Measurements

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Fundamental Scientific Measurements

	SI unit	abbreviation
Length	Meter	m
Mass (weight)	Kilogram	kg
Temperature	Kelvin	K
Time	Second	s
Amount	Mole	mol
Light intensity	Candela	cd
Electric current	Ampere	A

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Why Use the Metric System?

- Scientist use the metric system- easier to do conversions, communicate with others
- Case is very important- pay attention to detail

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Quality of Measurements Accuracy/Precision/Significant figures

Accuracy- how close a measurement is to the "true value"

Precision- how reproducible is the measurement

Significant Figures- The number of significant figures is based on the measurements taken

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Example of Making Measurements and Significant Figures

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Rules That Govern Significant Figures

1. All non-zero digits are significant.
2. Zeros between nonzero digits are significant.
3. Zeros to the left of the first nonzero digit are not significant.
4. Zeros at the end of a number that includes a decimal point are significant.

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Calculations with Significant Figures

1. In addition and subtraction an answer should have no more decimal places than the number with the fewest decimal places.

   Rounding off.

   Exact numbers-numbers which are not measurements and therefore not estimated.
2. For multiplication and division the answer is limited to the number of digits in the number with the fewest significant figures.

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Sample Test Question

- How many significant figures should the following calculation have?

  $25.0 \times 3.0 \times 4.88$

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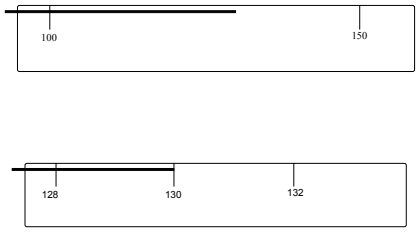
Why the End Zero Isn't Significant (without a decimal point)



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Conversions

Dimensional analysis

AKA unit cancellation method

1. Where are you?
2. Where do you want to go?
3. What conversion factors do you need to get there?

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Scientific Notation

- Very large numbers (Avogadro's number)
  602,200,000,000,000,000,000,000 atoms/mol
- Very small numbers (mass of an electron)
  0.000,000,000,000,000,000,000,000,000,9110 g
- Number between 1 and 9.99
- $\times 10^{n}$ where n is the number of decimal places you had to move to get to the number
- 6.022×10^{23} atoms/mol
- and 9.110×10^{-28} g

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Units of Measurement: Length

Fundamental Measurement

Length-

SI units of meter (m)

Sample question: If you are 5325 m away from an object, and walk 32.2 m closer to it, how far away will you be?

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Prefixes Most Often Used (Table 1.4)

– mega-	M	10^6	million
– kilo-	k	10^3	thousand
– centi-	c	10^{-2}	one-hundreth
– milli-	m	10^{-3}	one-thousandth
– micro-	μ	10^{-6}	one-millionth

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Units of Measurement: Volume

Derived Measurement

Volume- the amount of space matter occupies

SI unit is m^3 which you will never use

cc = cm^3 = mL

L = dm^3

Sample question: What is the volume of an object which is 2.3 cm x 6.3 cm x 25.0 cm?

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Units of Measurement: Mass

Fundamental Measurement

Mass- (weight) the amount of matter in an object

SI units of kg

use g much more often

Sample question: If you measure out 25.3 g of water, and divide it into 6 equal parts, what is the mass of each part?

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Units of Measurement: Density

Derived Measurement

Density- the amount of mass in a given amount of volume

units will vary depending on state:

g/L for gas

g/mL for liquid

g/cm^3 for solid

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Sample Test Questions

1. In a laboratory maintained at 20°C, you determine the mass of an apparent meteorite to be 56.981 g. The volume in triplicate determinations was found to be 19.03 cm^3. What is the density of the meteorite?
2. Given that mercury (Hg) has a density of 13.5347 g/mL, what is the mass of 50.0 mL of Hg?
3. What volume of lead is required to obtain 525.326 g lead if it has a density of 11.34 g/cm^3?

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Units of Measurement: Energy

Energy- the ability to do work

Kinetic Energy- energy from motion (vibration, etc)

Potential Energy- stored energy (bonds)

Heat- energy traveling from a warmer to a cooler object.

SI unit is joule (J)

1 cal = 4.18 J; 1 Cal = 1000 cal

Sample question: How many joules are in a 228 Cal Snickers bar?

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Units of Measurement: Temperature

Temperature- a measurement of the average kinetic energy in a sample (how fast molecules are vibrating)

SI unit is the Kelvin (K)

$^oC + 273 = K$

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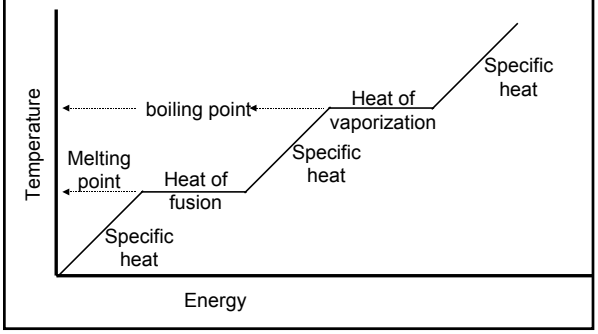
The Temperature-Energy Diagram and Changes in Physical State



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Sample Test Questions

1. How many calories of thermal energy are required to convert 75.6 g ice at -20°C to -5°C? The specific heat of ice is 1.00 cal/g°C
2. How many calories of thermal energy are required to convert 75.6 g ice at -0°C to water at 0°C? The heat of fusion for H_2O is 79.6 cal/g
3. How many calories of thermal energy are required to convert 63.25 g ice at -20°C to steam at 120°C? The heat of vaporization of water is 540 cal/g.

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Practice Conversion Problems

1. Over 40. billion kilograms of sulfuric acid are produced in the US each year. How many pounds per day is this?
2. What is the mass, in kilograms, of a 65-pound chimpanzee?