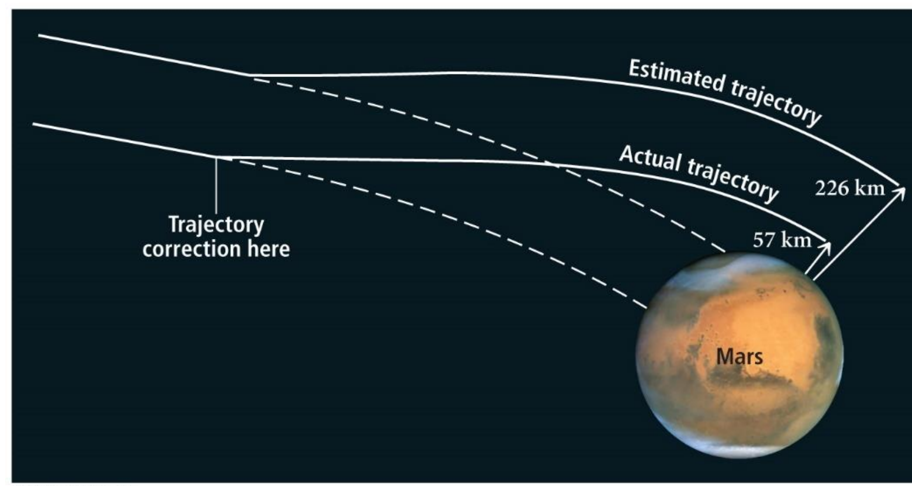
Chapter 2: Measurement, Problem Solving, and the Mole Concept

2.1: The Metric Mix-up: A $125 Million Unit Error

1998 – Mars Climate Orbiter

- Onboard Computers programmed in metric
- Ground engineers working in English units
- Corrections to trajectory 4.45 times too small
- Orbiter burned up in Mars' atmosphere



The Standard Units of Measurement

- Scientists have agreed on a set of international standard units for comparing all our measurements called the **SI units**
  - ✓ *Système International* = International System

Quantity	Unit	Symbol
length	meter	m
mass	kilogram	kg
time	second	s
temperature	kelvin	K

Temperature

- Measure of the average amount of kinetic energy caused by motion of the particles
  - ✓ higher temperature = larger average kinetic energy
- Heat flows from the matter that has ____________________

____

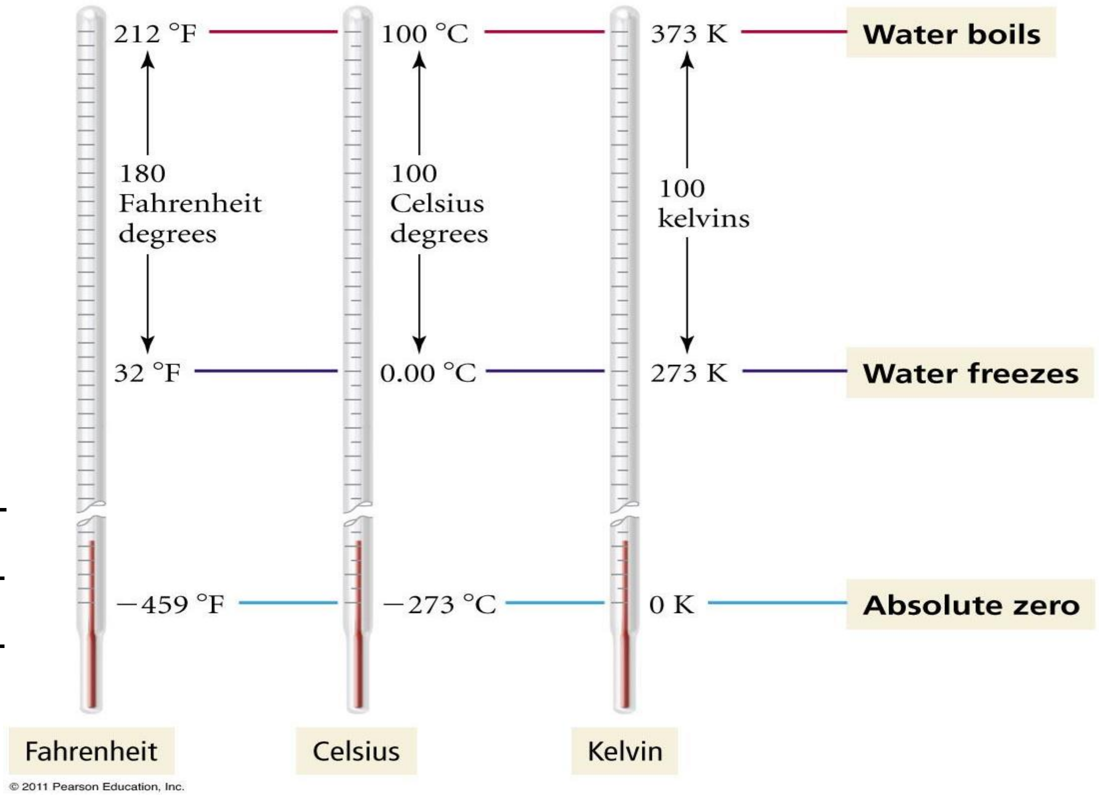
  - ✓ heat flows from hot object to cold
  - ✓ heat is exchanged through molecular collisions between the two materials

$$°C = \frac{(°F - 32)}{1.8}$$

$$K = °C + 273.15$$

Temperature Scales

- Fahrenheit scale, °F
  - ✓ used in the U.S.
- Celsius scale, °C
  - ✓ used in all other countries
- Kelvin scale, K
  - ✓ absolute scale
    - ➢ no negative numbers
  - ✓ ______________________________
    ____
    ____
  - ✓ 0 K = absolute zero



Common Prefix Multipliers in the SI System

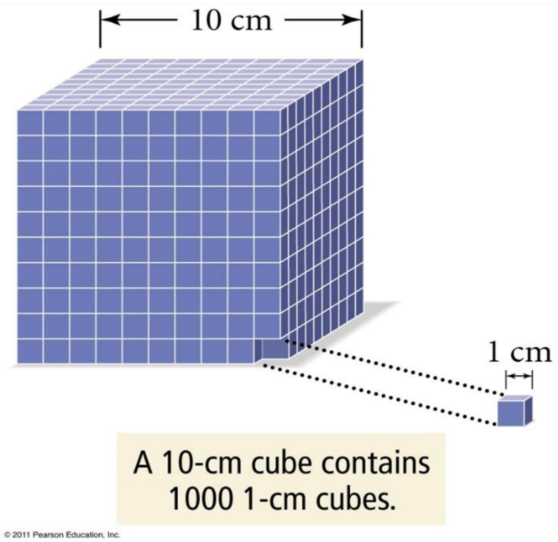
Prefix	Symbol	Decimal Equivalent	Power of 10
mega-	M	1,000,000	Base x 10^6
kilo-	k	1,000	Base x 10^3
deci-	d	0.1	Base x 10^{-1}
centi-	c	0.01	Base x 10^{-2}
milli-	m	0.001	Base x 10^{-3}
micro-	μ or mc	0.000 001	Base x 10^{-6}
nano-	n	0.000 000 001	Base x 10^{-9}
pico	p	0.000 000 000 001	Base x 10^{-12}

Volume

- Measure of the amount of space occupied
- SI unit = cubic meter (m^3)
- Commonly measure solid volume in cubic centimeters (cm^3)
- Commonly measure liquid or gas volume in milliliters (mL)

  ✓1 L is slightly larger than 1 quart
  ✓1 L = 1 dm^3 = 1000 mL = 10^3 mL
  ✓1 mL = 0.001 L = 10^{-3} L

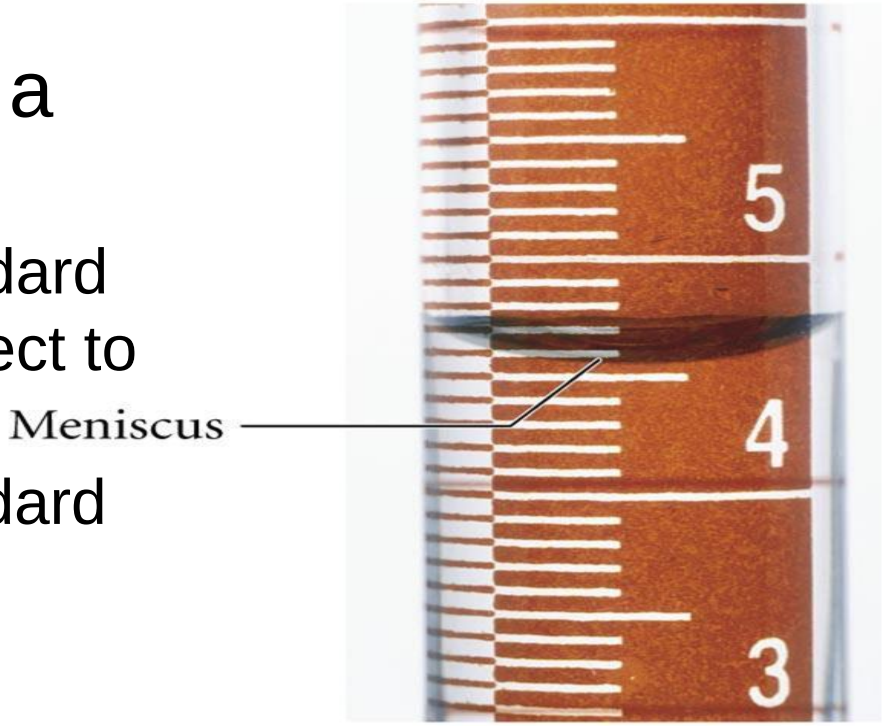
  ✓ **1 mL = 1 cm^3**



Measurement
and Significant Figures

What Is a Measurement?

- Quantitative observation
- Comparison to an agreed standard
- Every measurement has a number and a unit
  - The unit tells you what standard you are comparing your object to
  - The number tells you
    - what multiple of the standard the object measures
    - the uncertainty in the measurement



Reliability of Measurements: Precision and Accuracy

- Uncertainty comes from limitations of the instruments used for comparison, the experimental design, the experimenter, and nature's random behavior
- To understand how reliable a measurement is, we need to understand the limitations of the measurement
- **Accuracy**

  ____

  ____

  ____
- **Precision** is an indication of how close repeated measurements are to each other
  - ✓ how reproducible a measurement is

Precision and Accuracy

- Measurements are said to be
  - precise if they are consistent with one another;
  - accurate only if they are close to the actual value.

- Scientific measurements are reported so that __________

____

Consider the following reported value of 5.213:

- The **first three digits are certain**; the **last digit is estimated**.

Estimated value

5.213

Known with certainty

Estimation in Weighing

(a)

Markings every 1 g
Estimated reading 1.2 g

(b)

Markings every 0.1 g
Estimated reading 1.27 g

Estimation in Weighing

(a)

Report as 10.5 g

(b)

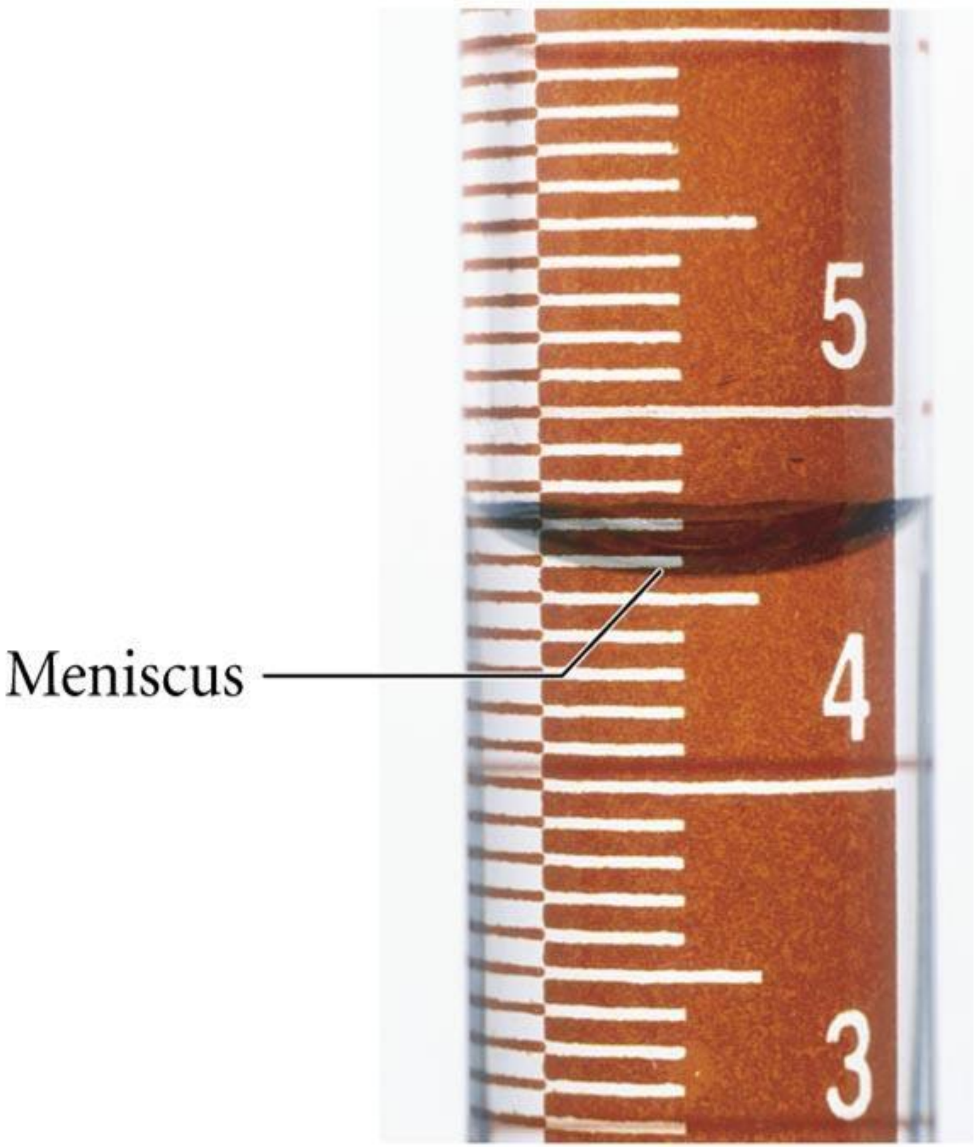
Report as 10.4977 g

Precision and Accuracy

Example 2.1

Reporting the Correct Number of Digits.

The graduated cylinder shown here has markings every 0.1 mL. Report the volume (which is read at the bottom of the meniscus) to the correct number of digits.



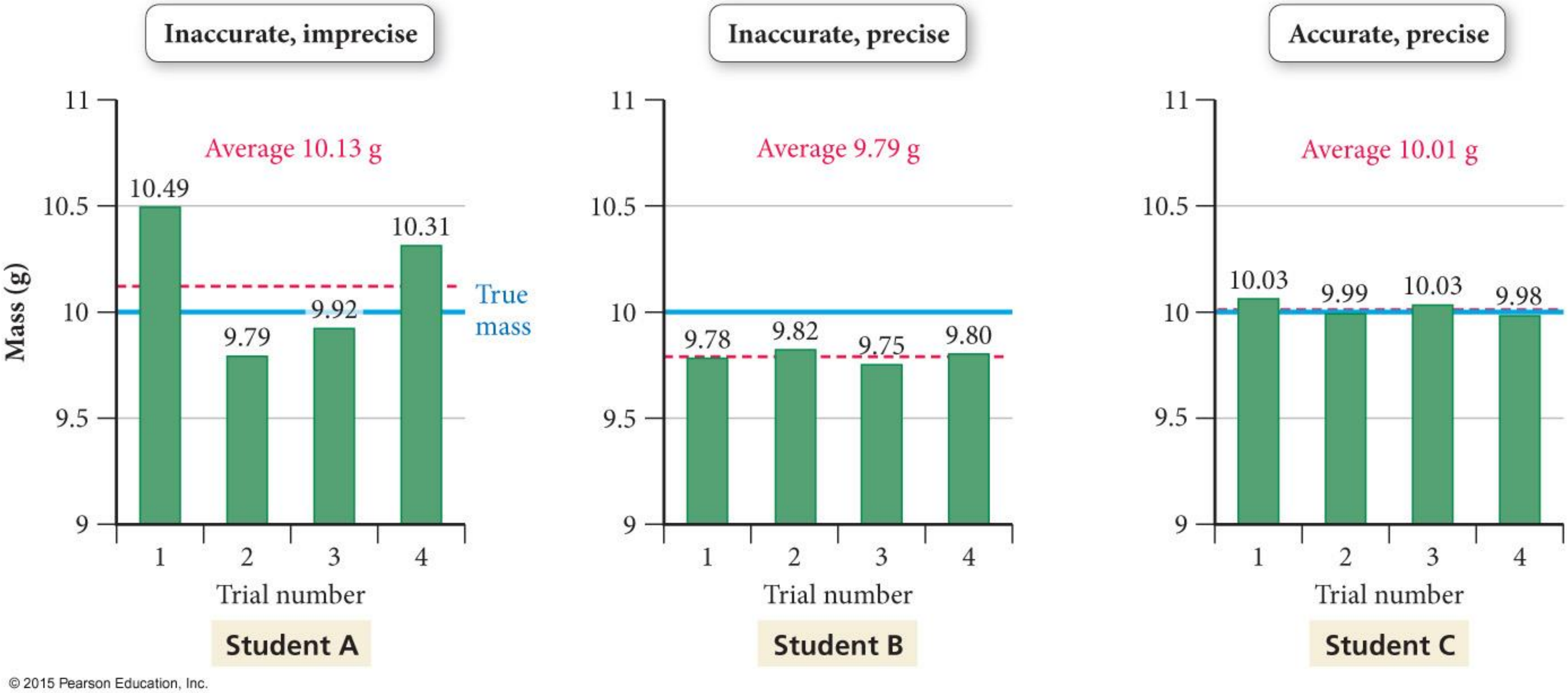
Precision and Accuracy: An Illustration Problem

Consider the results of three students who repeatedly weighed a lead block known to have a true mass of 10.00 g.

	Student A	Student B	Student C
Trial 1	10.49 g	9.78 g	10.03 g
Trial 2	9.79 g	9.82 g	9.99 g
Trial 3	9.92 g	9.75 g	10.03 g
Trial 4	10.31 g	9.80 g	9.98 g
Average	**10.13 g**	**9.79 g**	**10.01 g**

Precision and Accuracy: An Illustration Problem

Consider the results of three students who repeatedly



From the above data, what can you conclude about each of the students' recorded data?

Precision and Accuracy: An Illustration Problem

Lead block known to have a true mass of 10.00 g

- **<u>Student A's</u>** results are both __________________ (not close to the true value) and ________________________ (not consistent with one another).
  - ***Random error***

    ____

    ____

- **<u>Student B's</u>** results are ________________ (close to one another in value) but ____________________________.
  - ***Systematic error***

    ____

    ____

- **<u>Student C's</u>** results display little systematic error or random error—they are both _______________________ and _____________________.

Significant Figures

- Significant figures deal with writing numbers to reflect precision of their ______________________________.
- The precision of a **measurement** depends on the instrument used to make the measurement.
- The preservation of this precision during calculations can be accomplished by using significant figures.
- The greater the number of significant figures, the greater the certainty of the measurement.

Significant Figures

- The non-place-holding digits in a reported measurement are called **significant figures**
  - ✓ some zeros in a written number are only there to help you locate the decimal point

12.3 cm
has 3 sig. figs.
and its range is
12.2 to 12.4 cm

- Significant figures tell us the range of values to expect for repeated measurements
  - ✓ the more significant figures there are in a measurement, the smaller the range of values is

12.30 cm
has 4 sig. figs.
and its range is
12.29 to 12.31 cm

Rules of Significant Figures

1. Nonzero digits are always significant.

96	2 significant digits
61.4	3 significant digits

2. Zeros that are “sandwiched” between nonzero digits are significant.

5.02	3 significant digits
6004	4 significant digits

3. Zeros used as placeholders are NOT significant.

7000	1 significant digit
0.00783	3 significant digits

4. One or more final zeros used after the decimal point are significant.

4.7200	5 significant digits
0.250	3 significant digits

Using Significant Figures in Mathematical Operations:

<u>Multiplication and Division:</u>

The answer has the same number of significant figures as the least precise factor in the calculations.

3.05	x	1.3	=	3.965	= 4.0	correct ans
3 sig figs		2 sig figs		answer in calc		w/ 2 sig figs

<u>9.247 g</u> (4 sig figs)	=	.684962=	<u>0.685 g</u>	correct ans
13.5 cm^3 (3 sig figs)		(ans in calc)	cm^3	w/ 3 sig figs

Review

How many significant figures are in each of the following?

0.04450 m

5.0003 km

10 dm = 1 m

1.000×10^5 s

0.00002 mm

10,000 m

Exact Numbers

- Exact numbers have an unlimited number of significant figures.

- Exact counting of discrete objects
- Integral numbers that are part of an equation
- Defined quantities

- Some conversion factors are defined quantities, while others are not.

Intensive and Extensive Properties

- **Extensive properties** are properties whose value depends on amount of the substance
  - ✓extensive properties <u>canno</u>t be used to identify what *type* of matter something is
    - ➢if you are given a large glass containing 100 g of a clear, colorless liquid and a small glass containing 25 g of a clear, colorless liquid, are both liquids the same stuff?

- **Intensive properties** are properties whose value is independent of the amount of the substance
  - ✓intensive properties are often used to identify the type of matter
    - ➢samples with identical intensive properties are usually the same material

Density

Density = $\frac{\textbf{mass}}{\textbf{volume}}$ **Density (*d*) = $\frac{m}{V}$**

Density is a physical property: the ratio of mass to volume

- is an **intensive property**
  - The physical properties of **mass** and **volume** that determine a substance's density are EXTENSIVE.
- Units of Density

- Solids = g/cm^3 Liquids = g/mL Gases = g/L
  ✓ 1 cm^3 = 1 mL

- Volume of a solid can be determined by water displacement

- Density : solids > liquids >>> gases
  ✓ except ice is less dense than liquid water!

Density

$$\text{Density} = \frac{\text{Mass}}{\text{Volume}}$$

- For equal volumes, denser object has larger mass
- For equal masses, denser object has smaller volume
- Heating an object generally causes it to expand, therefore the density changes with temperature

TABLE 1.4 The Density of Some Common Substances at 20 °C

Substance	Density (g/cm^3)
Charcoal (from oak)	0.57
Ethanol	0.789
Ice	0.917 (at 0 °C)
Water	1.00 (at 4 °C)
Sugar (sucrose)	1.58
Table salt (sodium chloride)	2.16
Glass	2.6
Aluminum	2.70
Titanium	4.51
Iron	7.86
Copper	8.96
Lead	11.4
Mercury	13.55
Gold	19.3
Platinum	21.4

Calculations and Solving Chemical Problems

- Many problems in science involve using relationships to convert one unit of measurement to another
  - **unit conversion problems**.
- Using units as a guide to solving problems is
  - **dimensional analysis**.
- Units should always be included in calculations; they are multiplied, divided, and canceled like any other algebraic quantity.

Dimensional Analysis

- A **unit equation** is a statement of two equivalent quantities, such as

2.54 cm = 1 in.

- A **conversion factor** is a unit equation written in fraction form with the units we are ***converting from*** on the bottom and the units we are ***converting to*** on the top.

$$\frac{2.54\ \mathrm{cm}}{1\ \mathrm{in}} \quad \text{or} \quad \frac{1\ \mathrm{in}}{2.54\ \mathrm{cm}}$$

- Conversion factors are relationships between two units
  - ✓ may be exact or measured

Problem Solving and Dimensional Analysis

- Arrange conversion factors so the starting unit cancels
  - ✓ arrange conversion factors so the starting unit is on the bottom of the first conversion factor
- May string conversion factors
  - ✓ so you do not need to know every relationship, as long as you can find something else the starting and desired units are related to

$$\cancel{\text{given unit}} \times \frac{\text{desired unit}}{\cancel{\text{given unit}}} = \text{desired unit}$$

$$\cancel{\text{given unit}} \times \frac{\cancel{\text{related unit}}}{\cancel{\text{given unit}}} \times \frac{\text{desired unit}}{\cancel{\text{related unit}}} = \text{desired unit}$$

Dimensional Analysis

Units Raised to a Power:

- When building conversion factors for units raised to a power, remember to raise both the number and the unit to the power. For example, to convert from square inches to square centimeters, we construct the conversion factor as follows:

$$2.54\ \text{cm} = 1\ \text{in}$$

$$(2.54\ \text{cm})^2 = (1\ \text{in})^2$$

$$(2.54)^2\ \text{cm}^2 = 1^2\ \text{in}^2$$

$$6.45\ \text{cm}^2 = 1\ \text{in}^2$$

$$\frac{6.45\ \text{cm}^2}{1\ \text{in}^2} = 1$$

Problem Solving: Dimensional Analysis

Example:

The engineers involved in the Mars Climate Orbiter disaster entered the trajectory corrections in units of pound·second. Which conversion factor should they have multiplied their values by to conver them to the correcdt uniots of newton.second?

(1 pound·second = 4.45 newton·second)

(a) $\dfrac{1\text{ pound}\cdot\text{second}}{4.45\text{ newton}\cdot\text{second}}$

(b) $\dfrac{4.45\text{ newton}\cdot\text{second}}{1\text{ pound}\cdot\text{second}}$

(c) $\dfrac{1\text{ newton}\cdot\text{second}}{4.45\text{ pound}\cdot\text{second}}$

(d) $\dfrac{4.45\text{ pound}\cdot\text{second}}{1\text{ newton}\cdot\text{second}}$

Problem-Solving Strategy

- Identify the starting point (the ***given*** information).
  - Sort out information given in the problem.
- Identify the endpoint (what we must ***find***).
  - What is the problem asking you to solve for? What units does the answer need?
- Devise a way to use the given information to get the answer.
- Solve:
  - Most chemistry problems you will solve in this course are **unit conversion problems**.
  - Using units as a guide to solving problems (**dimensional analysis**)
    - Units should always be included in calculations; they are multiplied, divided, and canceled like any other algebraic quantity.
- Check whether the numerical value and its units make sense.

Example 2.3: Convert 1.76 yards to centimeters.

Note: 1.094 yd = 1m and 1 cm = 10^{-2} m

1. Sort into
   a. *Given*
   b. *Find*
2. Strategize: Devise a *conceptual plan* from the *given* units, using the appropriate conversion factors and ending with the *desired* units.
3. Solve: Begin with the *given* quantity. Multiply by the appropriate conversion factors, canceling units to arrive at the *find* quantity. Round to correct number of significant figures.
4. Check: Correct units? Does the answer make sense?

1. Sort
2. Strategize
3. Solve
4. Check

1. Sort
2. Strategize
3. Solve
4. Check

1. Sort
2. Strategize
3. Solve
4. Check

1. Sort

2. Strategize

3. Solve

4. Check

Moles

Counting Atoms by Moles

- If we can find the mass of a particular number of atoms, we can use this information to convert the mass of an element sample into the number of atoms in the sample

- A **mole** (mol) of anything contains 6.02214×10^{23} of those things.
  - Examples:
    - 1 mol of marbles corresponds to 6.02214×10^{23} marbles.
    - 1 mol of sand grains corresponds to 6.02214×10^{23} sand grains.

- This number is **Avogadro's number**.

Twenty-two copper pennies contain approximately 1 mol of copper atoms.

Chemical Packages - The Mole

- **Mole** = number of particles equal to the number of atoms in 12 g of C-12
  - ✓ 1 atom of C-12 weighs exactly 12 amu
  - ✓ 1 mole of C-12 weighs exactly 12 g

- The number of particles in 1 mole is called **Avogadro's Number = 6.0221421 x 10^{23}**
  - ✓ 1 mole of C atoms weighs 12.01 g and has 6.022 x 10^{23} atoms
    - ➢ the average mass of a C atom is 12.01 amu

Mole Conversions: Atoms to Moles or Moles to Atoms

- Converting between number of moles and number of atoms is similar to converting between dozens of eggs and number of eggs.

- For atoms, you use the conversion factor
  1 mol atoms = 6.022×10^{23} atoms.

- The conversion factors take the following forms:

$$\frac{1 \text{ mol atoms}}{6.022 \times 10^{23} \text{ atoms}} \quad \text{or} \quad \frac{6.022 \times 10^{23} \text{ atoms}}{1 \text{ mol atoms}}$$

Practice — A silver ring contains 1.1 x 10^{22} silver atoms. How many moles of silver are in the ring?

Converting between Mass and Amount (Number of Moles)

- To count atoms by weighing them, we need one other conversion factor—the mass of 1 mol of atoms.
- The mass of 1 mol of atoms of an element is the **molar mass**.
- **An element's molar mass in grams per mole is numerically equal to the element's atomic mass in atomic mass units (amu).**
- The lighter the atom, the less a mole weighs
- The lighter the atom, the more atoms there are in 1 g

Mole and Mass Relationships

Substance	***Weight of 1 atom***	***Pieces in 1 mole***	***Weight of 1 mole***
hydrogen	1.008 amu	6.022×10^{23} atoms	1.008 g
carbon	12.01 amu	6.022×10^{23} atoms	12.01 g
oxygen	16.00 amu	6.022×10^{23} atoms	16.00 g
sulfur	32.06 amu	6.022×10^{23} atoms	32.06 g
calcium	40.08 amu	6.022×10^{23} atoms	40.08 g
chlorine	35.45 amu	6.022×10^{23} atoms	35.45 g
copper	63.55 amu	6.022×10^{23} atoms	63.55 g

1 mole
sulfur
32.06 g

1 mole
carbon
12.01 g

Converting between Mass and Moles

- The molar mass of any element is the conversion factor between the mass (in grams) of that element and the amount (in moles) of that element.

- Example:

  12.01 g C atoms = 1 mol C atoms

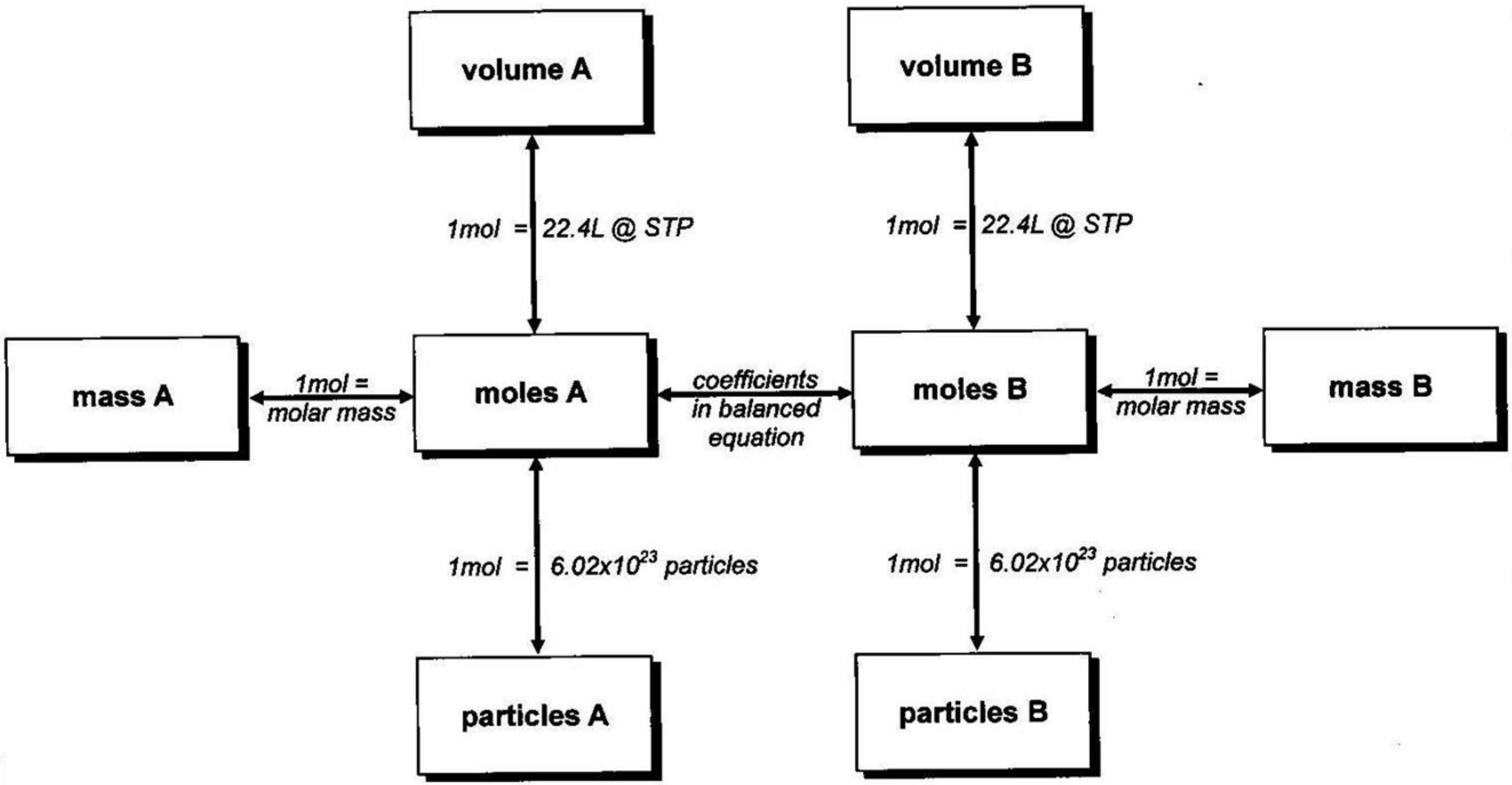
  or

  12.01 g C atoms/1 mol C atoms

  or

  1 mol C atoms/12.01 g C atoms

FLOWCHART FOR MOLE CONVERSIONS



Mass to Moles to Number of Particles: The Conceptual Plan

For an element,

Mass of element (grams) → Moles of element → Number of atoms

Divide by atomic mass

Multiply by Avogadro's number

For a molecule (compound),

Mass of molecule (grams) → Moles of molecules → Number of molecules

Divide by molar mass

Multiply by Avogadro's number

Number of Particles to Moles to Mass: The Conceptual Plan

For an element,

Number of atoms → **Moles of element** → **Mass of element (grams)**

Divide by Avogadro's number

Multiply by atomic mass

For a molecule (compound),

Number of molecules → **Moles of molecules** → **Mass of molecule (grams)**

Divide by Avogadro's number

Multiply by molar mass

Practice — Calculate the moles of sulfur in 57.8 g of sulfur

Practice — How many aluminum atoms are in a can weighing 16.2 g?