Science and Measurements

Chapter 1

Chapter 1 Instructional Goals

1. Explain, compare, and contrast the terms **scientific method**, **hypothesis**, and **experiment.**
2. Compare and contrast scientific **theory** and **scientific law**.
3. Define the terms **matter** and **energy**. Describe the three phases (states) of matter and the two forms of energy.
4. Describe and give examples of **physical properties** and **physical change.**
5. Perform unit conversion calculations.
6. Express and interpret numbers in scientific (exponential) notation.
7. Explain the difference between the terms **accurate** and **precise**.
8. Know and use the rules for **significant figures**.
   - Given a value, determine the number of significant figures.
   - Use the correct number of significant figures to report the results of calculations involving measured quantities.

What is Science?

Science is a ***method*** for gaining knowledge and understanding of reality.

It produces **generalizations** with ***predictive*** value.

There are two ways to do science:
scientific *theory* and **scientific *law***.

It is important to note that *both* methods are used to acquire predictive power and both begin with ***observation(s)***.

Scientific Theory

Other words for **theory** are ***model*** or ***explanation***.

Scientific theory uses models/explanations to make sense of observables.

- Often, a first guess at a model is proposed.
- The first guess is called a ***hypothesis***.

The hypothesis can usually be tested by experiment or additional observations.

If the hypothesis continues to be validated by experiment or new observations, it becomes ***theory***.

In the healthcare field, another word for theory or model is ***diagnosis***.

Scientific Law

A scientific law is simply ***a statement about something that generally occurs***.

Note that in using scientific law, ***no*** explanation (model) is given.

Scientific law can be contrasted with scientific theory that involves proposing a model or explanation for what is observed.

Chemistry

Chemistry is the study of matter and how it interacts with other ***matter*** and/or ***energy***.

Matter

Matter is anything that has mass and occupies space.

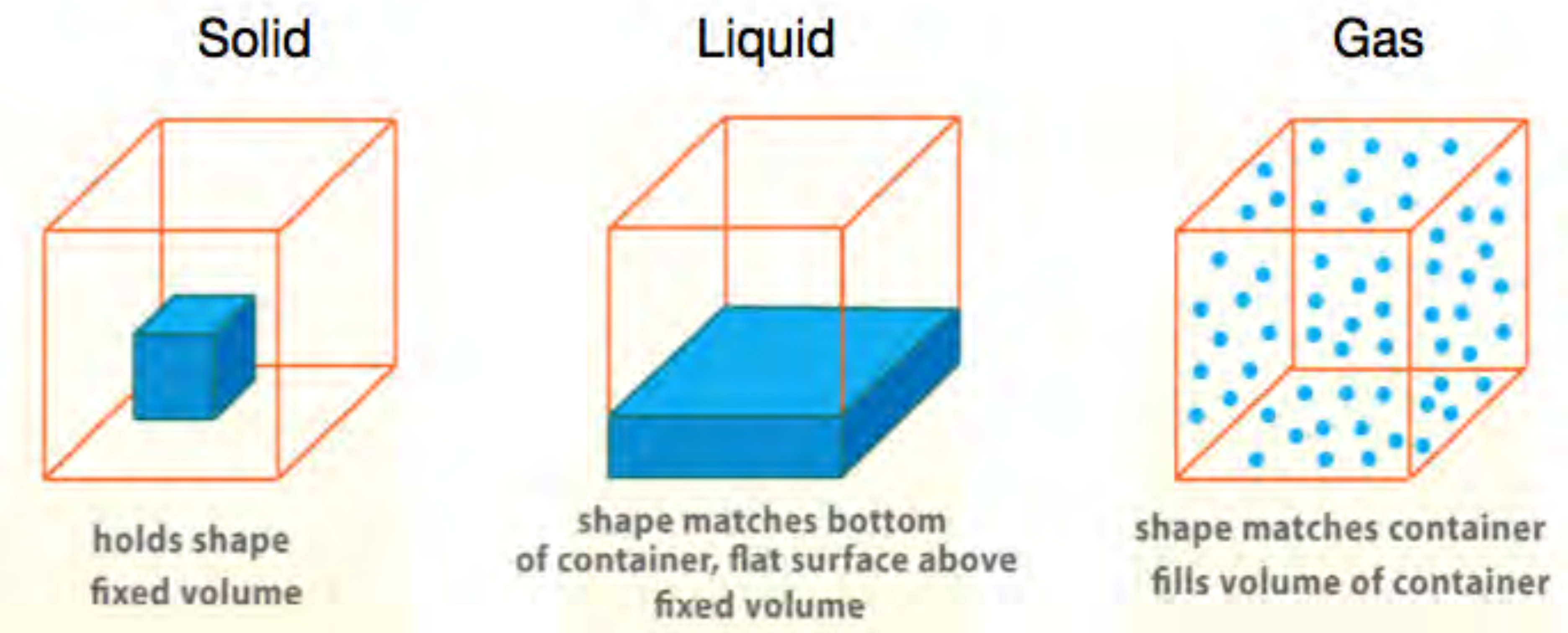
We can describe matter in terms of ***physical properties***, those characteristics that can be determined without changing it into a different substance.

- Example: Sugar is white, tastes sweet, and can be crushed into powder. Crushing sugar does not change sugar into something else.

- Matter can also be described in terms of its ***chemical properties***. Chemical properties of substances describe *how they are converted to new substances* in processes called chemical reactions.
  - Example: Carmalization of sugar

Matter

Matter is typically found in one of three different physical ***phases*** (sometimes called ***states***).



Example: Ice **Example: water** **Example: steam**

Matter

Changing the phase of matter, converting matter between solid, liquid, and gas is considered a *physical change* because the identity **does not** change.

- Examples of phase changes are: melting, boiling water to make steam, and melting an iron rod.

Energy

Energy is commonly defined as the ability to do ***work***.

Energy can be found in two forms, potential ***energy*** and ***kinetic*** *energy*.

Potential energy is ***stored*** energy; it has the potential to do work.

- An example of potential energy is water stored in a dam. If a valve is opened, the water will flow downhill and turn a paddle connected to a generator to create electricity.

Energy

Kinetic energy is the energy of ***motion***.

Any time matter is moving, it has kinetic energy.

An important law that is central to understanding nature is: **matter will exist in the lowest possible energy state**. Another way to say this is “if matter can lose energy, it will always do so.”

Understanding Check

Which are mainly examples of ***potential energy*** and which are mainly examples of ***kinetic energy***?

a) A mountain climber sits at the top of a peak.
b) A mountain climber rappels down a cliff.
c) A hamburger sits on a plate.
d) A nurse inflates a blood pressure cuff.

Units of Measurement

Units of Measurements

Measurements consist of two parts – a **number** and a **unit**.

SI Units and Their Symbols

Quantity	SI Unit Name	Symbol
Length	meter	m
Mass	kilogram	kg
Time	second	s
Temperature	Kelvin	K

Commonly Used Units and Their Symbols

Quantity	Unit Name	Symbol
Length	foot inch	ft in
Mass	gram pound	g lb
Volume	Liter	L
Temperature	Fahrenheit Celsius (or Centigrade)	^{0}F ^{0}C

Scientific Notation and Metric Prefixes

Scientific Notation

When making measurements, particularly in science and in the health sciences, there are many times when you must deal with very large or very small numbers.

Example: a typical red blood cell has a diameter of about 0.0000075 m.

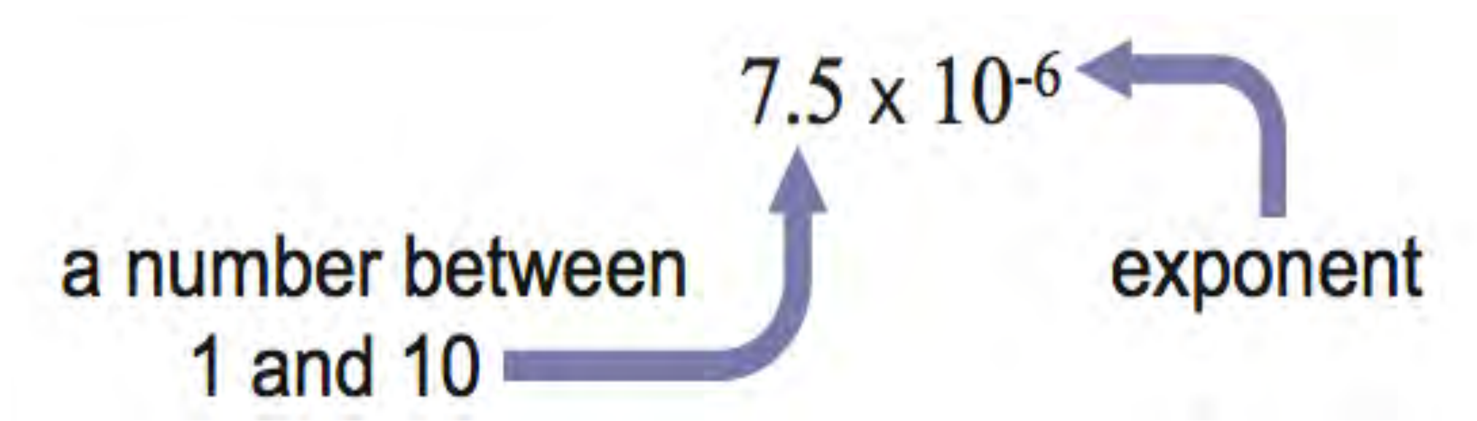
In ***scientific notation*** (exponential notation) this diameter is written 7.5×10^{-6} m.

$$\mathbf{0.0000075 = 7.5 \times 10^{-6}}$$

Scientific Notation

Values expressed in scientific notation are written as a number between **1** and **10** multiplied by a power of 10.

The superscripted number to the right of the ten is called an *exponent*.



- An exponent with a positive value tells you how many times to multiply a number by 10.

$$3.5 \times 10^4 = 3.5 \times 10 \times 10 \times 10 \times 10 = 35000$$

- An exponent with a negative value tells you how many times to divide a number by 10.

$$3.5 \times 10^{-4} = \frac{3.5}{10 \times 10 \times 10 \times 10} = 0.00035$$

Converting from Regular Notation to Scientific

1) Move the decimal point to the right of the first (right-most) non-zero number

- The exponent will be equal to the number of decimal places moved.

$$35000 = 3.5 \times 10^4$$

$$285.2 = 2.852 \times 10^2$$

$$8300000 = 8.3 \times 10^6$$

2) When you move the decimal point to the left, the exponent is positive.

Converting from Regular Notation to Scientific

$$0.00035 = 3.5 \times 10^{-4}$$

$$0.0445 = 4.45 \times 10^{-2}$$

$$0.000000003554 = 3.554 \times 10^{-8}$$

3) When you move the decimal point to the right, the exponent is negative.

Understanding Check

Convert each number into scientific notation.

a) 0.0144

b) 144

c) 36.32

d) 0.0000098

Converting from Scientific Notation to Regular Notation

You just learned how to convert from regular numerical notation to scientific notation. Now let's do the reverse; convert from scientific notation to regular notation.

Step 1: Note the value of the ***exponent***.

Step 2: The value of the exponent will tell you which direction ***and*** how many places to move the decimal point.

- If the value of the exponent is **positive**, remove the power of ten and move the decimal point that value of places to the ***right***.
- If the value of the exponent is **negative**, remove the power of ten and move the decimal point that value of places to the ***left***.

Example: Convert 3.7×10^5 into regular notation.

Step 1: Note the value of the ***exponent***: The exponent is **positive 5**.

Step 2: The value of the exponent will tell you which direction <u>and</u> how many places to move the decimal point. If the value of the exponent is **positive**, remove the power of ten and move the decimal point that value of places to the ***right***.

We will move the decimal point 5 places to the ***right***.

When the decimal point is **not shown** in a number, as in our answer, it is assumed to be ***after the right-most digit***.

Let's do another example: Convert 1.604 x 10^{-3} into regular notation.

Step 1: Note the value of the ***exponent***: The exponent is ***negative* 3**.

Step 2: The value of the exponent will tell you which direction <u>and</u> how many places to move the decimal point.

If the value of the exponent is **negative**, remove the power of ten and move the decimal point that value of places to the ***left***.

We will move the decimal point 3 places to the ***left***.

or 0.001604

Understanding Check

Convert the following numbers into regular notation.

a) 5.2789×10^{2}
b) 1.78538×10^{-3}
c) 2.34×10^{6}
d) 9.583×10^{-5}

Measurements and Significant Figures

There are three important factors to consider when making measurements:

1) **accuracy**
2) **precision**
3) **significant figures**

Accuracy is related to how close a measured value is to a true value.

Example: Suppose that a patient's temperature is taken twice and values of 98 °F and 102 °F are obtained. If the patient's true temperature is 103 °F, the second measurement is more *accurate*.

Precision is a measure of reproducibility.

Example: Suppose that a patient's temperature is taken three times and values of 98 °F, 99 °F and 97 °F are obtained. Another set of temperature measurements gives 90 °F, 100 °F and 96 °F.

The values in the first set of measurements are closer to one another, so they are more precise than the second set.

The quality of the equipment used to make a measurement is one factor in obtaining accurate and precise results.

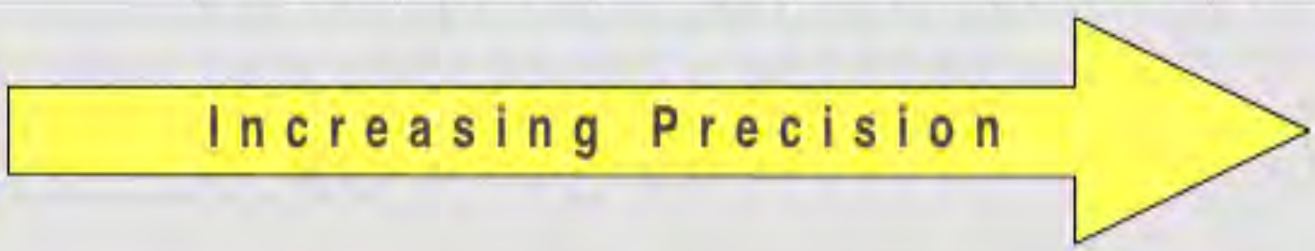
Significant Figures

One way to include information on the ***precision*** of a measured value (or a value that is calculated using measured values) is to report the value using the correct number of **significant figures**.

The ***precision of a measured value*** can be determined by the ***right****-most decimal place* reported.

- The names and precision of the decimal places **for the number 869.257** are shown below:

Digit in Number	8	6	9	2	5	7
Decimal Place Name	Hundreds (100's)	Tens (10's)	Ones (1's)	Tenths (1/10th's)	Hundredths (1/100th's)	Thousandths (1/1000th's)



A simple way to understand **significant figures** is to say that a digit is significant if we are ***sure*** of its value.

Method for Counting Significant Figures

We can look at a numerical value and determine the number of significant figures as follows:

- If the decimal point is ***present***, starting from the ***left***, count all numbers (including zeros) beginning with the first non zero number.

- If the decimal point is ***absent***, starting from the ***right***, count all numbers (including zeros) beginning with the first non zero number.

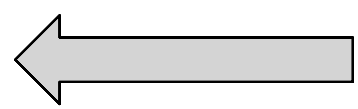
- When numbers are given in scientific notation, **do not** consider the power of **10**, only the value before “ $\mathbf{x\ 10^n}$.”

Example: If the botanist reported the age of the tree as **500 years**, how many significant figures are given?

Note that although the decimal point is implied to be after the right-most zero, it is **absent** (not shown explicitly), therefore we use the decimal point **absent** rule shown above; if the decimal point is **absent**, starting from the ***right***, count all numbers (including zeros) beginning with the first non zero number.

We will start inspecting each digit from right (to left) as shown by the arrow.
We will start counting when we get to the first non zero number.

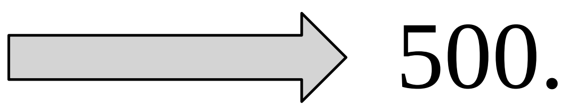
500



We do not count the first two zeros, but start counting at the **5**.
Therefore, there is **one** significant figure present.

Example: If the botanist reported the age of the tree as **500.** years (note the decimal point present), how many significant figures are given?

Note that in this case, the decimal point is **present** (shown), therefore we use the decimal point **present** rule shown above; if the decimal point is **present**, starting from the ***left***, count all numbers (including zeros) beginning with the first non zero number. We will start inspecting each digit from left to right as shown by the arrow. We will start counting when we get to the first non zero number.



We begin with the **5**, then count <u>all</u> numbers ***including zeros***. In this case, the two zeros are also significant. Therefore there are **three** significant figures present.

Example: If the botanist reported the age of the tree as **500.** years (note the decimal point present), how many significant figures are given?

Note that in this case, the decimal point is **present** (shown), therefore we use the decimal point **present** rule shown above; if the decimal point is **present**, starting from the ***left***, count all numbers (including zeros) beginning with the first non zero number. We will start inspecting each digit from left to right as shown by the arrow. We will start counting when we get to the first non zero number.

500.

We begin with the **5**, then count all numbers ***including zeros***. In this case, the two zeros are also significant. Therefore there are **three** significant figures present.

Outside of the science fields, “**500**” and “**500.**” are generally thought of as equivalent, however, the use of significant figures tells us that when we write “**500.**” (with the decimal point present) we know that number one hundred times more precisely than when we write “**500**” (without the decimal point). We have precision to the “ones” decimal place in “**500.**” vs. precision to the “hundreds” place in “**500**”.

Example: How many significant figures are contained in 0.00045?

Note that in this case, the decimal point is **present** (shown). We will start inspecting each digit from left to right as shown by the arrow. We will start counting when we get to the first non zero number.

0.00045

We begin with the **4**, then count all numbers ***including zeros***. Therefore there are **two** significant figures present.

Example: How many significant figures are contained in 0.0002600?

Note that in this case, the decimal point is **present** (shown). We will start inspecting each digit from left to right as shown by the arrow. We will start counting when we get to the first non zero number.

0.0002600

We begin with the **2**, then count all numbers ***including zeros***. Therefore there are **four** significant figures present.

Example: How many significant figures are contained in 7080?

If the decimal point is **absent**, starting from the ***right***, count all numbers (including zeros) beginning with the first non zero number. We will start inspecting each digit from right (to left) as shown by the arrow. We will start counting when we get to the first non zero number.

7080 

We do not count the first zero, but start counting at the **8,** and then count ***all*** *numbers* ***(including zeros)***. Therefore, there are **three** significant figures present.

Understanding Check:

Specify the number of significant figures in each of the values below.

a) 23.5

b) 0.0073000

c) 6.70

d) 48.50

e) 6200

f) 6200.

g) 6200.0

h) 0.6200

i) 0.62

j) 930

Significant Figures in Scientific Notation

When numbers are given in scientific notation, **do not** consider the power of 10, only the value before "**$x\ 10^n$**."

Examples: How many significant figures are contained in each of the values shown below?

a) 5×10^2 **one** significant figure

b) 5.0×10^2 **two** significant figures

c) 5.00×10^2 **three** significant figures

When converting back and forth from standard numerical notation to scientific notation, the number of significant figures used **should not change**.

Understanding Check

Write each measured value in ***scientific notation***, being sure to use *the correct number of significant figures*.

a) 5047
b) 87629.0
c) 0.00008
d) 0.07460

Calculations Involving Significant Figures

- When doing **multiplication or division** with measured values, the answer should have *the same number of significant figures* as the measured value with the least number of significant figures.

- When doing **addition or subtraction** with measured values, the answer should have the same *precision* as the least precise measurement (value) used in the calculation.

Example for Multiplication or Division:

- When doing **multiplication or division** with measured values, the answer should have *the same number of significant figures* as the measured value with the least number of significant figures.
  - If an object has a mass of 5.324 grams and a volume of 7.9 ml, what is it's density?

$$\text{density} = \frac{\text{mass}}{\text{volume}} = \frac{5.324\ \text{g}}{7.9\ \text{ml}} = 0.67\ \text{g/ml}$$

4 sig figs (5.324 g)

2 sig figs (7.9 ml)

2 sig figs (0.67 g/ml)

Example for Addition or Subtraction:

- When doing **<u>addition or subtraction</u>** with measured values, the answer should have the same *precision* as the least precise measurement (number) used in the calculation.
  - A book 50.85 mm thick, a box 168.3 mm thick and a piece of paper 0.037 mm thick are stacked on top of each other. What is the height of the stack?

	50.85 mm
	168.3 mm
+	0.037 mm
	219.187 mm

Least precise: precise to tenths

Round to tenths → 219.2 mm

Understanding Check

Each of the numbers below is measured. Solve the calculations and give the correct number of significant figures.

a) 0.12 x 1.77

b) 690.4 ÷ 12

c) 5.444 – 0.44

d) 16.5 + 0.114 + 3.55

Unit Conversions

Conversion Factors and the Factor Label Method

Typical Unit Conversion Problems

- A package weighs 3.50 kg (kilograms), what is the weight in lbs. (pounds)
- A student is 60.0 inches tall, what is the student's height in cm?
- The temperature in Cabo San Lucas, Mexico is 30.^{o}C, what is the temperature in ^{o}F?

To convert from one unit to another, we must know the ***relationship*** between the two units of measure.

Examples:

- A package weighs 3.50 kg (kilograms), what is the weight in lbs. (pounds)
  - 1kg = 2.20 lb

- A student is 60.0 inches tall, what is the students height in cm?
  - 1 inch = 2.54 cm

The *relationships* between units are called ***equivalence statements***.

Unit Relationships to Know:

- 1 milliliter (mL) = 1 cubic centimeter (cm^3)
- 1 inch (in) = 2.54 centimeters (cm)
- 1 kilogram (kg) = 2.20 pounds (lb)
- 4.184 Joule (J) = 1 calorie (cal)

The *relationships* between units are called ***equivalence statements***.

Unit Conversion Calculations: The Factor Label Method

A package weighs 3.50 kg (kilograms), what is the weight in lbs. (pounds)?

Equivalence statement: 1kg = 2.20 lb

$$\left(\frac{1\ \text{kg}}{2.20\ \text{lb}}\right) \quad \Longleftrightarrow \quad \left(\frac{2.20\ \text{lb}}{1\ \text{kg}}\right)$$

Conversion Factors

Equivalence statements can be written as ***conversion factors***.

$$3.50\ \cancel{\text{kg}} \left(\frac{2.20\ \text{lb}}{1\ \cancel{\text{kg}}}\right) = 7.70\ \text{lb}$$

$$\frac{3.50\ \cancel{\text{kg}}}{} \Bigg| \frac{2.20\ \text{lb}}{1\ \cancel{\text{kg}}} \Bigg| = 7.70\ \text{lb}$$

Some conversion factors have an ***<u>infinite</u>*** number of significant figures.

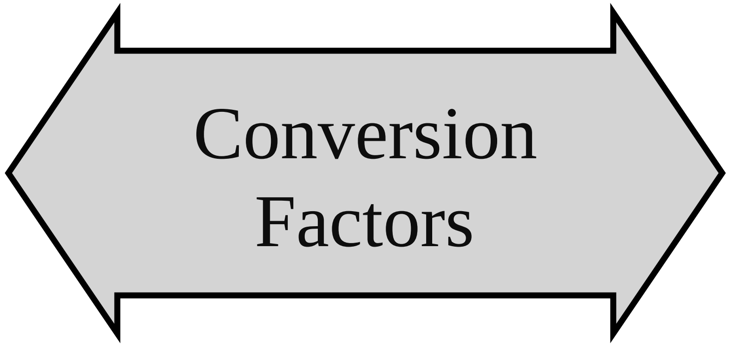
<u>Exact</u> (**defined or agreed upon**) conversion factors have an ***infinite*** number of significant figures.

- Examples of exact/defined conversion factors
  - 1 lb = 0.45359237 kg
  - 1 inch = 2.54 cm
  - 1 cg = 10^{-2}g
  - 1 ft = 12 inches
  - 1 ml = $1cm^3$

A student is 60.0 inches tall, what is the student’s height in cm?

Equivalence statement: 1 inch = 2.54 cm

$$\left(\frac{1\ \text{in}}{2.54\ \text{cm}}\right)$$



$$\left(\frac{2.54\ \text{cm}}{1\ \text{in}}\right)$$

$$\underbrace{60.0}_{\text{three significant figures}}\ \cancel{\text{in}}\ \underbrace{\left(\frac{2.54\ \text{cm}}{1\ \cancel{\text{in}}}\right)}_{\text{an infinite number of significant figures}} = 152\ \text{cm}$$

three significant figures

an **infinite** number of significant figures

$$\begin{array}{c|c|}
60.0\ \cancel{\text{in}} & 2.54\ \text{cm} \\
\hline
 & 1\ \cancel{\text{in}}
\end{array} = 152\ \text{cm}$$

Understanding Check:

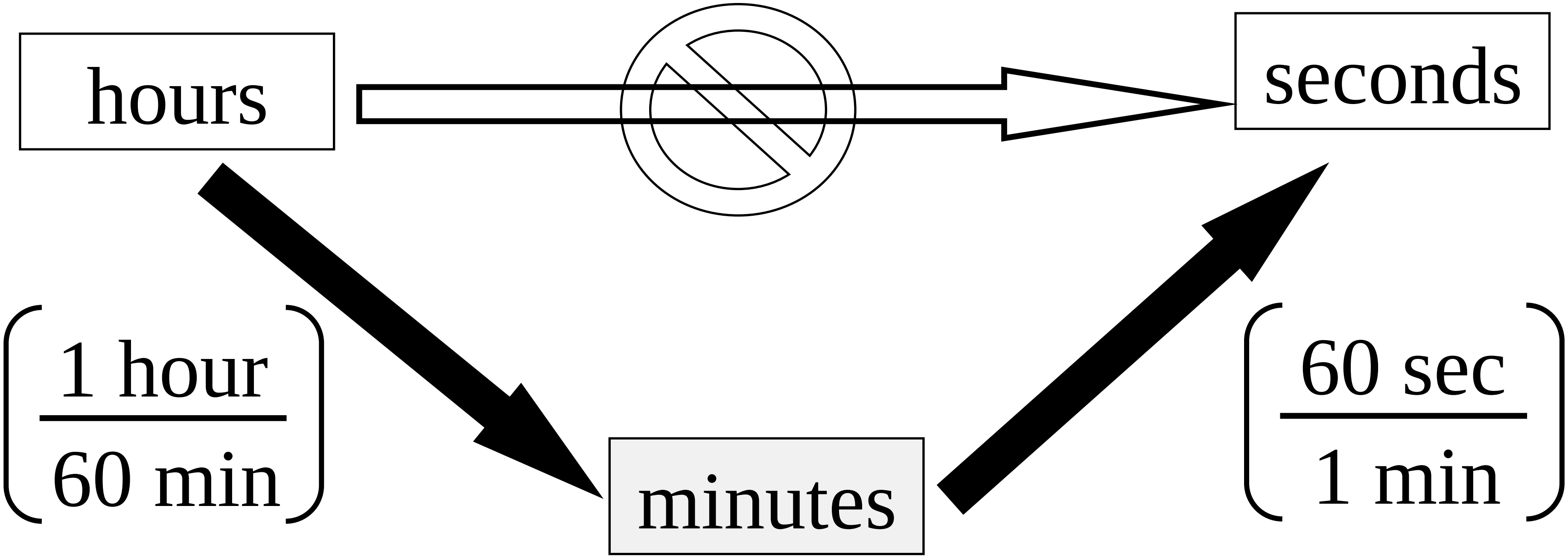
1) How many ft. (feet) in 379.3 in. (inches)?
   - 1 ft = 12 inches

2) How many eggs in 7.5 dozen?
   - 12 eggs = 1 dozen

3) How many calories in 514 joules?
   - 1 calorie = 4.184 joules

Sometimes it takes more than one step!

Example: How many seconds in 33.0 hours?

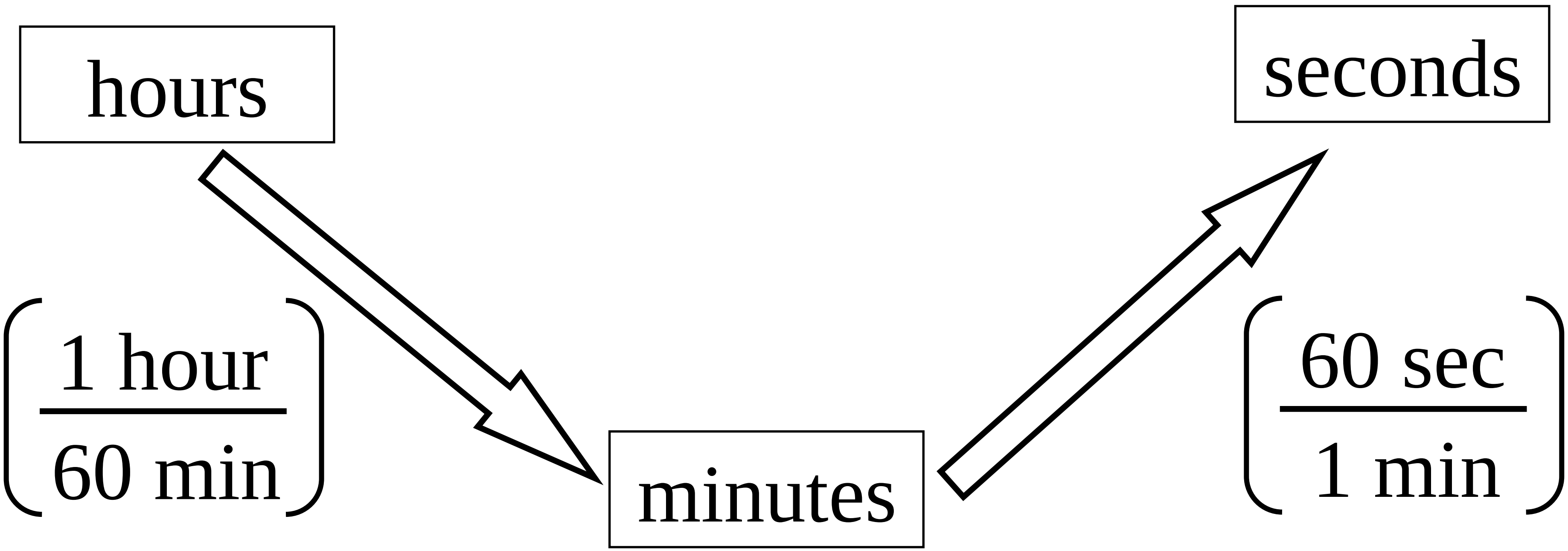


$$\frac{33.0\ \cancel{\text{hours}}}{} \left| \frac{60\ \text{min}}{1\ \cancel{\text{hour}}} \right| = \textbf{\textit{1980 min}}$$

$$\frac{\textbf{\textit{1980}}\ \cancel{\textbf{\textit{min}}}}{} \left| \frac{60\ \text{sec}}{1\ \cancel{\text{min}}} \right| = \textbf{\textit{119000 sec}}$$

If you wish, you can put these two calculations together:

Example: How many seconds in 33.0 hours?



$$\frac{33.0\ \text{hours}}{} \left|\ \frac{60\ \text{min}}{1\ \text{hour}}\ \right|\ \frac{60\ \text{sec}}{1\ \text{min}} \right| = 119000\ \text{sec}$$

Now you try a two step conversion:

How many inches in 5.5 meters given that:

1 inch = 2.54 cm

1 cm = .01 m

Metric Prefixes

Earlier, we used scientific notation to simplify working with very large or very small numbers.

Googol (10^{100})

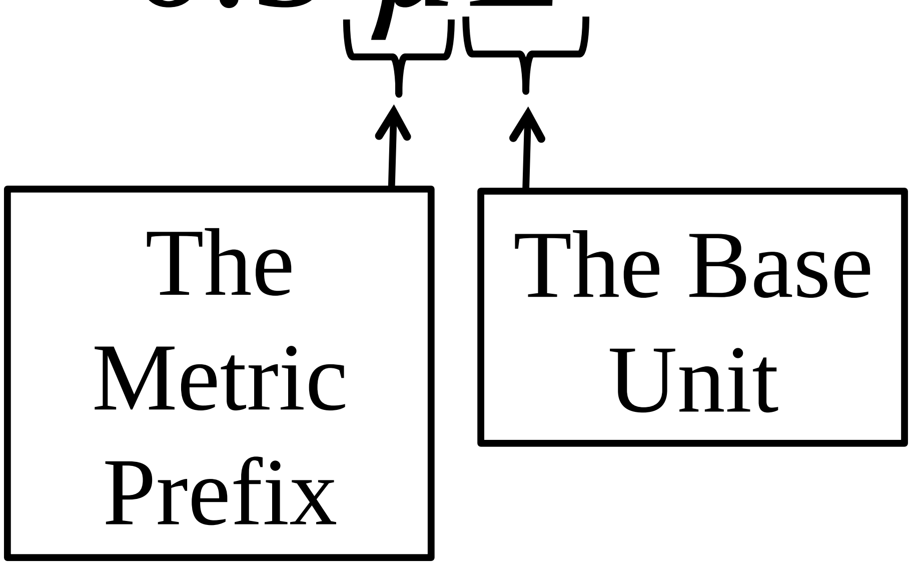
10, 000, 000, 000, 000,
000, 000, 000, 000, 000,
000, 000, 000, 000, 000,
000, 000, 000, 000, 000,
000, 000, 000, 000, 000,
000, 000, 000, 000, 000,
000, 000, 000, 000

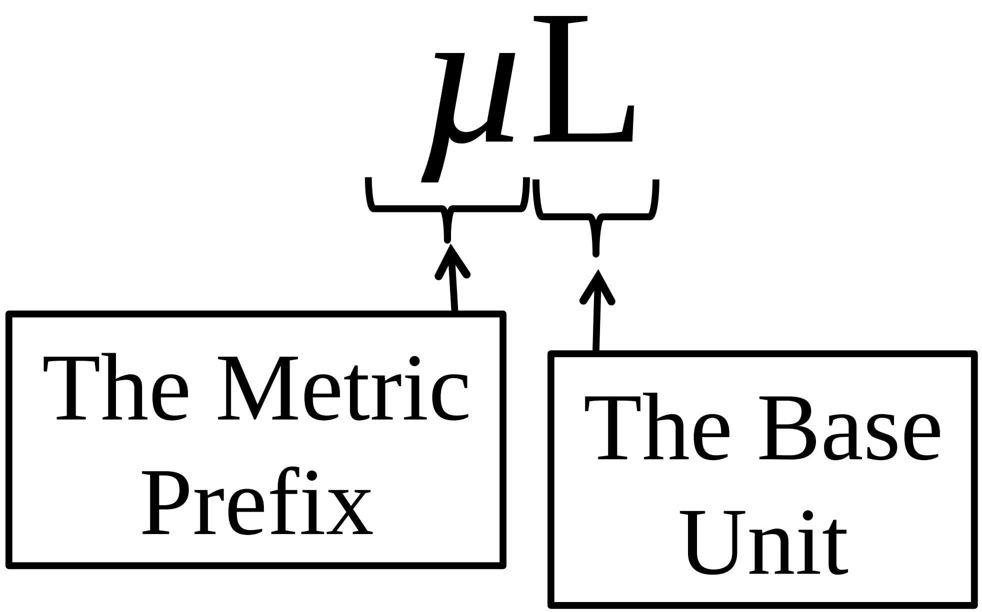
Another way to simplify working with large or small numbers is to use metric ***prefixes***.

Example: The volume of blood required for diabetics to measure blood glucose levels in modern glucometers is about 0.0000005 L.

It is more practical to use a metric prefix and say:

$$0.5\ \mu\text{L}$$





The metric prefix tells the fraction or multiple of the base unit(s).

- For example, 1×10^6 μL = 1 L

The base unit can be ***<u>any</u>*** metric unit:

- liter (L),gram (g), meter (m), joule (J), second (s), calorie (cal)…etc.

Unit Conversions Within The Metric System

Example: The volume of blood required to measure blood glucose levels in modern glucometers is about 0.0000005 L.

How can we convert that to μL ?

We need the relationship between L and μL to get the conversion factor.

Unit Conversions Within The Metric System

We will use the "Equality Table":

1 base unit =	
10 d (deci-)	0.1 da (deca-)
100 c (centi-)	.01 h (hecto)
1000 m (milli-)	.001 k (kilo)
1 x 10^6 μ (micro-)	1 x 10^{-6} M (mega-)
1 x 10^9 n (nano)	1 x 10^{-9} G (giga)

All these quantities are equal; any pair can used as conversion factors!!!

Example: What is the relationship between L (microliters) and liters (L)?

Equivalence statement: 1 L = 1 x 10^6 μL

1 base unit (Liters in this problem) =	
10 d (deci-)	0.1 da (deca-)
100 c (centi-)	.01 h (hecto)
1000 m (milli-)	.001 k (kilo)
1 x 10^6 μ (micro-)	1 x 10^{-6} M (mega-)
1 x 10^9 n (nano)	1 x 10^{-9} G (giga)

The equality table works for any unit!

1 base unit =	
10 d (deci-)	0.1 da (deca-)
100 c (centi-)	.01 h (hecto)
1000 m (milli-)	.001 k (kilo)
1×10^6 μ (micro-)	1×10^{-6} M (mega-)
1×10^9 n (nano)	1×10^{-9} G (giga)

The ***base*** ***unit*** could be gram (g), meter (m), liter (L), joule (J), second (s), mole (mol), calorie (cal)…etc.

Find the relationships between the following:

________ L = ________ mL

________ kg = ________ mg

________ nm = ________ m

________ cm = ________ mm

1 base unit =	
10 d (deci-)	0.1 da (deca-)
100 c (centi-)	.01 h (hecto)
1000 m (milli-)	.001 k (kilo)
1 x 10^6 μ (micro-)	1 x 10^{-6} M (mega-)
1 x 10^9 n (nano)	1 x 10^{-9} G (giga)

Example: How many μL (microliters) in 0.0000005 L?

Equivalence statement: 1 L = 1×10^6 μL

$$\left(\frac{1 \times 10^6\ \mu L}{1\ L}\right) \Longleftrightarrow \text{Conversion Factors} \Longleftrightarrow \left(\frac{1\ L}{1 \times 10^6\ \mu L}\right)$$

$$\frac{0.0000005\ L}{} \bigg| \bigg| =$$

1 base unit (Liters in this problem) =	
10 d (deci-)	0.1 da (deca-)
100 c (centi-)	.01 h (hecto)
1000 m (milli-)	.001 k (kilo)
1×10^6 μ (micro-)	1×10^{-6} M (mega-)
1×10^9 n (nano)	1×10^{-9} G (giga)

Example: How many mL (milliliters) in 0.0345 (kL) kiloliters ?

=

1 base unit (Liters in this problem) =	
10 d (deci-)	0.1 da (deca-)
100 c (centi-)	.01 h (hecto)
1000 m (milli-)	.001 k (kilo)
1 x 10^6 μ (micro-)	1 x 10^{-6} M (mega-)
1 x 10^9 n (nano)	1 x 10^{-9} G (giga)

You try one:

A vial contains 9758 mg of blood serum. Convert this into grams (g).

Temperature Unit Conversions

$$°F = (1.8 \times °C) + 32$$

$$°C = \frac{(°F - 32)}{1.8}$$

$$K = °C + 273.15$$

- Note: the 273.15, 32, and 1.8 are ***exact***.

Significant Figures in Equations with Mixed Operations:

When doing a calculation that involves ***only*** multiplication and/or division, you can do the entire calculation then round the answer to the correct number of significant figures at the end. The same is true for a calculation that involves ***only*** addition and/or subtraction.

But what about a calculation that involves mixed operations: <u>**both**</u> multiplication or division <u>***and***</u> addition or subtraction?

Significant Figures in Equations with Mixed Operations:

- $°F = (1.8 \times °C) + 32$

- $°C = \dfrac{(°F - 32)}{1.8}$

When doing calculations that involve **<u>both</u>** *multiplication or division* ***and*** *addition or subtraction*, **first** do a calculation for the operation *shown in parenthesis* and round that value to the correct number of significant figures, **then** use the rounded number to carry out the next operation.

On a warm summer day, the temperature reaches 85 °F. What is this temperature in °C?

$$°C = \frac{(°F - 32)}{1.8}$$

Note: First, you will do the subtraction (operation in parenthesis) and round the calculated value to the correct number of significant figures based on the rule for addition/ subtraction.

Next, you will divide that rounded number by *exactly* 1.8 (exactly 1.8 = 1.80000....) then round the calculated value to the correct number of significant figures using the rule for multiplication/division.

Example:

On a warm summer day, the temperature reaches 85 °F. What is this temperature in °C?