Chemistry 104 Chapter Two PowerPoint Notes

Measurements in Chemistry
Chapter 2

Chemistry 104
Professor Michael Russell

Physical Quantities

Measurable physical properties such as height, volume, and temperature are called ***Physical quantity***. A number and a unit of defined size is required to describe physical quantity.

Number Unit

61.2 kilograms

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Physical Quantities

A number alone doesn't say much!

Say an average textbook weighs 1.

The question would then be asked 1 what? 1 pound? 1 kilogram? 1 ounce?

You have to mention the unit of mass *along with the number* for the statement to be meaningful.

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Physical Quantities

Physical quantities measured using many different units. Mass can be measured in pounds, kilograms, ounces, etc.

To avoid confusion, scientists around the world have agreed to use a set of standard units known as the ***International System of Units*** or ***SI units*** for some common physical quantities.

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SI Units

In **SI Units**,

- mass measured in kilograms (kg)
- length measured in meters (m)
- volume measured in cubic meters (m^3)
- time measured in seconds (s).

Many other units derived from SI units.

- speed measured in meters per second (m/s)
- density measured in grams per cubic centimeter (g/cm^3).

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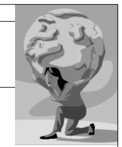
Measuring Mass



Mass is a measure of amount of matter in an object.

Weight is a measure of gravitational pull on an object.

At the same location, two objects with *identical masses* have *identical weights* (gravity pulls them equally).

Thus masses of objects determined by comparing the weight of the object to the weight of a known reference.

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Measuring Length

The Meter (m) is the standard measure of length or distance in both SI and metric system. One meter is 39.37 inches.

Centimeter (cm; $^1/_{100}$ m) and millimeter (mm; $^1/_{1000}$ m) commonly used for most measurements in chemistry and medicine.

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Measuring Volume

Volume is the amount of space occupied by an object.

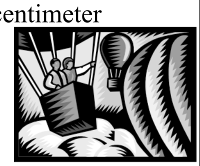
SI unit for volume is the cubic meter (m^3)

Liter (L) is commonly used in chemistry.

1 L = 0.001 m^3 = 1000 mL

A milliliter is often called a cubic centimeter

1 mL = 1 cm^3



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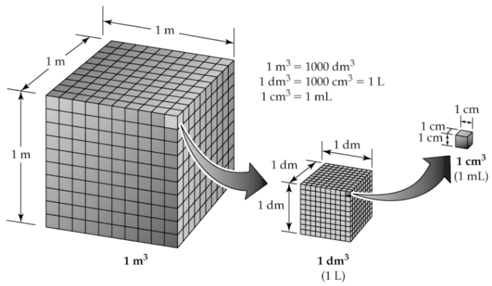

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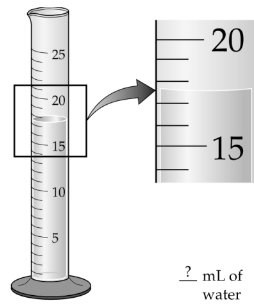
Measurement and Significant Figures

Every experimental measurement, no matter how precise, has a degree of uncertainty because there is a limit to the number of digits that can be determined.



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Measurement and Significant Figures

To indicate the precision of the measurement, the value recorded should use all the digits known with certainty *plus* one additional estimated digit that usually is considered *uncertain* by plus or minus 1 ($\pm$ 1)

The total number of digits used to express such a measurement is called the number of ***significant figures***.

Example: The quantity 65.07 g has four significant figures.

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Uncertain digit

54.07 g

A mass between 54.06 g and 54.08 g ($\pm$0.01 g)

Uncertain digit

54.071 38 g

A mass between 54.071 37 g and 54.071 39 g ($\pm$0.000 01 g)

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Rules for Determining Significant Figures

1. Zeroes in the middle of a number are significant. 69.08 g has four significant figures, 6, 9, 0, and 8.
2. Zeroes at the beginning of a number are not significant. 0.0089 g has two significant figure, 8 and 9.
3. Zeroes at the end of a number and after the decimal points are significant. 2.50 g has three significant figures 2, 5, and 0. 25.00 m has four significant figures 2, 5, 0, and 0.

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Rules for Determining Significant Figures

4. Zeroes at the end of a number and before an implied decimal points may or may not be significant. 1500 kg may have two, three, or four significant figures. Zeroes here may be part of the measurements or for simply to locate the unwritten decimal point.

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How Many Significant Figures?

94.072 g
0.0834 cm
0.02907 mL
138.200 m
23,000 kg
23,000. kg

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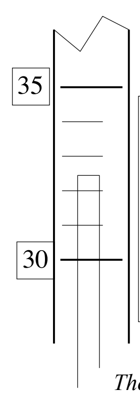
How Many Significant Figures?



Thermometer

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Scientific Notation

Scientific Notation is a convenient way to write a very small or a very large number.

Written as a product of a number between 1 and 10, times the number 10 raised to power.

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Two examples of converting standard numbers to scientific notations are shown below.

$$0.002\,15 = 2.15 \times \frac{1}{1000} = 2.15 \times \frac{1}{10 \times 10 \times 10}$$

$$= 2.15 \times \frac{1}{10^3} = 2.15 \times 10^{-3}$$

$$0.002\,15 = 2.15 \times 10^{-3}$$

Decimal point is moved three places to the right, so exponent is −3.

$$215. = 2.15 \times 10^2$$

Decimal point is moved two places to the left, so exponent is 2.

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Examples of converting scientific notations back to the standard numbers:

$$1.56 \times 10^{-8} = 0.000\,000\,015\,6$$

Negative exponent of –8,
so decimal point is moved to the left eight places.

$$3.7962 \times 10^4 = 37{,}962$$

Positive exponent of 4,
so decimal point is moved to the right four places.

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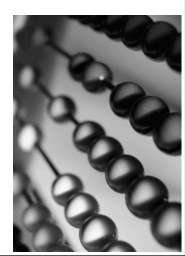
Rounding off Numbers

Often calculators produce large numbers as a result of a calculation *although* the number of significant figures is good only to a few numbers, less than the calculator has produced

In this case the large number may be rounded off to a smaller number *keeping only significant figures*.



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Rules for Rounding off Numbers

Rule 1 (For multiplication and divisions): The answer can't have more significant figures than either of the original numbers.

Three significant figures → 278 mi

Four significant figures → 11.70 gal

Three significant figures → 23.8 mi/gal

$$\frac{278 \text{ mi}}{11.70 \text{ gal}} = 23.8 \text{ mi/gal}$$

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Rules for Rounding off Numbers

Rule 2 (For addition and subtraction): The number can't have more digits after the decimal point than either of the original numbers.

Volume of water at start →	3.18? ?? L	← Two digits after decimal point
Volume of water addded →	+ 0.013 15 L	← Five digits after decimal point
Total volume of water →	3.19? ?? L	← Two digits after decimal point

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Rules for Rounding off Numbers

Rule 3: Once you decide how many numbers to keep, you *may* need to round off your answer:

If the first digit you remove is between 0 and 4, drop it and all remaining digits.

If the first digit you remove is between 5 and 9, round the number up by adding 1 to the digit to the left of the one you drop

Example: **2.4271** becomes **2.4** when rounded to two significant figures

Example: **4.5816** becomes **4.6** when rounded to two significant figures

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Problem Solving: Converting a Quantity from One Unit to Another

Factor-Label-Method: A quantity in one unit is converted to an equivalent quantity in a different unit by using a *conversion factor* that expresses the relationship between units.

Some conversion factors are considered *exact* and have *unlimited sig figs*.

Most conversion factors obey sig fig rules.

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Conversion factors between kilometers and miles

$$\frac{1\ \text{km}}{0.6214\ \text{mi}} = 1 \quad \text{or} \quad \frac{0.6214\ \text{mi}}{1\ \text{km}} = 1$$

These two quantities are the same.

These two quantities are the same.

$$\frac{1\ \text{km}}{0.6214\ \text{mi}} \quad \text{or} \quad \frac{0.6214\ \text{mi}}{1\ \text{km}}$$

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When solving a problem, set up an equation so that *all unwanted units cancel*, leaving only the desired unit. *For example*, we want to find out how many kilometers are there in 26.22 miles. We will get the correct answer if we multiply 26.22 mi by the conversion factor km/mi.

$$\underbrace{26.22\ \cancel{\text{mi}}}_{\text{Starting quantity}} \times \underbrace{\frac{1\ \text{km}}{0.6214\ \cancel{\text{mi}}}}_{\text{Conversion factor}} = \underbrace{42.20\ \text{km}}_{\text{Equivalent quantity}}$$

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Some **Exact Conversions**

Length
1 km = 1000 m = 10^5 cm = 10^{12} nm
12 in = 1 ft   5280 ft =1 mile
1 in = 2.54 cm

Volume
1 cm^3 = 1 mL

Mass
1 g = 1000 mg   1 kg = 1000 g   1 lb = 454 g

These conversions have ***unlimited*** sig figs by definition. Most other conversions inexact… and follow sig fig rules!

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Test yourself: How many hours in 3.5 weeks?

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Test yourself: How many quarters will a tourist need to travel 555 km? Car: 22 miles per gallon, gas: $1.37/gallon, 1.61 km = 1 mile

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Measuring Temperature

Temperature, the measure of how hot or cold an object is, is commonly reported either in Fahrenheit (^oF) or Celsius (^oC). The SI unit of temperature is, however, is the Kelvin (K).

Kelvin temperatures are ***always positive*** and they do not use the degree (^{o}) symbol.

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Freezing point of H_2O	Boiling point of H_2O
32^oF	212^oF
0^oC	100^oC

212 – 32 = 180 ^oF covers the same *range* of temperature as 100 - 0 = 100 ^oC.

Therefore, Celsius degree is exactly 180/100 = 1.8 times as large as Fahrenheit degree.

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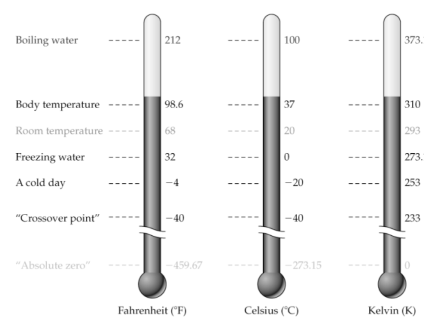

Comparison of the Fahrenheit, Celsius, and Kelvin Scales

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Converting between Fahrenheit and Celsius scales is similar to converting between different units of length or volume.

The following formulas can be used for the conversion:

$^oF = (9\ ^oF/ 5\ ^oC) \times\ ^oC + 32^oF$
$^oC = 5\ ^oC/ 9\ ^oF \times (^oF - 32^oF)$

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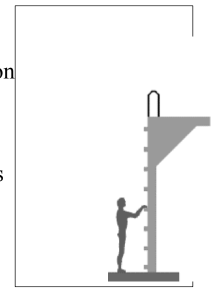
Energy and Heat

Energy is the capacity to do work or supply energy.

Two types of Energy:

1. ***Potential Energy***: stored energy
   Example: a swimmer standing on top of a diving platform



2. ***Kinetic Energy***: energy of motion
   Example: a swimmer who jumps from the top of the diving platform

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Potential Energy

In chemical reactions, the potential energy is often converted into heat.

Products have less potential energy than the reactants – the products are more *stable* than the reactants.

Stable products have very little potential energy and have little tendency to undergo further reaction.

SI unit of energy is the **Joule (J)** and the metric unit of energy is **calorie (cal)**.

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One Calorie (cal) is the amount of heat necessary to raise the temperature of 1 g of water by 1 ^oC.

One **food calorie** (**Cal**) = 1000 calories = 1 kcal

1000 cal = 1 kcal (kilocalorie)
1 food Calorie = 1 kcal = 1000 cal
1000 J = 1 kJ
1 cal = 4.184 J   1 kcal = 4.184 kJ

Not all substances have their temperature raised to the same extent when equal amounts of heat applied.

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The amount of heat needed to raise the temperature of 1 g of a substance by 1^{o}C is called the ***Specific Heat*** of the substance.

Unit of specific heat is cal/g ^{o}C

$$\text{Specific Heat} = \frac{\text{Heat (cal)}}{\text{Grams} * {}^{o}\text{C}}$$

Possible to calculate how much heat must be added or removed to accomplish a given temperature change of a given mass of a substance.

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<u>Example</u>: Calculate the amount of heat energy (in calories) required to heat 95 kg of water from 15 °C to 40 °C.

Specific heat of water = 1.0 cal/g °C.

Use: **Heat = mass*ΔT*(specific heat)**

where $\Delta T = T_{final} - T_{initial}$ = 40 - 15 = **25 °C**

and:

Heat = 95 kg * (1000 g/kg) * 25 °C * (1.0 cal/g °C)

Heat = 2.4 * 10^{6} cal

How many food calories (Cal)?

1 kcal = 1 Cal = 1000 cal

2.4 * 10^{6} cal * (1 Cal / 1000 cal) = 2.4 * 10^{3} Cal

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Density

Density relates the mass of an object with its volume.

Density is usually expressed in:

- Gram per cubic centimeter (g/cm^{3}) (solids)
- Gram per milliliter (g/mL) (liquids)

$$\text{Density} = \frac{\text{Mass (g)}}{\text{Volume (mL or cm}^3\text{)}}$$

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Test yourself: Mercury has a density of 13.6 g/mL. How many L of Hg are there in 42.7 kg of Hg?

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Specific Gravity

***Specific Gravity*:** density of a substance divided by the density of water at same temperature.

Specific Gravity is unitless

At normal temperature, the density of water is close to 1 g/mL. Thus, specific gravity of a substance at normal temperature is equal to the density.

$$\text{Specific gravity} = \frac{\text{Density of substance } (\cancel{\text{g/ml}})}{\text{Density of water at the same temperature } (\cancel{\text{g/ml}})}$$

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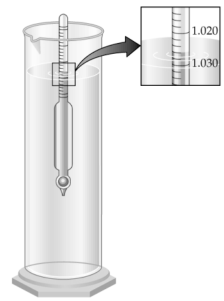
The specific gravity of a liquid can be measured using a hydrometer (a weighted bulb on the end of a calibrated glass tube)

The depth to which the hydrometer sinks indicates the fluid's specific gravity.



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End of Chapter 2

To review and study for Chapter 2, look at the "Concepts to Remember" at the end of Chapter Two