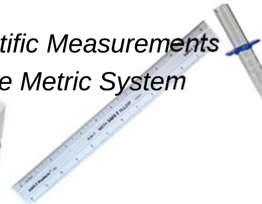


Chapter 2 / 3

Scientific Measurements

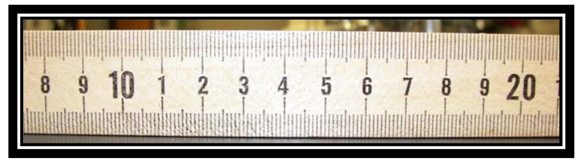
The Metric System


Taking Measurements of Real World Quantities

The precision of the measurement is limited by the measuring device.

You are allowed one digit of uncertainty in any measurement and no more.



Graduated Cylinder

Read to the bottom of the meniscus

This graduated cylinder is calibrated to the 1 mL mark. To what precision can measurements be reported?

"Tare" the balance to zero it.

The uncertainty digit is the last one given. What decimal place is this?

Analytical (mass) Balance

Computation using significant digits: (Lab manual p14)

1. Nonzero integers. Nonzero integers always count as significant figures.

The number 1457 has four nonzero integers, all of which count as significant figures.

2. Zeros. There are three classes of zeros:

Leading zeros are zeros that precede all of the nonzero digits. They never count as significant figures.

In the number 0.0025, the three zeros simple indicate the position of the decimal point. The number has only two significant figures, the 2 and the 5.

Captive zeros are zeros that fall between nonzero digits. They always count as significant figures.

The number 1.008 has four significant figures.

Trailing zeros are zeros at the right end of the number. They are significant only if the number contains a decimal point.

The number one hundred written as 100 has only one significant figure, but written as 100., it has three significant figures.


3. Exact numbers. Often calculations involve numbers that were not obtained using measured devices but were determined by counting: 10 experiments, 3 apples, 8 molecules. Such numbers are called exact numbers. They can be assumed to have an unlimited number of significant figures.

Exact numbers can also arise from definitions. For example, 1 inch is defined as exactly 2.54 centimeters. Thus in the statement 1 in. = 2.54 cm, neither 2.54 nor 1 limits the number of significant figures when it is used in a calculation.

Computations

Addition/Subtraction:

The result of an addition and/or subtraction may contain only as many decimal digits as the number with the fewest decimal digits in the computation.

Multiplication/Division:

The result of a multiplication and/or division may contain only as many total digits as the number with the fewest total significant digits in the computation.

When performing a compound calculation, try to avoid intermediate rounding as this tends to magnify the errors in your result (**Round only at the end but track the number of digits you will be allowed**).

Rounding:

Look at the number to the right of the last significant digit you will be reporting; if the number greater than 5, round your last digit up. If the number is less than 5, leave the last digit as is. If the first insignificant number is equal to 5 or is 5 followed by zeros; round the last significant digit up if odd, and leave it as is if even.

23.3500 rounds to 23.4 (to the tenths place)

23.4500 rounds to 23.4 (to the tenths place)

In the first case, the digit in the tenths place is odd, so it rounds up.

In the second case, the digit in the tenths place is even, so it stays the same.

Examples: Do the following calculations to the correct number of rounded significant digits. Include units.

1. 1.5 in + 34.465 in – 21.33 in
2. 4.55 cm * .043 cm * 9.233 cm
3. (23.655 cm – 19.2 cm) / 10.050 s

Answers

1. 1.5 in + 34.465 in – 21.33 in = 14.635 in = **14.6 in**
2. 4.55 cm * .043 cm * 9.233 cm = 1.806436…cm^3 = **1.8 cm^3**
3. (23.655 cm – 19.2 cm) / 10.050 s = 4.455 cm / 10.050 s = .4432835…cm/s = **0.44 cm/s**

Scientific Notation

- Allows for representation of very large or very small numbers conveniently.

- Allows the representation of a specific number of significant digits. (Ex. The number 200 represented to 2 significant digits is $2.0x10^2$)

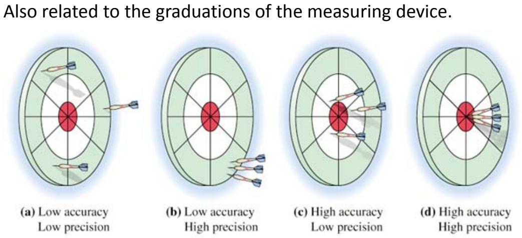
- Be sure to enter exponential (scientific) notation correctly in your calculator (use the exp or EE key. Do not use ^ if you can help it)
Ex. [5][.][3][EE][4] = $5.3x10^4$

Avoid [5][.][3][x][10][^][4]

Accuracy vs. Precision:

Accuracy: How well a measured value agrees with the accepted value.

Precision: How well measured values agree with one another. Also related to the graduations of the measuring device.



Random vs. Systematic Error

Dimensional Analysis (aka the factor-label method)

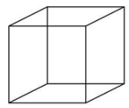
Converting units expresses the same quantity but in a different unit, so the value of the number changes (not the quantity it represents)

The conversion factors are "unit" conversions because they express the same magnitude of something.
1 in = 2.54 cm
Therefore (1 in/2.54 cm) = 1 or (2.54 cm / 1 in) = 1

When setting up conversion ratios, put the new unit in the position where you want it (i.e. numerator or denominator) then put the unit being converted into the opposite position
e.g. 3.557 cm (1 in / 2.54 cm)

Dimensional Analysis
(aka the factor-label method)

Watch converting volumes and areas that you ***convert in all dimensions***.



Example

53.5 in^3 (2.54 cm / 1 in)^3
53.5 in^3 (16.387064 cm^3 / 1 in^3)

Everything (including the units) gets raised to the power outside of the parenthesis.

1. Sample Calculations Using Significant Digits:

Solve: (6.223)(.015)(2.33+8.67)/(2.44-2.0) using proper significant digits.

2. Conversion Factors and Dimensional Analysis:

Ex. Convert a measurement of 6.114m/s into mi/hr using dimensional analysis and the fact that there are 2.54cm/in and 5280ft in a mi.

Answers:

1. (.093345)(11.00)/(.44) = 2.333625 = **2**

2. 6.114 m/s (3600s / 1hr)(100cm / 1m)
(1in / 2.54cm)(1ft / 12in)(1mi / 5280ft) = 13.6766… =
13.68mi/hr

1. Convert 5.70g/cm^3 to dg/in^3 to the correct number of significant digits given that 1 inch = 2.54 cm exactly.

Answer:

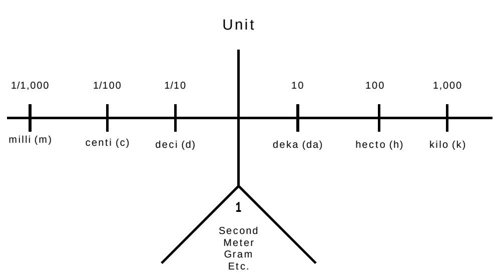
5.70g/cm^3 (10 dg / 1 g)(2.54 cm / 1 in)^3
= 934.0626… dg/in^3
= **9.34×10^2 dg/in^3**

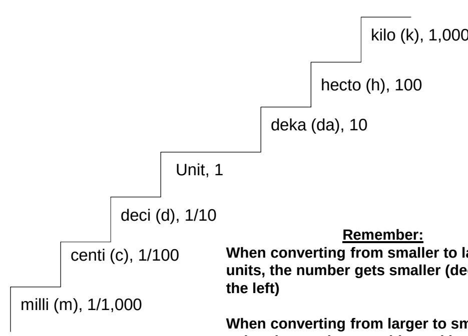
Metric System

Base 10 system of units.
Prefixes determine multiple of the base unit.

Base Units

	Unit	**Symbol**
Length	meters	m
Time	seconds	s
Mass	kilograms	kg
Quantity	moles	mol

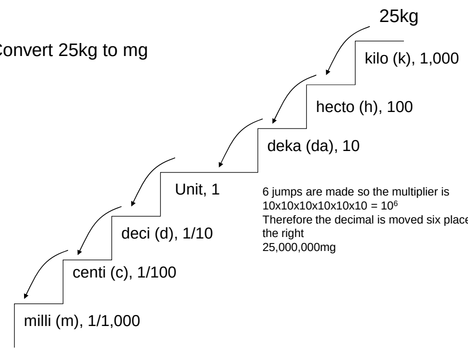




<u>Remember:</u>
When converting from smaller to larger units, the number gets smaller (decimal to the left)

When converting from larger to smaller units, the number gets bigger (decimal to the right)

Convert 25kg to mg



6 jumps are made so the multiplier is
$10x10x10x10x10x10 = 10^6$
Therefore the decimal is moved six places to the right
25,000,000mg

Other Important Prefixes You Should Know

	Multiple	Symbol
nano	10^{-9} (.000000001)	n
micro	10^{-6} (.000001)	µ (mu)
mega	10^{6} (1,000,000)	M

The Percent Concept

Expresses the amount of a single quantity compared to an entire sample.

It is given as parts out of 100

e.g. 25% = 25/100 = .25 (as a decimal)
40% = 40/100 = .40

Decimal Percent = Part / Whole
Percent = Decimal Percent x 100

e.g. Out of a collection of coins, 8 of them are pennies and the 16 are nickels. What is the decimal percent and percent of pennies in the collection

Answer: $\frac{8\ pennies}{24\ coins\ total} = .333 = 33\frac{1}{3}\% = 33.3\%$

Sample Problems

1. 22.0 is what percentage of 45.0?

2. 18.0 is 35.0% of what number?

3. 65.0% of what number is 16.0?

4. If shoes go on sale from \$75.00 to \$58.00, what is the percent decrease?

5. A rubidium (Rb) atom experiences a 39.9% decrease in radius when it loses an electron. If the original radius was 248.pm, what is the new "ionic" radius (in picometers)?

Sample Problems Answers

1. 22.0 is what percentage of 45.0? (22.0/45.0)x100 = 48.9%

2. 18.0 is 35.0% of what number? (18.0/.350) = 51.4

3. 65.0% of what number is 16.0? (16.0/.650) = 24.6

4. If shoes go on sale from \$75.00 to \$58.00, what is the percent decrease? (1 – 58.00/75.00) x100 = 22.7%
or (17.00/75.00)x100 = 22.7%

5. A rubidium (Rb) atom experiences a 39.9% decrease in radius when it loses an electron. If the original radius was 248.pm, what is the new "ionic" radius (in picometers)?
If it loses 39.9%, there is 60.1% left. (.601)(248.pm) = 149pm

Derived Units

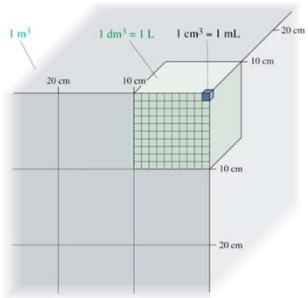
Derived units are units that are combinations of base units.
For instance, in physics the newton, N, is the unit of force.
It is derived from the base units kg m/s^2
So, 1N = 1kg m/s^2

Example: A joule is a derived unit of energy based on mass times length squared, divided by time squared. What are the base units of a joule?

Answer: 1 joule = kg m^2 / s^2

Volume

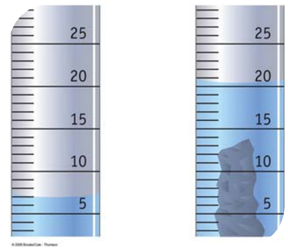
The liter and milliliter are also a derived units.
1 liter (L) = 1dm^3
1 milliliter (mL) = 1cm^3
There are 1,000 mL in one liter



Density

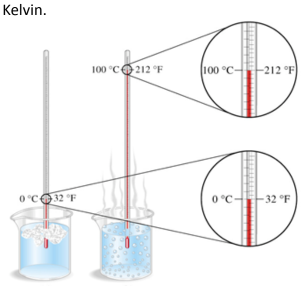
Density = mass / volume (D = m/V)

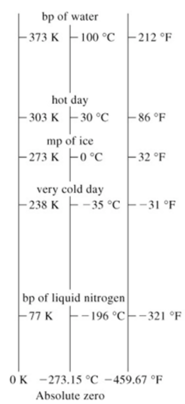
Determines whether an object will “sink or float” in another substance.
Water has a density of 1.00g/cm^3 (rounded)
Densities are sometimes determined by “water displacement”



The change in volume of the water when an object is submerged is the volume of the object and is called ΔV the density of the object is $D = m / \Delta V$
Mass is usually given in grams and volume in mL or cm^3

Temperature Conversions:
Units: Celsius, Fahrenheit and Kelvin
$^oF = 9/5\ ^oC + 32$ $\qquad$ $^oC = 5/9\ (^oF – 32)$
$K = ^oC + 273.15$
Remember there are no negative values for Kelvin.





Question: The “surface” temperature on the planet Neptune is133K, whereas on the planet Venus it is 737K. Convert these temperatures to Fahrenheit.

Answer:

$^oF_{Neptune}$ = [9(133-273.15)/5] + 32 = -220.27 = -220.

$^oF_{Venus}$ = [9(737-273.15)/5] + 32 = 866.93 = 867

Temperature Conversions:
Units: Celsius, Fahrenheit and Kelvin
$K = ^oC + 273.15$
$^oF = 9/5\ ^oC + 32$
$^oC = 5/9\ (^oF – 32)$

Heat and Specific Heat

Although heat and temperature are related, they are NOT the same thing

Heat: Energy transfer via random molecular motions resulting in gain or loss of internal energy.

Temperature: A measure of the average kinetic motion (energy) of a system of particles

When you compare a “hot” cup of coffee with a “warm” tub of water, which has greater heat energy? Which has greater temperature?

Specific heat (c) is the property of a substance that determines how much for a given amount of that substance the temperature will increase for an amount of heat energy either gained or lost.

$$q = m\ c\ \Delta T$$

Where q is the heat energy, m is the mass and ΔT is the temperature change.
For water c = 1 cal / g^{o}C or 4.184 J / g^{o}C

Specific Heat Example

Gold has a specific heat of 0.0305 cal / g^{o}C, if a 5.00g sample of gold at 22.0^{o}C absorbs 8.00 cal of energy, what will it’s final temperature be?

Answer: 8.00 cal = (5.00g)(0.0305 cal/g^{o}C)(ΔT)
ΔT = 52.5^{o}C
Final temperature = 22.0^{o}C + 52.5^{o}C = **74.5^{o}C**