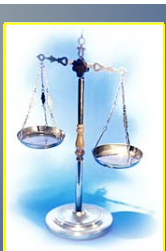


Chapter 3 “*Scientific Measurement*”

Click to add text

Section 3.1 Measurements and Their Uncertainty

- OBJECTIVES:
  - Convert measurements to scientific notation.

Section 3.1 Measurements and Their Uncertainty

- OBJECTIVES:
  - Distinguish among accuracy, precision, and error of a measurement.

Section 3.1 Measurements and Their Uncertainty

- OBJECTIVES:
  - Determine the number of significant figures in a measurement and in a calculated answer.

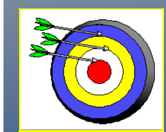
Measurements

- **Qualitative** measurements are words, such as heavy or hot
- **Quantitative** measurements involve numbers (quantities), and depend on:
  1) The reliability of the measuring instrument
  2) the care with which it is read – this is determined by YOU!
- **Scientific Notation**
  - Coefficient raised to power of 10
  - Reviewed earlier this semester!

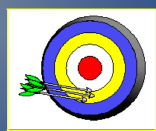
Accuracy, Precision, and Error

- It is necessary to make good, reliable measurements in the lab
- Accuracy – how close a measurement is to the true value
- Precision – how close the measurements are to each other (reproducibility)

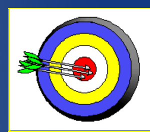
Precision and Accuracy



Neither accurate nor precise



Precise, but not accurate



Precise AND accurate

Accuracy, Precision, and Error

- Accepted value = the correct value based on reliable references
- Experimental value = the value measured in the lab

Accuracy, Precision, and Error

- Error = accepted value – exp. value
  - Can be positive or negative
- Percent error = the ***absolute value*** of the error divided by the accepted value, then multiplied by 100%

$$\% \text{ error} = \frac{|\text{ error }|}{\text{accepted value}} \times 100\%$$

Why Is there Uncertainty?

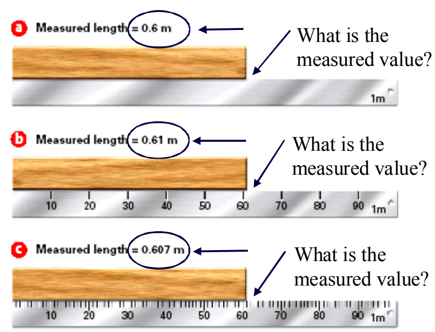
- Measurements are performed with instruments, and no instrument can read to an infinite number of decimal places
- *Which of the balances shown* has the greatest uncertainty in measurement?

Significant Figures in Measurements

- **Significant figures** in a measurement include all of the digits that are known, plus one more digit that is estimated.
- Measurements must be reported to the correct number of significant figures.

Figure 3.5 Significant Figures - Page 67

Which measurement is the best?



Rules for Counting Significant Figures

Non-zeros always count as significant figures:

3456 has

4 significant figures

Rules for Counting Significant Figures

Zeros

Leading zeroes do not count as significant figures:

0.0486 has

3 significant figures

Rules for Counting Significant Figures

Zeros

Captive zeroes always count as significant figures:

16.07 has

4 significant figures

Rules for Counting Significant Figures

Zeros

Trailing zeros are significant only if the number contains a written decimal point:

9.300 has

4 significant figures

Rules for Counting Significant Figures

Two special situations have an ***unlimited*** number of significant figures:

1. Counted items
   a) 23 people, or 425 thumbtacks
2. Exactly defined quantities
   b) 60 minutes = 1 hour

Sig Fig Practice #1

How many significant figures in the following?

1.0070 m → 5 sig figs

17.10 kg → 4 sig figs

100,890 L → 5 sig figs

3.29×10^3 s → 3 sig figs

0.0054 cm → 2 sig figs

3,200,000 → 2 sig figs

These come from measurements

5 dogs → unlimited ← This is a counted value

Significant Figures in Calculations

- In general a calculated answer <u>cannot</u> be more precise than the ***least precise*** measurement from which it was calculated.
- Ever heard that a chain is only as strong as the weakest link?
- Sometimes, calculated values need to be *rounded off*.

Rounding Calculated Answers

- **<u>Rounding</u>**
  - Decide *how many* significant figures are needed (more on this very soon)
  - Round to that many digits, counting from the *left*
  - Is the next digit less than 5? Drop it.
  - Next digit 5 or greater? Increase by 1

SAMPLE PROBLEM 3.1 - Page 69

Rounding Measurements

Round off each measurement to the number of significant figures shown in parentheses. <u>Write the answers in scientific notation.</u>

a. 314.721 meters (four)
b. 0.001 775 meter (two)
c. 8792 meters (two)

Be sure to answer the question completely!

1 Analyze *Identify the relevant concepts.*
Round off each measurement to the number of significant figures indicated. Then apply the rules for expressing numbers in scientific notation.

Rounding Calculated Answers

- <u>Addition and Subtraction</u>
  - The answer should be rounded to the same number of **<u>decimal places</u>** as the *least* number of decimal places in the problem.

SAMPLE PROBLEM 3.2 - Page 70

Significant Figures in Addition

Calculate the sum of the three measurements. Give the answer to the correct number of significant figures.

12.52 meters + 349.0 meters + 8.24 meters

1 Analyze *Identify the relevant concepts.*
Calculate the sum and then analyze each measurement to determine the number of decimal places required in the answer.

Rounding Calculated Answers

- <u>Multiplication and Division</u>
  - Round the answer to the same number of **<u>significant figures</u>** as the <u>least</u> number of significant figures in the problem.

SAMPLE PROBLEM 3.3 - Page 71

Significant Figures in Multiplication and Division

Perform the following operations. Give the answers to the correct number of significant figures.

a. 7.55 meters × 0.34 meter

b. 2.10 meters × 0.70 meter

c. 2.4526 meters ÷ 8.4

1 Analyze *Identify the relevant concepts.*

Perform the required math operation and then analyze each of the original numbers to determine the correct number of significant figures required in the answer.

Rules for Significant Figures in Mathematical Operations

- Multiplication and Division: # sig figs in the result equals the number in the *least precise* measurement used in the calculation.
  - 6.38 x 2.0 =
  - 12.76 → 13 (2 sig figs)

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Sig Fig Practice #2

Calculation	Calculator says:	Answer
3.24 m x 7.0 m	22.68 m^2	23 m^2
100.0 g ÷ 23.7 cm^3	4.219409283 g/cm^3	4.22 g/cm^3
0.02 cm x 2.371 cm	0.04742 cm^2	0.05 cm^2
710 m ÷ 3.0 s	236.6666667 m/s	240 m/s
1818.2 lb x 3.23 ft	5872.786 lb·ft	5870 lb·ft
1.030 g x 2.87 mL	2.9561 g/mL	2.96 g/mL

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Rules for Significant Figures in Mathematical Operations

- Addition and Subtraction: The number of decimal places in the result equals the number of decimal places in the *least precise* measurement.
  - 6.8 + 11.934 =
  - 18.734 → 18.7 (3 sig figs)

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Sig Fig Practice #3

Calculation	Calculator says:	Answer
3.24 m + 7.0 m	10.24 m	10.2 m
100.0 g - 23.73 g	76.27 g	76.3 g
0.02 cm + 2.371 cm	2.391 cm	2.39 cm
713.1 L - 3.872 L	709.228 L	709.2 L
1818.2 lb + 3.37 lb	1821.57 lb	1821.6 lb
2.030 mL - 1.870 mL	0.16 mL	0.160 mL

*Note the zero that has been added.

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Section 3.3 The International System of Units

- OBJECTIVES:
  - List SI units of measurement and common SI prefixes.

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Section 3.3
The International System of Units

- OBJECTIVES:
  - Distinguish between the ***mass*** and ***weight*** of an object.

Section 3.3
The International System of Units

- OBJECTIVES:
  - Convert between the Celsius and Kelvin temperature scales.

International System of Units

- Measurements depend upon units that serve as reference standards
- The standards of measurement used in science are those of the Metric System

International System of Units

- Metric system is now revised and named as the International System of Units (SI), as of 1960
- It has simplicity, and is based on 10 or multiples of 10
- **7 base units,** but only five commonly used in chemistry: meter, kilogram, kelvin, second, and mole.

The Fundamental SI Units
(Le Système International, SI)

Physical Quantity	Name	Abbreviation
Length	Meter	m
Mass	Kilogram	kg
Temperature	Kelvin	K
Time	Second	s
Amount of substance	Mole	mol
Not commonly used in chemistry:		
Luminous intensity	Candela	cd
Electric current	Ampere	A

Nature of Measurements

Measurement - quantitative observation consisting of 2 parts:

- Part 1 – **number**
- Part 2 - scale (unit)
- Examples:
  - **20** grams
  - **6.63 x 10^{34}** Joule seconds

International System of Units

- Sometimes, non-SI units are used
  - Liter, Celsius, calorie
- Some are **derived** units
  - They are made by joining other units
  - Speed = miles/hour (distance/time)
  - Density = grams/mL (mass/volume)

Length

- In SI, the basic unit of length is the **meter (m)**
  - Length is the distance between two objects – measured with ruler
- We make use of prefixes for units larger or smaller

SI Prefixes – Page 74 Common to Chemistry

Prefix	Unit Abbreviation	Meaning	Exponent
Kilo-	k	thousand	10^{3}
Deci-	d	tenth	10^{-1}
Centi-	c	hundredth	10^{-2}
Milli-	m	thousandth	10^{-3}
Micro-	μ	millionth	10^{-6}
Nano-	n	billionth	10^{-9}

Volume

- The space occupied by any sample of matter.
- Calculated for a solid by multiplying the length x width x height; thus derived from units of length.
- SI unit = **cubic meter (m^3)**
- Everyday unit = Liter (L), which is non-SI. (Note: 1mL = $1cm^3$)

Devices for Measuring Liquid Volume

- Graduated cylinders
- Pipets
- Burets
- Volumetric Flasks
- Syringes

The Volume Changes!

- Volumes of a solid, liquid, or gas will *generally* increase with temperature
- Much more prominent for GASES
- Therefore, measuring instruments are calibrated for a specific temperature, usually **20 ^{0}C**, which is about room temperature

Units of Mass

- **Mass** is a measure of the quantity of matter present
  - *Weight* is a force that measures the pull by gravity- it changes with location
- Mass is constant, regardless of location

Working with Mass

- The SI unit of mass is the **kilogram (kg),** even though a more convenient everyday unit is the gram
- Measuring instrument is the balance scale

Units of Temperature

- Temperature is a measure of how hot or cold an object is. **(Measured with a thermometer.)**
- Heat moves from the object at the higher temperature to the object at the lower temperature.
- We use two units of temperature:
  - Celsius – named after Anders Celsius
  - Kelvin – named after Lord Kelvin

Units of Temperature

- Celsius scale defined by two readily determined temperatures:
  - Freezing point of water = 0 °C
  - Boiling point of water = 100 °C
- Kelvin scale does not use the degree sign, but is just represented by K
  - absolute zero = 0 K (thus no negative values)
  - formula to convert: K = °C + 273

SAMPLE PROBLEM 3.4 - Page 78

Converting Between Temperature Scales

Normal human body temperature is 37°C. What is that temperature in kelvins?

1 **Analyze** *List the known and the unknown.*

Known

- Temperature in °C = 37°C

Unknown

- Temperature in K = ? K

Use the known value and the equation K = °C + 273 to calculate the temperature in kelvins.

Units of Energy

- Energy is the capacity to do work, or to produce heat.
- Energy can also be measured, and two common units are:
  1) Joule (J) = the SI unit of energy, named after James Prescott Joule
  2) calorie (cal) = the heat needed to raise 1 gram of water by 1 °C

Units of Energy

- Conversions between joules and calories can be carried out by using the following relationship:

$$1\ \text{cal} = 4.184\ \text{J}$$

Section 3.3 Conversion Problems

- OBJECTIVE:
  - Construct conversion factors from equivalent measurements.

Section 3.3 Conversion Problems

- OBJECTIVE:
  - Apply the techniques of dimensional analysis to a variety of conversion problems.

Section 3.3 Conversion Problems

- OBJECTIVE:
  - Solve problems by breaking the solution into steps.

Section 3.3 Conversion Problems

- OBJECTIVE:
  - Convert complex units, using dimensional analysis.

Conversion factors

- A "ratio" of equivalent measurements
- Start with two things that are the same: one meter is one hundred centimeters
- write it as an equation

  1 m = 100 cm
- We can divide on each side of the equation to come up with two ways of writing the number "1"

Conversion factors

$$\frac{1\text{ m}}{100\text{ cm}} = \frac{\cancel{100\text{ cm}}}{\cancel{100\text{ cm}}}$$

Conversion factors

$$\frac{1\text{ m}}{100\text{ cm}} = 1$$

Conversion factors

$$\frac{1\text{ m}}{100\text{ cm}} = 1$$

$$\frac{\cancel{1\text{ m}}}{\cancel{1\text{ m}}} = \frac{100\text{ cm}}{1\text{ m}}$$

Conversion factors

$$\frac{1\text{ m}}{100\text{ cm}} = 1$$

$$1 = \frac{100\text{ cm}}{1\text{ m}}$$

Conversion factors

- A unique way of writing the number 1
- In the same system they are <u>defined</u> quantities so they have an *unlimited number of significant figures*
- Equivalence statements always have this relationship:

big # small unit = small # big unit

1000 mm = 1 m

Practice by writing the <u>two possible conversion factors</u> for the following:

- Between kilograms and grams
- between feet and inches
- using 1.096 qt. = 1.00 L

What are they good for?

- We can multiply by the number "one" creatively to change the units.
- Question: 13 inches is how many yards?
- We know that 36 inches = 1 yard.
- 13 inches x $\frac{\text{1 yard}}{\text{36 inches}}$ =

What are they good for?

- We can multiply by a conversion factor to change the units .
- Problem: 13 inches is how many yards?
- Known: 36 inches = 1 yard.
- $\frac{\text{1 yard}}{\text{36 inches}}$ = 1
- 13 ~~inches~~ x $\frac{\text{1 yard}}{\text{36 inches}}$ = 0.36 yards

Conversion factors

- Called <u>conversion factors</u> because they allow us to convert units.
- really just multiplying by ***one***, in a creative way.

Dimensional Analysis

- A way to analyze and solve problems, by using units (or dimensions) of the measurement
- Dimension = a unit (such as g, L, mL)
- Analyze = to solve
  - Using the units to solve the problems.
- If the units of your answer are right, chances are you did the math right!

Dimensional Analysis

- Dimensional Analysis provides an alternative approach to problem solving, instead of with an equation or algebra.
- A ruler is 12.0 inches long. How long is it in cm? (1 inch = 2.54 cm)
- How long is this in meters?
- A race is 10.0 km long. How far is this in miles, if:
  - 1 mile = 1760 yards
  - 1 meter = 1.094 yards

Converting Between Units

- Problems in which measurements with one unit are converted to an equivalent measurement with another unit are easily solved using dimensional analysis
- Sample: Express 750 dg in grams.
- Many complex problems are best solved by breaking the problem into manageable parts.

Converting Between Units

- Let's say you need to clean your car:
  1) Start by vacuuming the interior
  2) Next, wash the exterior
  3) Dry the exterior
  4) Finally, put on a coat of wax
- What problem-solving methods can help you solve complex word problems?
  - ✓ Break the solution down into steps, and use more than one conversion factor if necessary

Converting Complex Units?

- Complex units are those that are expressed as a ratio of two units:
  - Speed might be meters/hour
- Sample: Change 15 meters/hour to units of centimeters/second
- How do we work with units that are squared or cubed? (cm^3 to m^3, etc.)

SAMPLE PROBLEM 3.9 - Page 86

Converting Ratios of Units

The mass per unit volume of a substance is a property called density. The density of manganese, a metallic element, is 7.21 g/cm^3. What is the density of manganese expressed in units kg/m^3?

1 Analyze *List the knowns and the unknown.*

Knowns
- density of manganese = 7.21 g/cm^3
- 10^3 g = 1 kg
- 10^6 cm^3 = 1 m^3

Unknown
- density manganese = ? kg/m^3

The desired conversion is $g/cm^3 \longrightarrow kg/m^3$. The mass unit in the numerator must be changed from grams to kilograms: $g \longrightarrow kg$. In the denominator, the volume unit must be changed from cubic centimeters to cubic meters: $cm^3 \longrightarrow m^3$. Note that the relationship between cm^3 and m^3 was determined from the relationship between cm and m. Cubing the relationship 10^2 cm = 1 m yields $(10^2\ cm)^3 = (1\ m)^3$, or $10^6\ cm^3 = 1\ m^3$.

Section 3.4 Density

- OBJECTIVES:
  - Calculate the density of a material from experimental data.

Section 3.4 Density

- OBJECTIVES:
  - Describe how density varies with temperature.

Density

- Which is heavier- a pound of lead or a pound of feathers?
  - Most people will answer lead, but the weight is exactly the same
  - They are normally thinking about equal volumes of the two
- The relationship here between mass and volume is called Density

Density

- The formula for density is:

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

- Common units are: **g/mL**, or possibly **g/cm³**, (or **g/L** for gas)
- Density is a physical property, and does not depend upon sample size

Table 3.6 - Page 90

Note temperature and density units

Densities of Some Common Materials

Solids and Liquids		Gases	
Material	Density at 20°C (g/cm³)	Material	Density at 20°C (g/L)
Gold	19.3	Chlorine	2.95
Mercury	13.6	Carbon dioxide	1.83
Lead	11.4	Argon	1.66
Aluminum	2.70	Oxygen	1.33
Table sugar	1.59	Air	1.20
Corn syrup	1.35–1.38	Nitrogen	1.17
Water (4°C)	1.000	Neon	0.84
Corn oil	0.922	Ammonia	0.718
Ice (0°C)	0.917	Methane	0.665
Ethanol	0.789	Helium	0.166
Gasoline	0.66–0.69	Hydrogen	0.084

Density and Temperature

- What happens to the density as the temperature of an object increases?
  - Mass remains the same
  - Most substances increase in volume as temperature increases
- Thus, density generally decreases as the temperature increases

Density and Water

- **Water** is an important exception to the previous statement.
- Over certain temperatures, the volume of water increases as the temperature decreases (Do you want your water pipes to freeze in the winter?)
  - Does ice float in liquid water?
  - Why?

SAMPLE PROBLEM 3.10 - Page 91

Calculating Density

A copper penny has a mass of 3.1 g and a volume of 0.35 cm³. What is the density of copper?

1 Analyze *List the knowns and the unknown.*

Knowns	Unknown
• mass = 3.1 g	• density = ? g/cm³
• volume = 0.35 cm³	

Use the known values and the following definition of density.

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

SAMPLE PROBLEM 3.11 - Page 92

Using Density to Calculate Volume

What is the volume of a pure silver coin that has a mass of 14 g? The density of silver (Ag) is 10.5 g/cm³.

1 Analyze *List the knowns and the unknown.*

Knowns	Unknown
• mass of coin = 14 g	• volume of coin = ? cm³
• density of silver = 10.5 g/cm³	

You can solve this problem by using density as a conversion factor. You need to convert the mass of the coin into a corresponding volume. The density gives the following relationship between volume and mass.

$$1\ \text{cm}^3\ \text{Ag} = 10.5\ \text{g Ag}$$

Based on this relationship, you can write the following conversion factor.

$$\frac{1\ \text{cm}^3\ \text{Ag}}{10.5\ \text{g Ag}}$$

Notice that the known unit is in the denominator and the unknown unit is in the numerator.