per cubic meter	$\frac{kg}{m^3}$	$\frac{\text{mass}}{\text{volume}}$
Molar mass	M	kilograms per mole	$\frac{kg}{mol}$	$\frac{\text{mass}}{\text{amount of substance}}$
Molar volume	V_m	cubic meters per mole	$\frac{m^3}{mol}$	$\frac{\text{volume}}{\text{amount of substance}}$
Energy	E	joule	J	force × length

Derived SI Units, *continued*

Volume

- **Volume** is the amount of space occupied by an object.
  - The derived SI unit is cubic meters, m^3
  - The cubic centimeter, cm^3, is often used
  - The liter, L, is a non-SI unit
  - $1\ L = 1000\ cm^3$
  - $1\ mL = 1\ cm^3$

End Of Slide

Derived SI Units, *continued*

Density

- **Density** is the ratio of mass to volume, or mass divided by volume.

$$\textit{density} = \frac{\textit{mass}}{\textit{volume}} \quad \text{or} \quad D = \frac{m}{V}$$

  - The derived SI unit is kilograms per cubic meter, kg/m^3
  - g/cm^3 or g/mL are also used
  - Density is a characteristic physical property of a substance.

End Of Slide

Derived SI Units, *continued*

Density

- Density can be used as one property to help identify a substance

Solids	Density at 20°C (g/cm^3)	Liquids	Density at 20°C (g/mL)
cork	0.24*	gasoline	0.67*
butter	0.86	ethyl alcohol	0.791
ice	0.92†	kerosene	0.82
sucrose	1.59	turpentine	0.87
bone	1.85*	water	0.998
diamond	3.26*	sea water	1.025**
copper	8.92	milk	1.031*
lead	11.35	mercury	13.6

† measured at 0°C
* typical density
** measured at 15°C

Derived SI Units, *continued*

Sample Problem A

A sample of aluminum metal has a mass of 8.4 g. The volume of the sample is 3.1 cm^3. Calculate the density of aluminum.

Derived SI Units, *continued*

Sample Problem A Solution

Given: mass (m) = 8.4 g
volume (V) = 3.1 cm^3

Unknown: density (D)

Solution:

$$\textit{density} = \frac{\textit{mass}}{\textit{volume}} = \frac{8.4\ \text{g}}{3.1\ \text{cm}^3} = \boxed{2.7\ \text{g}/\text{cm}^3}$$

End Of Slide

Conversion Factors

- A **conversion factor** is a ratio derived from the equality between two different units that can be used to convert from one unit to the other.
  - example: How quarters and dollars are related

$$\frac{4\ \text{quarters}}{1\ \text{dollar}} = 1 \qquad \frac{1\ \text{dollar}}{4\ \text{quarters}} = 1$$

$$\frac{0.25\ \text{dollar}}{1\ \text{quarters}} = 1 \qquad \frac{1\ \text{quarter}}{0.25\ \text{dollar}} = 1$$

Conversion Factors, *continued*

- **Dimensional analysis** is a mathematical technique that allows you to use units to solve problems involving measurements.
- quantity sought = quantity given × conversion factor
- example: the number of quarters in 12 dollars

number of quarters = 12 dollars × conversion factor

$$? \text{ quarters} = 12 \text{ dollars} \times \frac{4 \text{ quarter}}{1 \text{ dollar}} = 48 \text{ quarters}$$

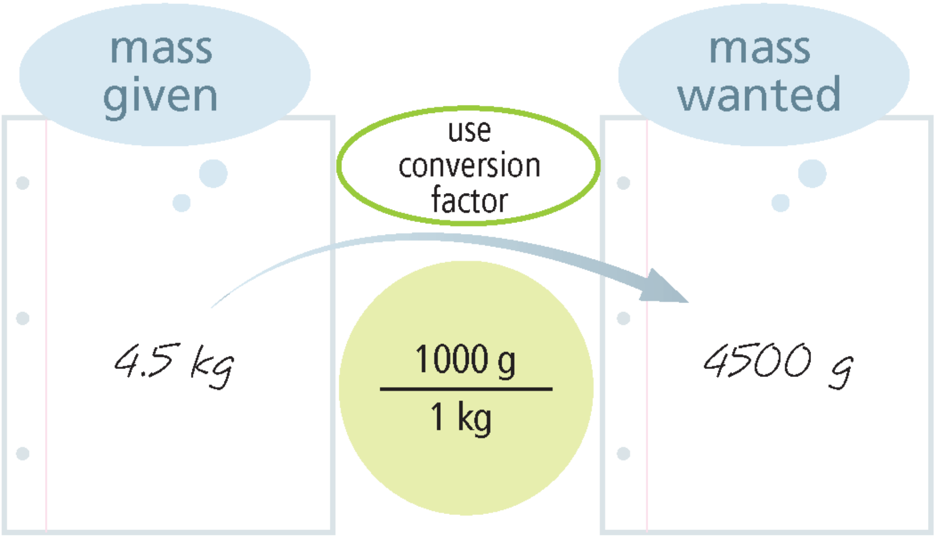
End Of Slide

Using Conversion Factors

1. Identify the quantity and unit given and the unit that you want to convert to.

2. Using the equality that relates the two units, set up the conversion factor that cancels the given unit and leaves the unit that you want to convert to.

3. Multiply the given quantity by the conversion factor. Cancel units to verify that the units left are the ones you want for your answer.



Conversion Factors, *continued*

Deriving Conversion Factors

- You can derive conversion factors if you know the relationship between the unit you have and the unit you want.
  - example: conversion factors for meters and decimeters

$$\frac{1\ \text{m}}{10\ \text{dm}} \qquad \frac{0.1\ \text{m}}{\text{dm}} \qquad \frac{10\ \text{dm}}{\text{m}}$$

SI Conversions

Key problem-solving approach:
SI conversions

Prefix	Arrow toward base unit	Arrow away from base unit
giga (g)	$\times 10^9$	$\div 10^9$
mega (M)	$\times 10^6$	$\div 10^6$
kilo (k)	$\times 10^3$	$\div 10^3$
centi (c)	$\div 10^2$	$\times 10^2$
milli (m)	$\div 10^3$	$\times 10^3$
micro (μ)	$\div 10^6$	$\times 10^6$
nano (n)	$\div 10^9$	$\times 10^9$
pico (p)	$\div 10^{12}$	$\times 10^{12}$

Base Unit*

meter, m
seconds, s
ampere, A
mole, mol
candela, cd

*Kilogram, the base unit for mass, does not appear in this list because it has a different set of conversion values (1 kg = 1000 g).

Conversion Factors, *continued*

Sample Problem B

Express a mass of 5.712 grams in milligrams and in kilograms.

End Of Slide

Conversion Factors, *continued*

Sample Problem B Solution

Express a mass of 5.712 grams in milligrams and in kilograms.

Given: 5.712 g

Unknown: mass in mg and kg

Solution: mg

$$1\ \text{g} = 1000\ \text{mg}$$

Possible conversion factors: $\frac{1000\ \text{mg}}{\text{g}}$ and $\frac{1\ \text{g}}{1000\ \text{mg}}$

$$5.712\ \text{g} \times \frac{1000\ \text{mg}}{\text{g}} = 5712\ \text{mg}$$

End Of Slide

Conversion Factors, *continued*

Sample Problem B Solution, *continued*

Express a mass of 5.712 grams in milligrams and in kilograms.

Given: 5.712 g

Unknown: mass in mg and kg

Solution: kg

$$1\,000 \text{ g} = 1 \text{ kg}$$

Possible conversion factors: $\frac{1000 \text{ g}}{\text{kg}}$ and $\frac{1 \text{ kg}}{1000 \text{ g}}$

$$5.712 \cancel{\text{g}} \times \frac{1 \text{ kg}}{1000 \cancel{\text{g}}} = 0.005712 \text{ kg}$$

Lesson Starter

- Look at the specifications for electronic balances. How do the instruments vary in precision?
- Discuss using a beaker to measure volume versus using a graduated cylinder. Which is more precise?

Objectives

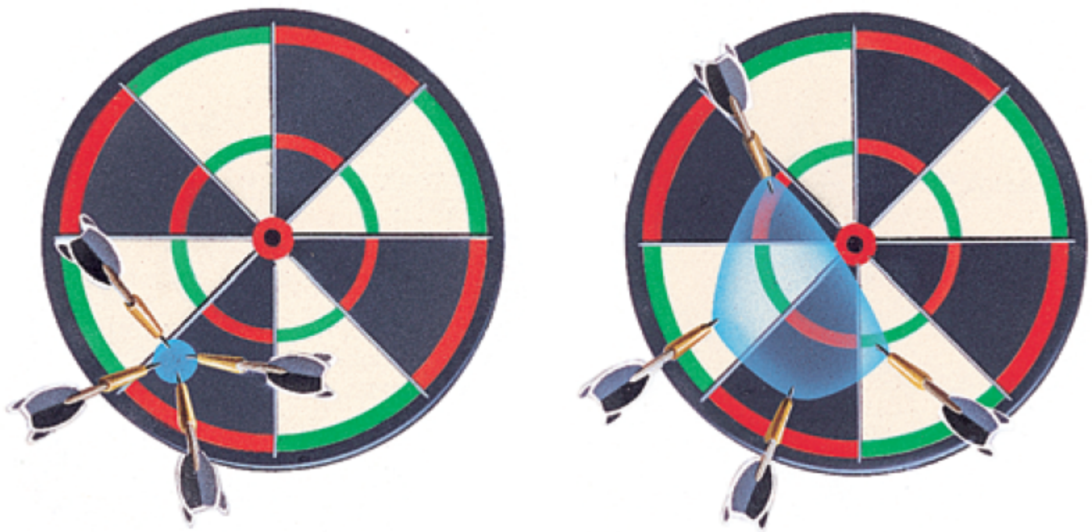
- **Distinguish** between accuracy and precision.
- **Determine** the number of significant figures in measurements.
- **Perform** mathematical operations involving significant figures.
- **Convert** measurements into scientific notation.
- **Distinguish** between inversely and directly proportional relationships.

Accuracy and Precision

- **Accuracy** refers to the closeness of measurements to the correct or accepted value of the quantity measured.

- **Precision** refers to the closeness of a set of measurements of the same quantity made in the same way.

Accuracy and Precision



a Darts within the bull's-eye mean high accuracy and high precision.

b Darts clustered within a small area but far from the bull's-eye mean low accuracy and high precision.

c Darts scattered around the target and far from the bull's-eye mean low accuracy and low precision.

Accuracy and Precision, *continued*

Percentage Error

- **Percentage error** is calculated by subtracting the accepted value from the experimental value, dividing the difference by the accepted value, and then multiplying by 100.

$$\text{Percentage error} = \frac{Value_{experimental} - Value_{accepted}}{Value_{accepted}} ? 100$$

End Of Slide

Accuracy and Precision, *continued*

Sample Problem C

A student measures the mass and volume of a substance and calculates its density as 1.40 g/mL. The correct, or accepted, value of the density is 1.30 g/mL. What is the percentage error of the student's measurement?

Accuracy and Precision, *continued*

Sample Problem C Solution

$$\text{Percentage error} = \frac{Value_{experimental} - Value_{accepted}}{Value_{accepted}} \times 100$$

$$= \frac{1.40\ g/mL - 1.30\ g/mL}{1.30\ g/mL} \times 100 = 7.7\%$$

Accuracy and Precision, *continued*

Error in Measurement

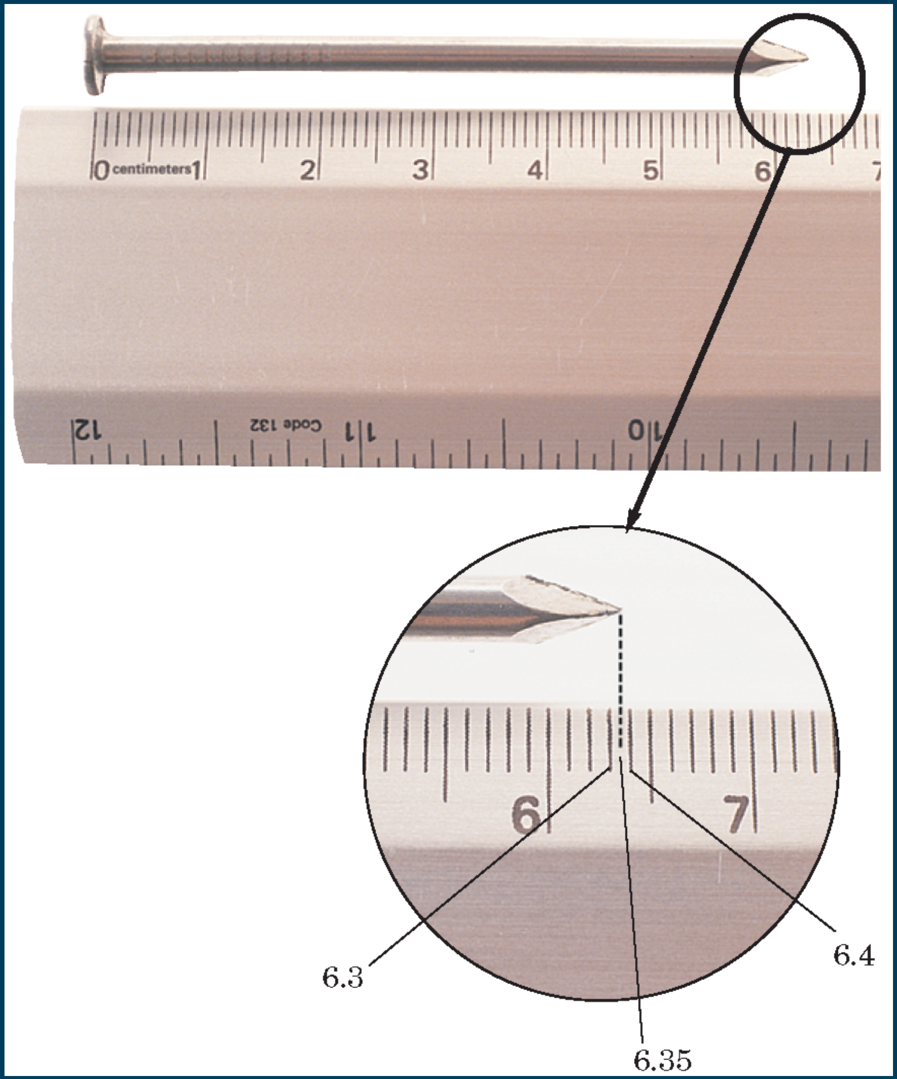
- Some error or uncertainty always exists in any measurement.
    - skill of the measurer
    - conditions of measurement
    - measuring instruments

End Of Slide

Significant Figures

- **Significant figures** in a measurement consist of all the digits known with certainty plus one final digit, which is somewhat uncertain or is estimated.
- The term significant **does not** mean certain.

Reporting Measurements Using Significant Figures



Significant Figures, *continued*

Determining the Number of Significant Figures

Rule	Examples
1. Zeros appearing between nonzero digits are significant.	a. 40.7 L has three significant figures. b. 87 009 km has five significant figures.
2. Zeros appearing in front of all nonzero digits are not significant.	a. 0.095 897 m has five significant figures. b. 0.0009 kg has one significant figure.
3. Zeros at the end of a number and to the right of a decimal point are significant.	a. 85.00 g has four significant figures. b. 9.000 000 000 mm has 10 significant figures.
4. Zeros at the end of a number but to the left of a decimal point may or may not be significant. If a zero has not been measured or estimated but is just a placeholder, it is not significant. A decimal point placed after zeros indicates that they are significant.	a. 2000 m may contain from one to four significant figures, depending on how many zeros are placeholders. **For measurements given in this text, assume that 2000 m has one significant figure.** b. 2000. m contains four significant figures, indicated by the presence of the decimal point.

Significant Figures

Click below to watch the Visual Concept.

Visual Concept

Rules for Determining Significant Zeros

Click below to watch the Visual Concept.

Visual Concept

Significant Figures, *continued*

Sample Problem D

How many significant figures are in each of the following measurements?

a. 28.6 g

b. 3440. cm

c. 910 m

d. 0.046 04 L

e. 0.006 700 0 kg

End Of Slide

Significant Figures, *continued*

Sample Problem D Solution

a. 28.6 g

There are no zeros, so all three digits are significant.

b. 3440. cm

By rule 4, the zero is significant because it is immediately followed by a decimal point; there are 4 significant figures.

c. 910 m

By rule 4, the zero is not significant; there are 2 significant figures.

Significant Figures, *continued*

Sample Problem D Solution, *continued*

d. 0.046 04 L

By rule 2, the first two zeros are not significant; by rule 1, the third zero is significant; there are 4 significant figures.

e. 0.006 700 0 kg

By rule 2, the first three zeros are not significant; by rule 3, the last three zeros are significant; there are 5 significant figures.

Significant Figures, *continued*

Rounding

If the digit following the last digit to be retained is:	then the last digit should:	Example (rounded to three significant figures)
greater than 5	be increased by 1	42.68 g ⟶ 42.7 g
less than 5	stay the same	17.32 m ⟶ 17.3 m
5, followed by nonzero digit(s)	be increased by 1	2.7851 cm ⟶ 2.79 cm
5, not followed by nonzero digit(s), and preceded by an odd digit	be increased by 1	4.635 kg ⟶ 4.64 kg (because 3 is odd)
5, not followed by nonzero digit(s), and the preceding significant digit is even	stay the same	78.65 mL ⟶ 78.6 mL (because 6 is even)

Significant Figures, *continued*

Addition or Subtraction with Significant Figures

- When adding or subtracting decimals, the answer must have the same number of digits to the right of the decimal point as there are in the measurement having the fewest digits to the right of the decimal point.

Addition or Subtraction with Significant Figures

- For multiplication or division, the answer can have no more significant figures than are in the measurement with the fewest number of significant figures.

Significant Figures, *continued*

Sample Problem E

Carry out the following calculations. Express each answer to the correct number of significant figures.

a. 5.44 m - 2.6103 m

b. 2.4 g/mL × 15.82 mL

Significant Figures, *continued*

Sample Problem E Solution

a. 5.44 m - 2.6103 m = 2.84 m

There should be two digits to the right of the decimal point, to match 5.44 m.

b. 2.4 g/mL × 15.82 mL = 38 g

There should be two significant figures in the answer, to match 2.4 g/mL.

End Of Slide

Significant Figures, *continued*

Conversion Factors and Significant Figures

- There is no uncertainty exact conversion factors.
- Most exact conversion factors are defined quantities.

Scientific Notation

- In **scientific notation,** numbers are written in the form $M \times 10^n$, where the factor M is a number greater than or equal to 1 but less than 10 and n is a whole number.

  - example: 0.000 12 mm = 1.2×10^{-4} mm

  - Move the decimal point four places to the right and multiply the number by 10^{-4}.

End of Slide

Scientific Notation, *continued*

1. Determine M by moving the decimal point in the original number to the left or the right so that only one nonzero digit remains to the left of the decimal point.

2. Determine n by counting the number of places that you moved the decimal point. If you moved it to the left, n is positive. If you moved it to the right, n is negative.

Scientific Notation, *continued*

Mathematical Operations Using Scientific Notation

1. Addition and subtraction —These operations can be performed only if the values have the same exponent (n factor).

example: 4.2×10^4 kg + 7.9×10^3 kg

$$\begin{array}{r} 4.2 \times 10^4 \text{ kg} \\ +0.79 \times 10^4 \text{ kg} \\ \hline 4.99 \times 10^4 \text{ kg} \end{array}$$

rounded to 5.0×10^4 kg

or

$$\begin{array}{r} 7.9 \times 10^3 \text{ kg} \\ +42 \times 10^3 \text{ kg} \\ \hline 49.9 \times 10^3 \text{ kg} \end{array} = 4.99 \times 10^4 \text{ kg}$$

rounded to 5.0×10^4 kg

End of Slide

Scientific Notation, *continued*

Mathematical Operations Using Scientific Notation

2. Multiplication —The M factors are multiplied, and the exponents are added algebraically.

example: $(5.23 \times 10^{6}\ \mu m)(7.1 \times 10^{-2}\ \mu m)$

$= (5.23 \times 7.1)(10^{6} \times 10^{-2})$

$= 37.133 \times 10^{4}\ \mu m^{2}$

$= 3.7 \times 10^{5}\ \mu m^{2}$

End Of Slide

Scientific Notation, *continued*

Mathematical Operations Using Scientific Notation

3. Division — The M factors are divided, and the exponent of the denominator is subtracted from that of the numerator.

example: $\dfrac{5.44 \times 10^{7}\ \text{g}}{8.1 \times 10^{4}\ \text{mol}}$

$$= \frac{5.44}{8.1} \times 10^{7-4}\ \text{g/mol}$$

$$= 0.6716049383 \times 10^{3}$$

$$= 6.7 \times 10^{2}\ \text{g/mol}$$

End Of Slide

Using Sample Problems

- **Analyze**
  The first step in solving a quantitative word problem is to read the problem carefully at least twice and to analyze the information in it.

- **Plan**
  The second step is to develop a plan for solving the problem.

- **Compute**
  The third step involves substituting the data and necessary conversion factors into the plan you have developed.

Using Sample Problems, *continued*

- **Evaluate**

  Examine your answer to determine whether it is reasonable.

  1. Check to see that the units are correct.
  2. Make an estimate of the expected answer.
  3. Check the order of magnitude in your answer.
  4. Be sure that the answer given for any problem is expressed using the correct number of significant figures.

End Of Slide

Using Sample Problems, *continued*

Sample Problem F

Calculate the volume of a sample of aluminum that has a mass of 3.057 kg. The density of aluminum is 2.70 g/cm3.

Using Sample Problems, *continued*

Sample Problem F Solution

1. Analyze

Given: mass = 3.057 kg, density = 2.70 g/cm3

Unknown: volume of aluminum

2. Plan

The density unit is g/cm3, and the mass unit is kg.

conversion factor: 1000 g = 1 kg

Rearrange the density equation to solve for volume.

$$D = \frac{m}{V} \quad \Rightarrow \quad V = \frac{m}{D}$$

Using Sample Problems, *continued*

Sample Problem F Solution, *continued*

3. Compute

$$V = \frac{3.057\ \text{kg}}{2.70\ \text{g/cm}^3} \times \frac{1000\ \text{g}}{\text{kg}}$$

$= 1132.222\ldots\ \text{cm}^3$ (calculator answer)

round answer to three significant figures

$$V = 1.13 \times 10^3\ \text{cm}^3$$

End of Slide

Using Sample Problems, *continued*

Sample Problem F Solution, *continued*

4. Evaluate

Answer: $V = 1.13 \times 10^3\ \text{cm}^3$

- The unit of volume, cm^3, is correct.
- An order-of-magnitude estimate would put the answer at over 1000 cm^3.

$$\frac{3}{2} \times 1000$$

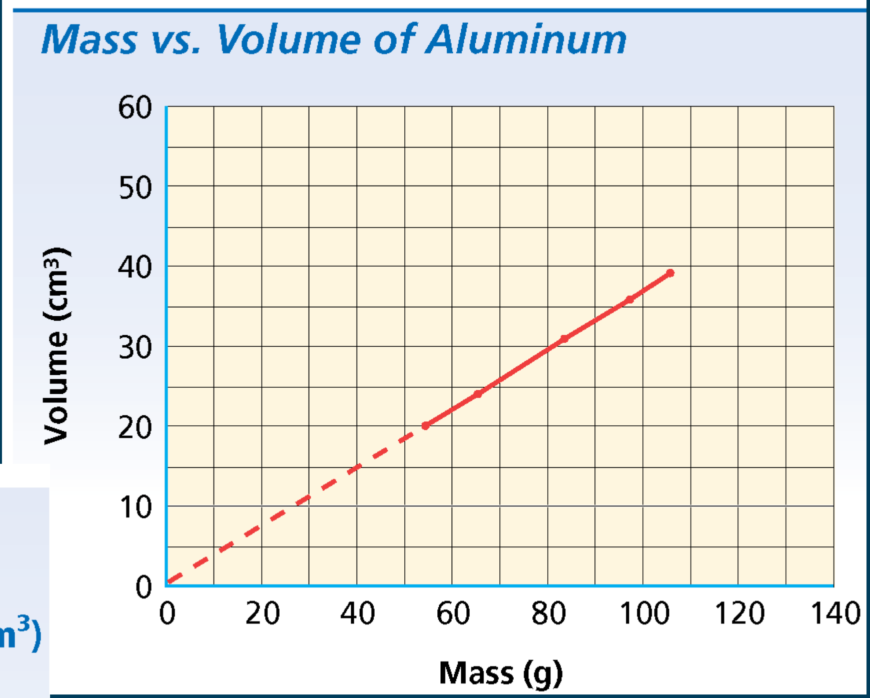
- The correct number of significant figures is three, which matches that in 2.70 g/cm.

End Of Slide

Direct Proportions

- Two quantities are **directly proportional** to each other if dividing one by the other gives a constant value.
- $y \propto x$
  - read as "y is proportional to x."

Direct Proportion



Mass-Volume Data for Aluminum at 20°C

Mass (g)	Volume (cm^3)	$\frac{m}{V}$ (g/cm^3)
54.4	20.1	2.70
65.7	24.15	2.72
83.5	30.9	2.70
97.2	35.8	2.71
105.7	39.1	2.70

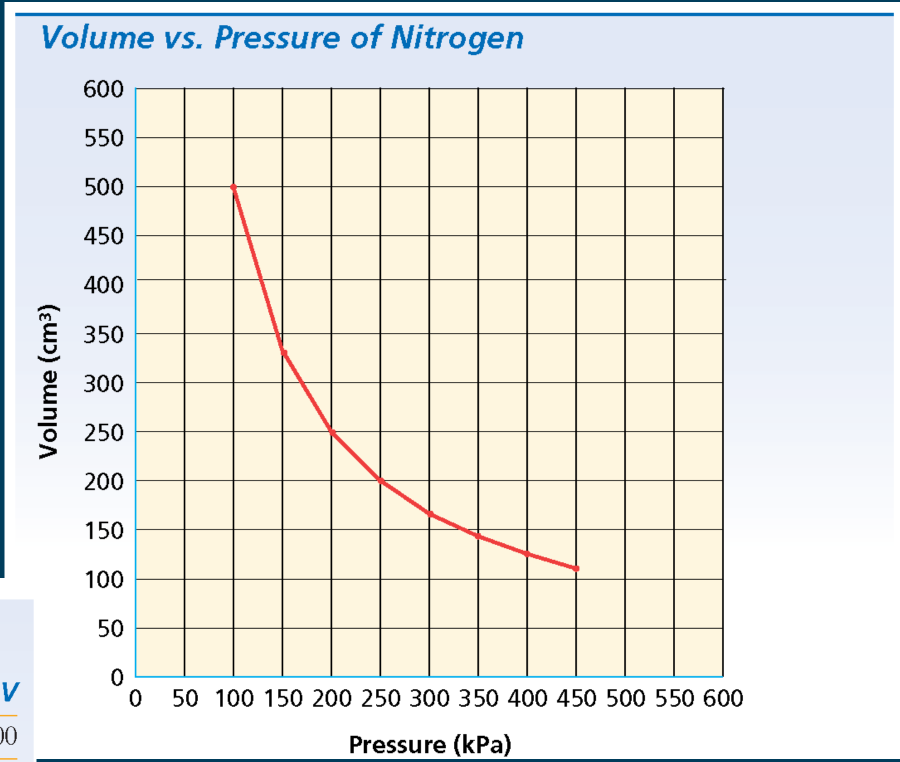
Inverse Proportions

- Two quantities are **inversely proportional** to each other if their product is constant.

- $y \propto \frac{1}{x}$

  - read as "y is proportional to 1 divided by x."

Inverse Proportion



Pressure-Volume Data for Nitrogen at Constant Temperature

Pressure (kPa)	Volume (cm^3)	$P \times V$
100	500	50 000
150	333	49 500
200	250	50 000
250	200	50 000
300	166	49 800
350	143	50 500
400	125	50 000
450	110	49 500