Measurement

International System of Units (SI)

- revised metric system proposed in 1960
- widely used in science
- 7 base units

SI Base Units

Length	Meter	m
Mass	Kilogram	kg
Time	Second	s
Electrical current	Ampere	A
Temperature	Kelvin	K
Amount of Substance	Mole	mol
Luminous intensity	Candela	cd

SI Prefixes

Tera-	T	10^{12}	Deci-	d	10^{-1}
Giga-	G	10^{9}	Centi-	c	10^{-2}
Mega-	M	10^{6}	Milli-	m	10^{-3}
Kilo-	k	10^{3}	Micro-	μ	10^{-6}
			Nano-	n	10^{-9}
			Pico-	p	10^{-12}

Derived units in SI

measured in terms of one or more base units

volume

m x m x m = (m^3) = 1000 (dm^3)

$1 dm^3$ = 1 liter (L)

density

mass/volume = (kg/dm^3) = (g/cm^3) = (g/ml)

Density

The mass of a substance that occupies one unit of volume

$$\textbf{Density} = \frac{\text{mass}}{\text{volume}} = \frac{\text{kg}}{\text{dm}^3} = \frac{\text{g}}{\text{cm}^3} = \frac{\text{g}}{\text{ml}}$$

Example

What is the density of a piece of concrete that has a mass of 8.76 g and a volume of 3.07 cm³

$$\textbf{Density} = \frac{\text{mass}}{\text{volume}} = \frac{8.76\text{g}}{3.07\ \text{cm}^3} = 2.85\text{g/cm}^3$$

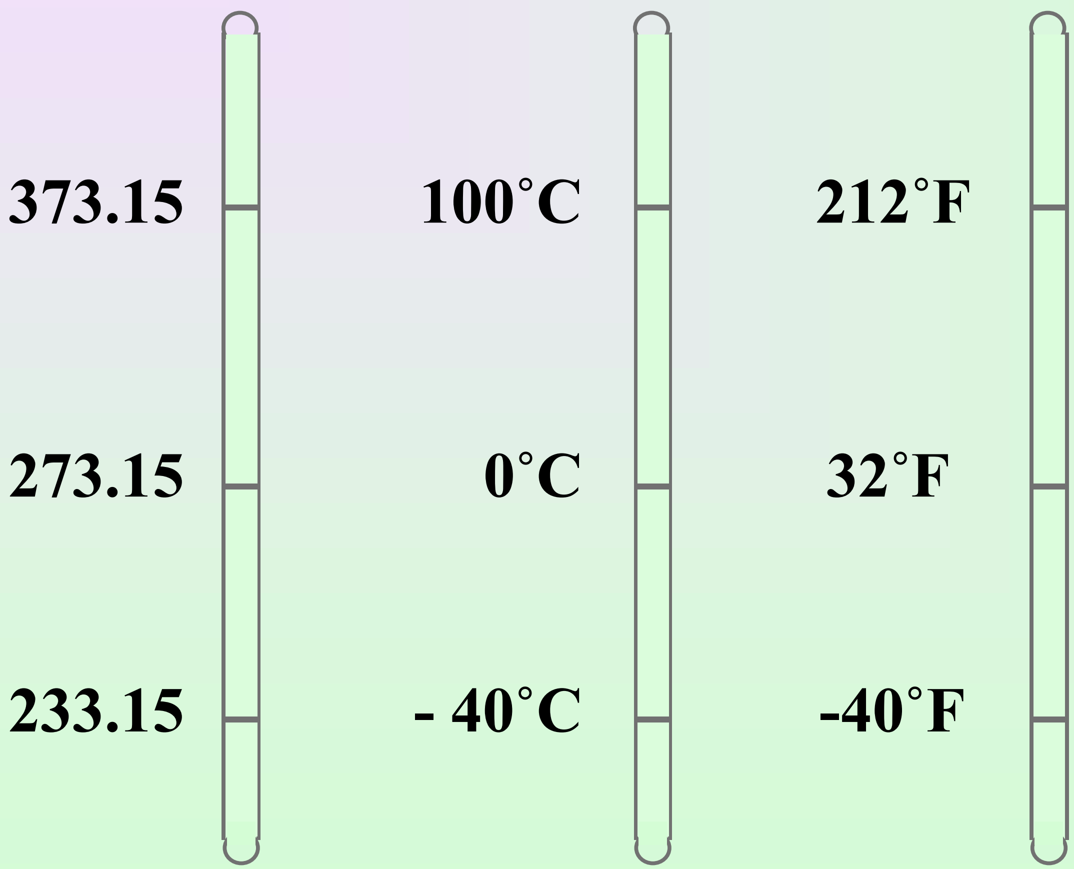
Temperature

There are three systems for measuring temperature that are widely used:

Kelvin scale	$K = C^\circ + 273.15$
Celsius scale	$C^\circ = K - 273.15$
Fahrenheit scale	$F^\circ = C^\circ (9/5) + 32$
Used mainly in engineering	$C^\circ = (F^\circ - 32)\ 5/9$

Temperature

Kelvin scale **Celsius scale** **Fahrenheit scale**



Handling Numbers

In chemistry we deal with very large and very small numbers

Scientific Notation

is a way of dealing with numbers that are either extremely large or extremely small

$N \times 10^n$

where **N** is a number between 1 and 10 and **n** is an exponent that can be a positive or negative integer

Exponents

$$100 = 10 \times 10 = 10^2$$

$$0.1 = \frac{1}{10} = 10^{-1}$$

$$0.001 = \frac{1}{10} \times \frac{1}{10} \times \frac{1}{10} = 10^{-3}$$

Example

Express 568.762 in scientific notation.

$568.762 = 5.68762 \times 10^2$

note that the decimal point moved to the left by two places and n = 2.

Example

Express 0.00000772 in scientific notation.

$0.00000772 = 7.72 \times 10^{-6}$

note that the decimal point moved to the right by six places and $n = -6$.

Scientific Notation

To add or subtract using scientific notation, first write each quantity with the same exponent n. Then add or subtract the N parts of the numbers; the exponent parts remain the same.

Example

$$(4.31 \times 10^4) + (3.9 \times 10^3) =$$

$$(4.31 \times 10^4) + (0.39 \times 10^4)$$

$$= 4.70 \times 10^4$$

Scientific Notation

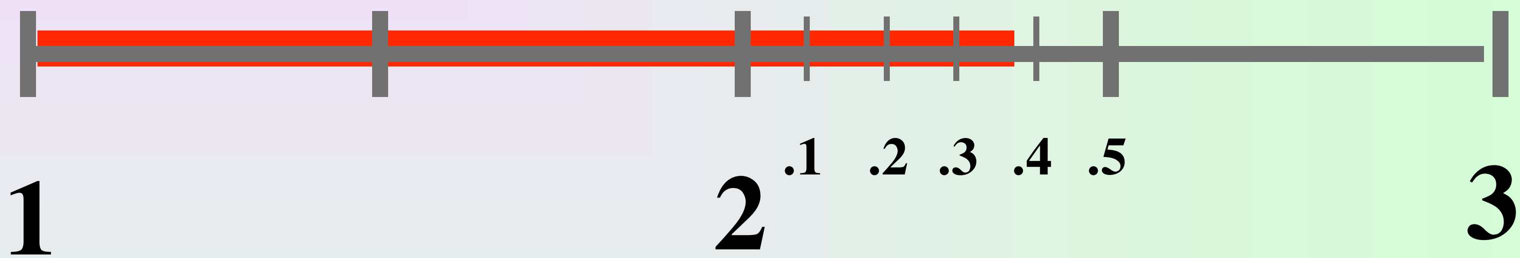
To multiply numbers expressed in scientific notation, multiply the N parts of the numbers in the usual way, but add the exponent n's together.

Example

$(8.0 \times 10^4) \times (5.0 \times 10^2) = 40.0 \times 10^6$

$= 4.0 \times 10^7$

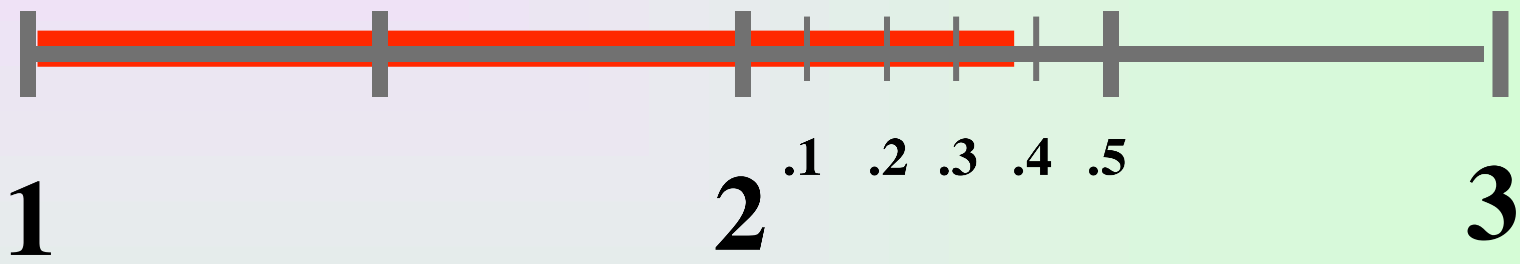
Uncertainty in measurement



2.36mm 2.37mm

middle value ?

Uncertainty in measurement



2.36mm **2.37mm**

middle value ?

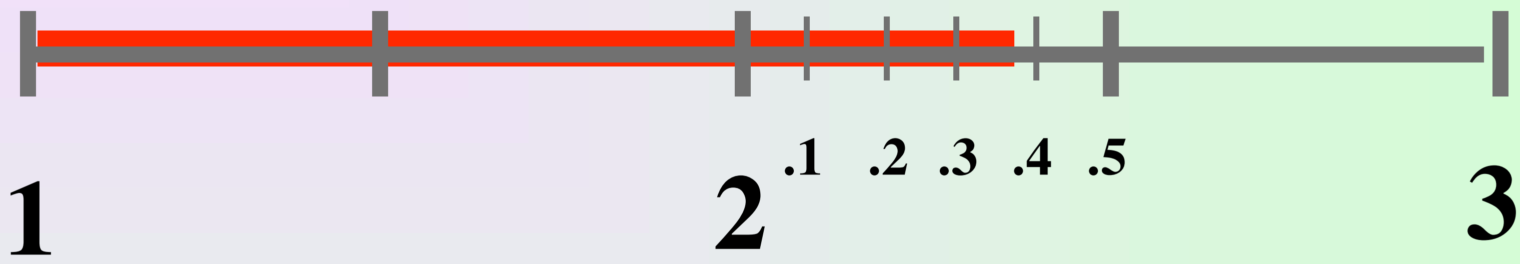
There is uncertainty with this degree of accuracy

Accuracy - the closeness of a measurement to the true value

precision - the reproducibility of a series of measurements

A series of measurements can be precise with out being accurate

Uncertainty in measurement



2.37mm

The first two measured numbers are called *certain* digits

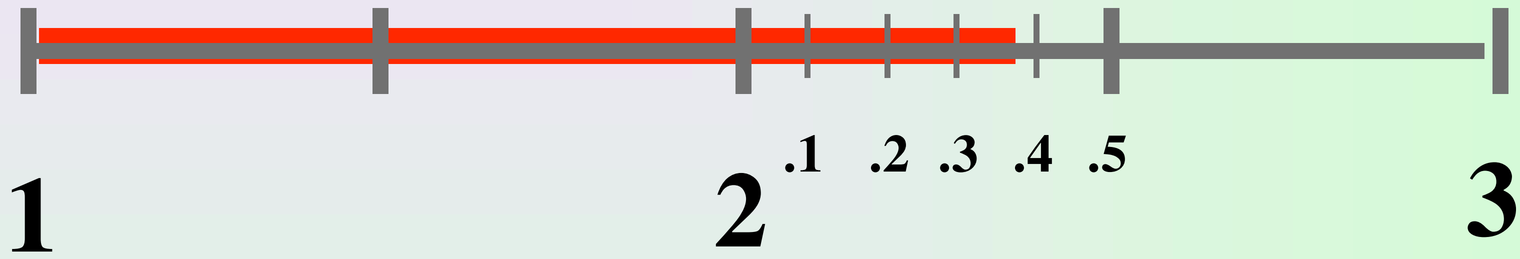
The the digit to the right of the 3 is called an *uncertain* digit

a measurement always has some degree of *uncertainty*

We customarily report a measurement by recording all the certain digits plus the first uncertain digit.

these numbers are called significant figures

Significant Figures



2.37mm

three significant figures in this measurement

Example

What is the difference between the measurements 25.00 ml and 25. ml .

They convey different information:

25.00 ml   a volume between 24.99ml and 25.01 ml

25 ml   a volume between 24 ml and 26ml

Rules for Significant Figures

digits other than zero are always significant

67.8 g **3 significant figures**

98. g **2 significant figures**

one or more final zeros used after the decimal point are always significant

4.700 m **4 significant figures**

82.0 m **3 significant figures**

Rules for Significant Figures

zeros between two other significant digits are always significant

5.029 cm **4 significant figures**

zeros used solely for spacing the decimal point are not significant

0.00783 ml **3 significant figures**

0.34 g/ml **2 significant figures**

Rules for Significant Figures

If the zeros follow nonzero digits,there is ambiguity if no decimal point is given

300 N	**significant figures ?**
300. N	**3 significant figures**
300.0 N	**4 significant figures**

Avoid ambiguity by expressing measurements in scientific notation

3.0×10^2 N	**2 significant figures**

Adding Significant Figures in Calculations

A result can only be as accurate as the least significant measurement

4.37 g

+ 1.002 g

5.37~~2~~ g **3 significant figures**

Multiplying Significant Figures in Calculations

A result can only be as accurate as the least significant measurement

Volume = l x w x h = (1.87cm)(1.413cm)(1.207cm)

= 3.19cm

3 significant figures

Using Significant Figures in Calculations

A result can only be as accurate as the least significant measurement

21.mm - 13.8mm = 7mm

1 significant figures

21 has no decimal place

Significant Figure Rules

Plus or minus (±) notation may be used to express the amount of uncertainty there is in a measurement

51.56 cm^3 ± 0.02 cm^3

51.58 cm^3 largest possible value

51.54 cm^3 smallest possible value

Rounding Off Rules

In a series of calculations, carry the extra digits through the final result, then round.

If the digit following the last reportable digit is:

- **4 or less, you drop it** — **1.33 to 1.3**

- **5 or more, you increase the last reportable digit by one** — **1.36 to 1.4**

Dealing with Propagation of Error

If the digit following the last reportable digit is:

- **4 or less, you drop it**
- **6 or more, you increase the last reportable digit by one**
- **5, you use the arbitrary odd-even rule**
  - **If the last reportable digit is even, you leave it unchanged**
  - **If the last reportable digit is odd, you increase it by one.**

Example

Round off 108.75 and 108.65 to four significant figures.

108.75 = 108.8

odd

108.65 = 108.6

even

Percent Error

Observed value

the value based on laboratory measurements

True value

the value based on accepted references

Absolute error

the difference between the observed value and the true value

(observed value - true value)

Percent Error

$$\% \text{ Error} = \frac{\text{absolute error}}{\text{true value}} \times 100\%$$

Example

the boiling point of methanol is 65˚C. Your measured boiling point of methanol is 66.0˚C.what is the percent error in your measurement.

$$\% \text{ Error} = \frac{\text{absolute error}}{\text{true value}} \times 100\%$$

$$\% \text{ Error} = \frac{66^\circ C - 65^\circ C}{65^\circ C} \times 100\%$$

$$= 1.\ 5\%$$

The Unit-Factor Method of Solving Problems

also called "dimensional analysis"

it is a good idea to carry units in a calculation to ensure that the answer to the problem has the correct units

The Unit-Factor Method

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

dividing both sides of the equation by 2.54cm

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

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

we create an expression called a unit-factor

Example

What is the length of a 2.85cm pin in inches?

$$2.85\ \text{cm} \times \frac{1\text{in}}{2.54\text{cm}} = 1.12\ \text{in}$$

Example

Where were you a billion seconds ago ?

$$1\times10^{9}\ \text{sec} \times \frac{1\ \text{min}}{60\ \text{sec}} \times \frac{1\ \text{hour}}{60\ \text{min}} \times \frac{1\ \text{day}}{24\ \text{hours}}$$

$$\times \frac{1\ \text{year}}{365\ \text{days}} = 31.7\ \text{years}$$