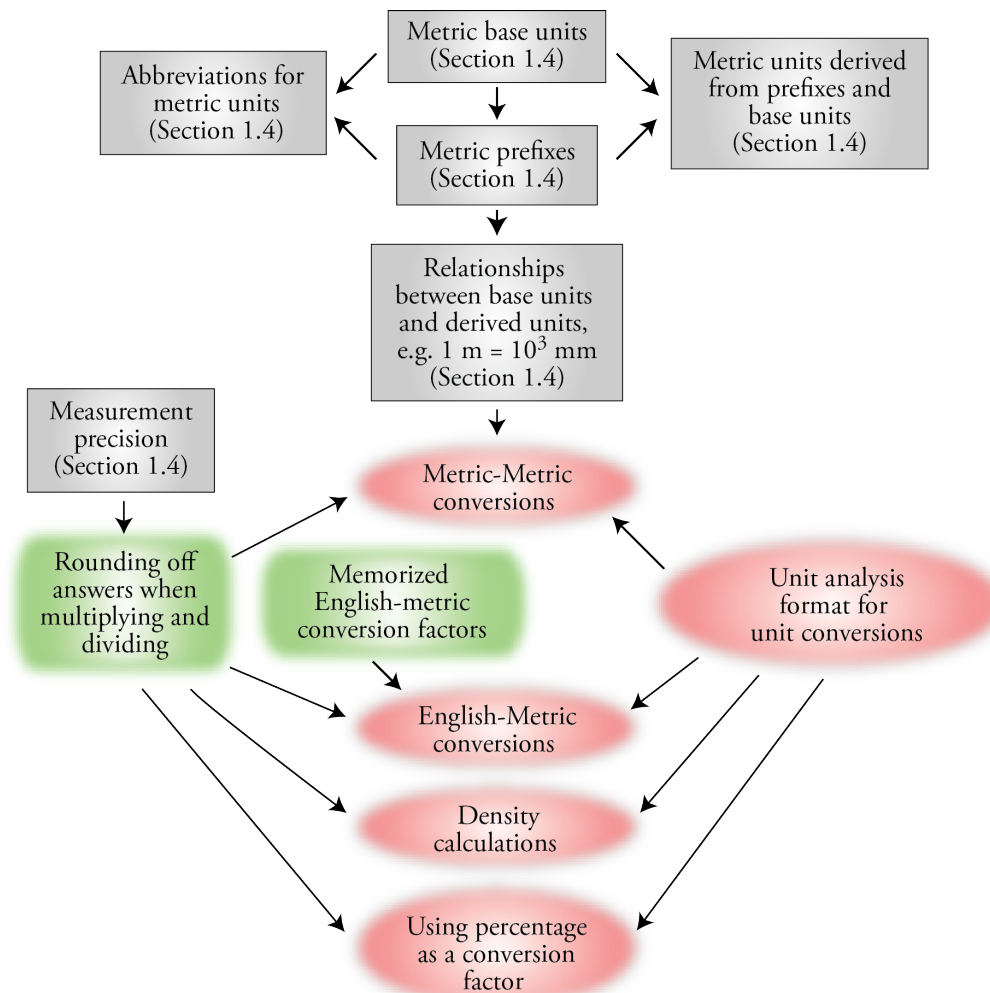




Chapter 2  
Unit Conversions

***An Introduction to Chemistry***  
by Mark Bishop

# Chapter Map



# Unit Conversions



*All science requires mathematics. The knowledge of mathematical things is almost innate in us. . . [Mathematics] is the easiest of sciences, a fact which is obvious in that no one's brain rejects it...*

Roger Bacon (c. 1214-c. 1294)

*Stand firm in your refusal to remain conscious during algebra. In real life, I assure you, there is no such thing as algebra.*

Fran Lebowitz (b. 1951)

# General “Unit Analysis” Procedure and Terminology

- Let’s convert 0.00003617 kg to milligrams (mg).
- Be patient...take small steps.
- **Step 1:** Identify the unit or units you want, and set that equal to the unit or units you are given. (See Appendix A for a list of units and their abbreviations.)

$$\begin{array}{c} \text{Desired unit} \\ \diagdown \\ ? \text{ mg} = 0.00003617 \text{ kg} \\ \diagup \\ \text{Given value} \end{array}$$


# General Procedure and Terminology

- **Step 2:** Multiply the expression to the right of the equals sign by one or more conversion factors that cancel the unwanted units and generate the desired unit.
  - Units can be cancelled like algebraic variables.
  - Set up the skeleton of the next conversion with the first unit you want to cancel in correct position.

Skeleton to cancel the first unit

$$? \text{ mg} = 0.00003617 \text{ ~~kg~~} \left( \frac{\text{ / }}{\text{ ~~kg~~}} \right)$$

# General Procedure and Terminology



- **Step 2 (cont.):**
  - Ask yourself, “Do I know a conversion factor that will take me directly from the unit I have to the unit I want?”.
  - In this case, “Do I know how many mg there are per kg?”.
  - If the answer is no, ask yourself, “What type of unit do I have, and what type of unit do I want?”.
  - In this case, we are converting from one SI mass unit to another SI mass unit.
  - Apply strategies you have learned for different types of conversions.

# General Procedure and Terminology

- **Step 2 (cont.):**


- When converting from one SI unit to another SI unit for the same type of measurement (such as both mass), convert from the unit you have to the base unit and then from the base unit to the unit you want. See Table 1.1 on page 11 of the text.

Skeleton to convert from the unit  
you have to the base unit

$$? \text{ mg} = 0.00003617 \text{ kg} \left( \frac{\cancel{\text{g}}}{\cancel{\text{kg}}} \right) \left( \frac{\text{mg}}{\cancel{\text{g}}} \right)$$

Skeleton to convert from the base unit  
to the unit you want

# Metric-Metric Conversion Factors



- The relationships between metric (SI) units can be derived from the metric prefixes. (See Table 1.2 for a useful list of metric prefixes.)
- These relationships can easily be translated into conversion factors.
  - For example, two possible sets of conversion factors for relating milliliters to liters can be obtained from the definition of the prefix *milli*.
  - *Milli* is defined as  $10^{-3}$ , so  $1 \text{ mL} = 10^{-3} \text{ L}$ .
  - If a milliliter is 1/1000 of a liter, there must be 1000 milliliters in a liter, so  $10^3 \text{ mL} = 1 \text{ L}$ .



# Metric-Metric Conversion Factors

- The relationship between milliliters and liters yields two possible sets of conversion factors.

$$10^3 \text{ mL} = 1 \text{ L} \text{ leads to } \frac{10^3 \text{ mL}}{1 \text{ L}} \quad \text{or} \quad \frac{1 \text{ L}}{10^3 \text{ mL}}$$

$$1 \text{ mL} = 10^{-3} \text{ L} \text{ leads to } \frac{1 \text{ mL}}{10^{-3} \text{ L}} \quad \text{or} \quad \frac{10^{-3} \text{ L}}{1 \text{ mL}}$$

# General Procedure and Terminology

- **Step 2 (cont.):**

- Set up your conversion factors to cancel the units that you do not want and generate the units that you do want.
- Because the k in kg represents kilo, and because kilo is defined as  $10^3$ , there must be  $10^3$  g per kg.

Converts given SI unit to SI base unit

$$? \text{ mg} = 0.00003617 \text{ kg} \left( \frac{10^3 \text{ g}}{1 \text{ kg}} \right)$$

# General Procedure and Terminology

- **Step 2 (cont.):.**

- Our next step is to set up the skeleton for the conversion of grams to milligrams.

$$? \text{ mg} = 0.00003617 \cancel{\text{kg}} \left( \frac{10^3 \cancel{\text{g}}}{1 \cancel{\text{kg}}} \right) \left( \frac{\quad}{\text{g}} \right)$$

Skeleton to convert SI base unit  
to desired SI unit

# General Procedure and Terminology

- **Step 2 (cont.):.**

- We can use either of the following conversion factors.

$$\left( \frac{1 \text{ mg}}{10^{-3} \text{ g}} \right) \text{ or } \left( \frac{10^3 \text{ mg}}{1 \text{ g}} \right)$$

- I recommend using the positive exponentials, so I recommend the second conversion factor.
- The positive exponent goes with the smaller unit.

$$? \text{ mg} = 0.00003617 \text{ kg} \left( \frac{10^3 \cancel{\text{g}}}{1 \cancel{\text{kg}}} \right) \left( \frac{10^3 \text{ mg}}{1 \cancel{\text{g}}} \right)$$

Converts SI base unit to desired SI unit

# General Procedure and Terminology

- **Step 3:** Check to be sure you used correct conversion factors and that your units cancel to yield the desired unit.

$$? \text{ mg} = 0.00003617 \text{ kg} \left( \frac{10^3 \text{ g}}{1 \text{ kg}} \right) \left( \frac{10^3 \text{ mg}}{1 \text{ g}} \right)$$

# General Procedure and Terminology

- **Step 4:** Do the calculation, rounding your answer to the correct number of significant figures and combining it with the correct unit. (See Section 2.2 for the rules for rounding.)

Desired unit

Converts given SI unit to metric base unit

$$? \text{ mg} = 0.00003617 \text{ kg} \left( \frac{10^3 \text{ g}}{1 \text{ kg}} \right) \left( \frac{10^3 \text{ mg}}{1 \text{ g}} \right) = 36.17 \text{ mg}$$

Given value

Converts SI base unit to desired metric unit

# General “Unit Analysis” Procedure and Terminology

- Let’s convert 254 meters (m) to feet (ft).
- **Step 1:** Identify the unit or units you want, and set that equal to the unit or units you are given.

Desired unit  
/

$$? \text{ ft} = 254 \text{ m}$$

/

Given value

# General Procedure and Terminology

- Set up the skeleton of the next conversion with the first unit you want to cancel in the correct position.

$$? \text{ ft} = 254 \cancel{\text{ m}} \left( \frac{\quad}{1 \cancel{\text{ m}}} \right)$$

- Ask yourself, “Do I know a conversion factor that will take me directly from the unit I have to the unit I want?”. In this case, “Do I know how many feet there are per meter?”.
- If the answer is no, ask yourself, “What type of unit do I have, and what type of unit do I want?”.
- In this case, we are converting from an SI length unit to an English length unit.
- Apply strategies you have learned for different types of conversions.



# English-Metric Conversion Factors

**Table 2.1**

Type of Measurement	Probably Most Useful to Know	Others Useful to Know		
Length	$\frac{2.54 \text{ cm}}{1 \text{ in.}}$	$\frac{1.609 \text{ km}}{1 \text{ mi}}$	$\frac{39.37 \text{ in.}}{1 \text{ m}}$	$\frac{1.094 \text{ yd}}{1 \text{ m}}$
Mass	$\frac{453.6 \text{ g}}{1 \text{ lb}}$	$\frac{2.205 \text{ lb}}{1 \text{ kg}}$		
Volume	$\frac{3.785 \text{ L}}{1 \text{ gal}}$	$\frac{1.057 \text{ qt}}{1 \text{ L}}$		

# General Procedure and Terminology

- **Step 2 (cont.):**

- We will use 2.54 cm/1 in., which is exact.
- We can view our problem as three steps: m to cm, cm to in., and in. to ft.
- We can think of the steps one at a time.

$$? \text{ ft} = 254 \cancel{\text{ m}} \left( \frac{\quad}{1 \cancel{\text{ m}}} \right)$$

$$? \text{ ft} = 254 \cancel{\text{ m}} \left( \frac{10^2 \cancel{\text{ cm}}}{1 \cancel{\text{ m}}} \right)$$

# General Procedure and Terminology

- **Step 2 (cont.):.**

- Now we can do the core conversion from centimeters to inches.

$$? \text{ ft} = 254 \cancel{\text{ m}} \left( \frac{10^2 \cancel{\text{ cm}}}{1 \cancel{\text{ m}}} \right) \left( \frac{\quad}{\cancel{\text{ cm}}} \right)$$

$$? \text{ ft} = 254 \cancel{\text{ m}} \left( \frac{10^2 \cancel{\text{ cm}}}{1 \cancel{\text{ m}}} \right) \left( \frac{1 \text{ in.}}{2.54 \cancel{\text{ cm}}} \right)$$

# General Procedure and Terminology

- **Step 2 (cont.):.**

- The last conversion is from inches to feet.

$$? \text{ ft} = 254 \cancel{\text{ m}} \left( \frac{10^2 \cancel{\text{ cm}}}{1 \cancel{\text{ m}}} \right) \left( \frac{1 \cancel{\text{ in.}}}{2.54 \cancel{\text{ cm}}} \right) \left( \frac{\quad}{\cancel{\text{ in.}}} \right)$$

$$? \text{ ft} = 254 \cancel{\text{ m}} \left( \frac{10^2 \cancel{\text{ cm}}}{1 \cancel{\text{ m}}} \right) \left( \frac{1 \cancel{\text{ in.}}}{2.54 \cancel{\text{ cm}}} \right) \left( \frac{1 \text{ ft}}{12 \cancel{\text{ in.}}} \right)$$

# General Procedure and Terminology

- **Step 3:** Check to be sure you used correct conversion factors and that your units cancel to yield the desired unit.

$$? \text{ ft} = 254 \cancel{\text{ m}} \left( \frac{10^2 \cancel{\text{ cm}}}{1 \cancel{\text{ m}}} \right) \left( \frac{1 \cancel{\text{ in.}}}{2.54 \cancel{\text{ cm}}} \right) \left( \frac{1 \text{ ft}}{12 \cancel{\text{ in.}}} \right)$$


- **Step 4:** Do the calculation, rounding your answer to the correct number of significant figures and combining it with the correct unit.

$$? \text{ ft} = 254 \cancel{\text{ m}} \left( \frac{10^2 \cancel{\text{ cm}}}{1 \cancel{\text{ m}}} \right) \left( \frac{1 \cancel{\text{ in.}}}{2.54 \cancel{\text{ cm}}} \right) \left( \frac{1 \text{ ft}}{12 \cancel{\text{ in.}}} \right) = 833 \text{ ft}$$

# Reporting Values from Measurements

- The numbers in a values should reflect both the magnitude (size) of the value and the value's uncertainty. (e.g. 185 lb)
- Measurements never give exact values. ( $185 \pm 2$  lb)
- Unless we are told otherwise, we assume that values from measurements have an uncertainty of plus or minus one in the last decimal place reported. (e.g. 184-186 lb)
  - 5 g says could be from 4 g to 6 g
  - 5.00 g says 4.99 g to 5.01 g
  - 5.000 g says 4.999 g to 5.001 g

# Reporting Values from Calculations



- If a calculation is performed using all exact numbers and if the answer is not rounded off, the answer is exact, but this is a rare occurrence.
- The answers derived from calculations are usually not exact.
- As we do for measurements, unless we are told otherwise, we assume that values from calculations have an uncertainty of plus or minus one in the last decimal place reported.
- The main purpose of this lesson is to show you a simple way to round off the answers to your calculations in a way that reflects the proper degree of uncertainty.

# Reporting Values from Calculations (Example)

- For example, Consider the following calculation that converts 4.9800 g to lb.

$$? \text{ lb} = 4.9800 \text{ g} \left( \frac{1 \text{ lb}}{453.6 \text{ g}} \right) = 0.01098 \text{ lb (or } 1.098 \times 10^{-2} \text{ lb)}$$

- 4.98 divided by 453.6 is 0.01097883597884, which suggests an uncertainty of 0.0000000000000001 lb.
- The value 453.6 g suggests an uncertainty of 0.1 g, which is about 0.0002 lb, so we do not want to report our answer as 0.01097883597884 lb, which suggests an uncertainty of 0.0000000000000001 lb.
- We need a simple technique for rounding our answers to calculations to reflect the uncertainty of the numbers used in the calculations.



# Rounding Answers from Multiplication and Division Step 1

- **Step 1:** Determine whether each value is exact, and ignore exact values.
  - Exact values
    - Numbers that come from definitions are exact.
    - Numbers derived from counting are exact.
  - Do Step 2 for values that are not exact.
    - Values that come from measurements are never exact. (e.g. 4.9800 g)
    - We will assume that values derived from calculations are not exact unless otherwise indicated.

# Exact Values

- Numbers that come from definitions are exact.
  - Numbers derived from the definitions of the metric prefixes are exact, such as

$$\frac{10^3 \text{ g}}{1 \text{ kg}}$$

- Numbers in English-English conversion factors with the same type of unit top and bottom are exact, such as

$$\frac{12 \text{ in.}}{1 \text{ ft}}$$

- The 2.54 in the following conversion factor is exact

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

- Numbers derived from counting are exact.  $\frac{5 \text{ toes}}{1 \text{ foot}}$

# Rounding Answers from Multiplication and Division Step 1 (Numbers that are not exact)

- Do Step 2 for values that are not exact.
  - Values that come from measurements are never exact.
  - We will assume that values derived from calculations are not exact unless otherwise indicated.
  - Except for 2.54 cm/1 in. (in which the 2.54 is defined and exact), we will assume that all of the English-metric conversion factors that we see have numbers that were calculated and rounded. For example, the 453.6 in the following conversion factor is not exact.

$$\frac{453.6 \text{ g}}{1 \text{ lb}}$$

# Rounding Answers from Multiplication and Division Step 1 Example

- The following shows how we can convert the mass of a hydrogen atom from micrograms to pounds.

$$? \text{ lb} = 1.67 \times 10^{-18} \cancel{\mu\text{g}} \left( \frac{1 \cancel{\text{g}}}{10^6 \cancel{\mu\text{g}}} \right) \left( \frac{1 \text{ lb}}{453.6 \cancel{\text{g}}} \right)$$

- **Step 1:** Determine whether each value is exact, and ignore exact values.
  - The mass of a hydrogen atom is not defined, and masses cannot be counted, so the  $1.67 \times 10^{-18}$  is not exact.
  - The  $10^6$  number comes from the definition of micro, so it is exact. We will ignore this number when rounding.
  - Except for 2.54 cm per inch, all of the English-metric conversion factors that we will see are calculated and rounded, so the 453.6 is not exact.

# Rounding Answers from Multiplication and Division Step 2

- **Step 2:** Determine the number of significant figures in each value that is not exact.

- All non-zero digits are significant.

$\begin{array}{c} 1 \quad 3 \quad 5 \\ 11.275 \text{ g} \\ 2 \quad 4 \end{array}$  ——— Five significant figures

- Zeros between nonzero digits are significant.

A zero between nonzero digits

$\begin{array}{c} 1 \quad | \quad 3 \quad 5 \\ 10.275 \text{ g} \\ 2 \quad 4 \end{array}$  ——— Five significant figures

# Rounding Answers from Multiplication and Division Step 2 (cont.)

- **Step 2:** Determine the number of significant figures in each value that is not exact.
  - Zeros to the left of nonzero digits are not significant.

Not significant figures

$0.000102 \text{ kg}$  which can be described as  $1.02 \times 10^{-4} \text{ kg}$

Both have three significant figures.

# Rounding Answers from Multiplication and Division Step 2 (cont.)

- **Step 2:** Determine the number of significant figures in each value that is not exact.
  - Zeros to the right of nonzero digits in numbers that include decimal points are significant.

Five significant figures — 10.200 g      20.0 mL — Three significant figures

1 3 5      1 3  
2 4      2

Unnecessary for reporting size of value,  
but do reflect degree of uncertainty.

# Rounding Answers from Multiplication and Division Step 2 (cont.)

- Zeros to the right of nonzero digits in numbers without decimal points are ambiguous for significant figures.

$220 \text{ kg}$

Precise to  $\pm 1 \text{ kg}$  or  $\pm 10 \text{ kg}$ ?  
Two or three significant figures?

Important for reporting size of value, but unclear about degree of uncertainty

$2.2 \times 10^2 \text{ kg}$

Use scientific notation to remove ambiguity.

$2.20 \times 10^2 \text{ kg}$



# Rounding Answers from Multiplication and Division Step 2 Example

- The following shows how we can convert the mass of a hydrogen atom from micrograms to pounds.

$$? \text{ lb} = 1.67 \times 10^{-18} \cancel{\mu\text{g}} \left( \frac{1 \cancel{\text{g}}}{10^6 \cancel{\mu\text{g}}} \right) \left( \frac{1 \text{ lb}}{453.6 \cancel{\text{g}}} \right)$$

- **Step 2:** Determine the number of significant figures in each value that is not exact.
  - For  $1.67 \times 10^{-18}$ , the uncertainty is reflected by the 1.67, which has three nonzero digits and thus three significant figures.
  - The 453.6 has four nonzero digits and thus four significant figures.

# Rounding Answers from Multiplication and Division Step 3

- **Step 3:** When multiplying and dividing, round your answer off to the same number of significant figures as the value used with the fewest significant figures.
  - If the digit to the right of the final digit you want to retain is less than 5, round down (the last digit remains the same).

26.221 rounded to three significant figures is 26.2

|  
First digit dropped is less than 5

# Rounding Answers from Multiplication and Division Step 3 (cont.)

- **Step 3:** When multiplying and dividing, round your answer off to the same number of significant figures as the value used with the fewest significant figures.
  - If the digit to the right of the final digit you want to retain is 5 or greater, round up (the last significant digit increases by 1).

26.272 rounded to three significant figures is 26.3

|  
First digit dropped is greater than 5.

26.2529 rounded to three significant figures is 26.3

|  
First digit dropped is equal to 5.

26.15 rounded to three significant figures is 26.2

|  
First digit dropped is equal to 5.

# Rounding Answers from Multiplication and Division Step 3 Example

- The following shows how we can convert the mass of a hydrogen atom from micrograms to pounds.

$$? \text{ lb} = 1.67 \times 10^{-18} \cancel{\mu\text{g}} \left( \frac{1 \cancel{\text{g}}}{10^6 \cancel{\mu\text{g}}} \right) \left( \frac{1 \text{ lb}}{453.6 \cancel{\text{g}}} \right)$$

- **Step 3:** When multiplying and dividing, round your answer off to the same number of significant figures as the value used with the fewest significant figures.
  - $1.67 \times 10^{-18}$  has three significant figures.
  - 453.6 has four significant figures.
  - Our answer should have three significant figures.

# Rounding Answers from Multiplication and Division Step 3 Example

- **Step 3:** When multiplying and dividing, round your answer off to the same number of significant figures as the value used with the fewest significant figures.
  - The calculated result is  $3.681657848325 \times 10^{-27}$ .
  - We round to three significant figures, and because the first number we are dropping is less than 5 (i.e. 1 in this case), we round our answer to  $3.68 \times 10^{-27}$ .

$$? \text{ lb} = 1.67 \times 10^{-18} \cancel{\mu\text{g}} \left( \frac{1 \cancel{\text{g}}}{10^6 \cancel{\mu\text{g}}} \right) \left( \frac{1 \text{ lb}}{453.6 \cancel{\text{g}}} \right) = 3.68 \times 10^{-27} \text{ lb}$$

Example 2.4: The average human body contains 5.2 L of blood. What is this volume in quarts?

$$? \text{ qt} = 5.2 \cancel{\text{L}} \left( \frac{1 \cancel{\text{gal}}}{3.785 \cancel{\text{L}}} \right) \left( \frac{4 \text{ qt}}{1 \cancel{\text{gal}}} \right) = 5.5 \text{ qt}$$

- We input the following into the calculator:

$$5.2 \div 3.785 \times 4 =$$

- The answer on the display of the calculator is 5.49537649. We do not write this down until we decide on the number of significant figures in our answer.
- 5.2 has 2 sig. figs., 3.785 has 4 sig. figs., and 4 is exact, so we round our answer to 2 significant figures.
- The first digit we are dropping is 9, so we round up to 5.5.

# Example 2.5: How many minutes does it take an ant walking at 0.01 m/s to travel 6.0 feet across a picnic table?

$$\begin{array}{ccccccc}
 \text{2 significant figures} & & \text{Exact} & & & & \text{1 significant figure} \\
 \diagdown & & \diagdown & & & & \diagdown \\
 ? \text{ min} = 6.0 \text{ ft} & \left( \frac{12 \text{ in.}}{1 \text{ ft}} \right) & \left( \frac{2.54 \text{ cm}}{1 \text{ in.}} \right) & \left( \frac{1 \text{ m}}{10^2 \text{ cm}} \right) & \left( \frac{1 \text{ s}}{0.01 \text{ m}} \right) & \left( \frac{1 \text{ min}}{60 \text{ s}} \right) & = 3 \text{ min} \\
 & & \diagup & \diagup & \diagup & \diagdown & \\
 & & \text{Exact} & \text{1 significant figure} & \text{Exact} & & 
 \end{array}$$

- We input the following into the calculator:

$$6 \times 12 \times 2.54 \div 100 \div .01 \div 60 =$$

- The answer on the display of the calculator is 3.048.
- 6.0 has 2 sig. figs., 0.01 has 1 sig. fig., and 12, 2.54,  $10^2$ , and 60 are exact, and so we round our answer to 1 significant figure.
- The first digit we are dropping is 0, so we round down to 3.

# Rounding Answers from Addition and Subtraction

- **Step 1:** Determine whether each value is exact, and ignore exact values.
  - Skip exact values.
  - Do Step 2 for values that are not exact.
- **Step 2:** Determine the number of decimal positions for each value that is not exact.
- **Step 3:** Round your answer to the same number of decimal positions as the inexact value with the fewest decimal places.



## Example 2.6

- A laboratory procedure calls for you to determine the mass of an unknown liquid. Let's suppose that you weigh a 100-mL beaker on a new electronic balance and record its mass as 52.3812 g. You then add 10 mL of the unknown liquid to the beaker and discover that the electronic balance has stopped working. You find a 30-year-old balance in a cupboard, use it to weigh the beaker of liquid, and record that mass as 60.2 g. What is the mass of the unknown liquid?

$$60.2 \text{ g beaker with liquid} - 52.3812 \text{ g beaker} = 7.8188 \text{ g liquid}$$

## Example 2.6

60.2 g beaker with liquid – 52.3812 g beaker = 7.8188 g liquid

- 60.2 is uncertain in the tenth position ( $\pm 0.1$ ).
- 52.3812 is uncertain in the fourth decimal position ( $\pm 0.0001$ ).
- We round our answer off to the tenth position (7.8 g)

60.2 g beaker with liquid – 52.3812 g beaker = 7.8 g liquid

# Density




- ***Mass density*** is mass divided by volume. It is usually just called density.

$$\text{density} = \frac{\text{mass}}{\text{volume}}$$

- Density usually has the units of g/mL for solids and liquids and g/L for gases.

# Uses for Density



- It is possible to identify an unknown substance by comparing its density at a particular temperature to the densities of known substances at that temperature. (See Table 2.2 for a list of densities.)
- Density can be used as a unit analysis conversion factor that converts mass to volume or volume to mass.

## Example 2.7: What is the volume of 25.00 kg water at 20 °C?

- Don't forget to take small steps.
  - Set the unit you want to get equal to the value you are given.

$$? \text{ L H}_2\text{O} = 25.00 \text{ kg H}_2\text{O}$$

- Set up the skeleton of the first conversion factor.

$$? \text{ L H}_2\text{O} = 25.00 \cancel{\text{ kg}} \text{ H}_2\text{O} \left( \frac{\quad}{\cancel{\text{ kg}}} \right)$$

- Ask yourself whether you know a conversion factor that will take you directly from the unit that you have to the unit you want. (In this case, let's assume that the answer is no.)

The background of the slide features a dark blue gradient. On the left, there is a silhouette of a lighthouse on a small island. On the right, there are several water molecules represented as red and white spheres. The text is overlaid on the top left in a light blue, sans-serif font.

## Example 2.8: What is the volume of 25.00 kg water at 20 °C?

- If you can't see how to make your conversion in one step, ask yourself what type of unit you're converting from and what type of unit are you converting to.
- In this case, we are converting from mass into volume, which is our tip-off that we need the density, which you are likely to get from a table of densities.
- The density of water at 20 °C is 0.9982 g/mL.

## Example 2.8: What is the volume of 25.00 kg water at 20 °C?

- Like any other conversion factor, density can be used in two forms. The first form below can be used to convert from volume in mL to mass in g, and the second form can be used to convert mass in g to volume in mL. We will use the second form.

$$\frac{0.9982 \text{ g H}_2\text{O}}{1 \text{ mL}} \quad \text{or} \quad \frac{1 \text{ mL}}{0.9982 \text{ g H}_2\text{O}}$$

## Example 2.8: What is the volume of 25.00 kg water at 20 °C?

- Before we can convert from g to mL, we need to convert from kg to g.
- After converting to mL with the density, we convert from mL to L.
- Our plan is to use three conversion factors to make the following conversions.

kg → g → mL → L

$$? \text{ L H}_2\text{O} = 25.00 \text{ kg H}_2\text{O} \left( \frac{\cancel{\text{g}}}{\cancel{\text{kg}}} \right) \left( \frac{1 \cancel{\text{ mL H}_2\text{O}}}{0.9982 \cancel{\text{ g H}_2\text{O}}} \right) \left( \frac{\text{L}}{\cancel{\text{ mL}}} \right)$$



# Example 2.8: What is the volume of 25.00 kg water at 20 °C?

- Let's go back to the beginning of our setup.

$$? \text{ L H}_2\text{O} = 25.00 \text{ kg H}_2\text{O} \left( \frac{\quad}{\text{kg}} \right)$$

- We get the conversion factor that will take us from kg to g from the definition of kilo, which is  $10^3$ .

$$? \text{ L H}_2\text{O} = 25.00 \text{ kg H}_2\text{O} \left( \frac{10^3 \text{ g}}{1 \text{ kg}} \right)$$

- Next, we set up the skeleton of the next conversion factor.

$$? \text{ L H}_2\text{O} = 25.00 \text{ kg H}_2\text{O} \left( \frac{10^3 \text{ g}}{1 \text{ kg}} \right) \left( \frac{\quad}{\text{g H}_2\text{O}} \right)$$

## Example 2.8: What is the volume of 25.00 kg water at 20 °C?

- Our core conversion is to convert g to mL using the density.

$$? \text{ L H}_2\text{O} = 25.00 \cancel{\text{ kg H}_2\text{O}} \left( \frac{10^3 \cancel{\text{ g}}}{1 \cancel{\text{ kg}}} \right) \left( \frac{1 \text{ mL H}_2\text{O}}{0.9982 \cancel{\text{ g H}_2\text{O}}} \right)$$

- We now set up the skeleton for the next conversion.

$$? \text{ L H}_2\text{O} = 25.00 \cancel{\text{ kg H}_2\text{O}} \left( \frac{10^3 \cancel{\text{ g}}}{1 \cancel{\text{ kg}}} \right) \left( \frac{1 \cancel{\text{ mL H}_2\text{O}}}{0.9982 \cancel{\text{ g H}_2\text{O}}} \right) \left( \frac{\quad}{\cancel{\text{ mL}}} \right)$$

- We get the conversion factor that will take us from mL to L from the definition of milli, which is  $10^{-3}$ .

$$? \text{ L H}_2\text{O} = 25.00 \cancel{\text{ kg H}_2\text{O}} \left( \frac{10^3 \cancel{\text{ g}}}{1 \cancel{\text{ kg}}} \right) \left( \frac{1 \cancel{\text{ mL H}_2\text{O}}}{0.9982 \cancel{\text{ g H}_2\text{O}}} \right) \left( \frac{1 \text{ L}}{10^3 \cancel{\text{ mL}}} \right)$$

Example 2.8: What is the volume of 25.00 kg water at 20 °C?

$$\begin{aligned} ? \text{ L H}_2\text{O} &= 25.00 \text{ kg H}_2\text{O} \left( \frac{10^3 \text{ g}}{1 \text{ kg}} \right) \left( \frac{1 \text{ mL H}_2\text{O}}{0.9982 \text{ g H}_2\text{O}} \right) \left( \frac{1 \text{ L}}{10^3 \text{ mL}} \right) \\ &= 25.05 \text{ L H}_2\text{O} \end{aligned}$$

- Check to be sure that you have used conversion factors for which you have high confidence that they are correct.
- Check to be sure that your units cancel.
- Use the calculator to complete the calculation (25.0450811).
- Report your answer to the correct significant figures.
- Be sure to add the correct unit to your answer.

Example 2.9: An empty 2-L graduated cylinder is found to have a mass of 1124.2 g. Liquid methanol,  $\text{CH}_3\text{OH}$ , is added to the cylinder, and its volume measured as 1.20 L. The total mass of the methanol and the cylinder is 2073.9 g, and the temperature of the methanol is  $20^\circ\text{C}$ . What is the density of methanol at this temperature?

- One of the most useful initial strategies for working word problems is to write down the values that you are given, separating them from the words.
- Sometimes you can just write down the number and unit. For example,
  - 1124.2 g
  - 1.20 L
  - 2073.9 g
- If the values refer to more than one thing, it's best to identify what your values refer to. For our example,
  - 1124.2 g grad cyl
  - 1.20 L  $\text{CH}_3\text{OH}$
  - 2073.9 g grad cyl/ $\text{CH}_3\text{OH}$

Example 2.9: An empty 2-L graduated cylinder is found to have a mass of 1124.2 g. Liquid methanol, CH<sub>3</sub>OH, is added to the cylinder, and its volume measured as 1.20 L. The total mass of the methanol and the cylinder is 2073.9 g, and the temperature of the methanol is 20 °C. What is the density of methanol at this temperature?

- Unless you are told otherwise, describe the density of solids and liquids with units of g/mL. When starting your unit analysis setup, be sure that it's clear that g is on the top and mL is on the bottom.

$$\frac{? \text{ g}}{\text{mL}} = \quad \text{not } ? \text{ g/mL} =$$

- Because we want our answer to contain a ratio of two units, we start the right side of our setup with a ratio of two units. Our next little step is to put a line on the right side of the equals sign.

$$\frac{? \text{ g}}{\text{mL}} = \frac{\quad}{\quad}$$

Example 2.9: An empty 2-L graduated cylinder is found to have a mass of 1124.2 g. Liquid methanol, CH<sub>3</sub>OH, is added to the cylinder, and its volume measured as 1.20 L. The total mass of the methanol and the cylinder is 2073.9 g, and the temperature of the methanol is 20 °C. What is the density of methanol at this temperature?

- Because we want mass on the top when we are done and volume on the bottom, we put our mass unit on the top and our volume unit on the bottom.
- The mass of the methanol is found by subtracting the mass of the cylinder from the total mass of the cylinder and the methanol.

$$\frac{? \text{ g}}{\text{mL}} = \frac{(2073.9 - 1124.2) \text{ g}}{1.20 \text{ L}}$$

Example 2.9: An empty 2-L graduated cylinder is found to have a mass of 1124.2 g. Liquid methanol, CH<sub>3</sub>OH, is added to the cylinder, and its volume measured as 1.20 L. The total mass of the methanol and the cylinder is 2073.9 g, and the temperature of the methanol is 20 °C. What is the density of methanol at this temperature?

- We now have units of g/L instead of g/mL, so we have to convert L to mL.
- Because we want to cancel L, and because it is on the bottom of the first ratio, the skeleton of the next conversion factor has the L on top.

$$\frac{? \text{ g}}{\text{mL}} = \frac{(2073.9 - 1124.2) \text{ g}}{1.20 \cancel{\text{L}}} \left( \frac{\cancel{\text{L}}}{\phantom{\text{L}}} \right)$$

- The m in mL represents milli, which is 10<sup>-3</sup>, so there must be 10<sup>3</sup> mL per L.

$$\frac{? \text{ g}}{\text{mL}} = \frac{(2073.9 - 1124.2) \text{ g}}{1.20 \cancel{\text{L}}} \left( \frac{1 \cancel{\text{L}}}{10^3 \text{ mL}} \right)$$

Example 2.9: An empty 2-L graduated cylinder is found to have a mass of 1124.2 g. Liquid methanol,  $\text{CH}_3\text{OH}$ , is added to the cylinder, and its volume measured as 1.20 L. The total mass of the methanol and the cylinder is 2073.9 g, and the temperature of the methanol is  $20^\circ\text{C}$ . What is the density of methanol at this temperature?

- Because the order of operations calls for multiplication and division to be done before addition and subtraction, it's a good habit to always put the numbers for additions and subtractions in parentheses when doing the calculations on your computer or calculator. One way to do this calculation is below. (Different calculators and computers call for inputting scientific notation for numbers in different ways, but in this case, we can use 1000 for  $10^3$ .)

$$(2073.9 - 1124.2) \div 1.2 \div 1000 =$$

- The computer shows the answer as 0.791416666666666666666666666666666667, but we need to round this to the correct significant figures.





# Conversion Factors from Percentages and Percentage Calculations



- *Percentage by mass* is a value that tells us the number of mass units of the part for each 100 mass units of the whole.
- For example, our bodies contain about 8.0% by mass of blood. The part is blood, and the whole is body.
- The 8.0% by mass of blood tells us that that for every 100 kilograms of body, there are 8.0 kilograms of blood, or for every 100 pounds of body, there are 8.0 pounds of blood.
- In chemistry, it is common to assume that any percentage is a percentage by mass unless you are specifically told otherwise.

# Conversion Factors from Percentages and Percentage Calculations

- For the conversion factors derived from percentages, we can use any mass units we want in the ratio as long as the units are the same for the part and for the whole.
- 8.0% by mass of blood in our bodies leads to the following:

$$\frac{8.0 \text{ kg blood}}{100 \text{ kg body}} \quad \text{or} \quad \frac{8.0 \text{ g blood}}{100 \text{ g body}} \quad \text{or} \quad \frac{8.0 \text{ lb blood}}{100 \text{ lb body}} \quad \text{or} \quad \dots$$

# Conversion Factors from Percentage by Volume



- *Percentage by volume* is a value that tells us the number of volume units of the part for each 100 volume units of the whole.
- Thus, when we are told that 3.2% by volume of our blood goes to our brain, this means that for every 100 liters of blood total, 3.2 liters of blood goes to the brain and that for every 100 milliliters of blood total, 3.2 milliliters of blood goes to the brain.

# Conversion Factors from Percentage by Volume

- For the conversion factors derived from percentages by volume, we can use any volume units we want in the ratio as long as the units are the same for the part and for the whole.
- 3.2% blood to our brain leads to

$$\frac{3.2 \text{ L blood to brain}}{100 \text{ L blood total}} \quad \text{or} \quad \frac{3.2 \text{ mL blood to brain}}{100 \text{ mL blood total}} \quad \text{or}$$
$$\frac{3.2 \text{ qt blood to brain}}{100 \text{ qt blood total}} \quad \text{or} \quad \dots$$

# Conversion Factors from Percentages and Percentage Calculations

- Mass percentages and volume percentage can be used as unit analysis conversion factors to convert between units of the part and units of the whole.

$$\text{For } X\% \text{ by mass} \quad \frac{X \text{ (any mass unit) part}}{100 \text{ (same mass unit) whole}}$$

$$\text{For } X\% \text{ by volume} \quad \frac{X \text{ (any volume unit) part}}{100 \text{ (same volume unit) whole}}$$

Example 2.11: Your body is about 8.0% blood. If you weigh 145 pounds, what is the mass of your blood in kilograms?

- It's a good strategy to write down the values you are given, separating them from the words. We are given two things, 8.0% blood and 145 lb.
  - It's important to recognize that the percentage provides a ratio of two units. Because both *kg* and *lb* are mentioned in the problem, we could use either of the following conversion factors.

$$\frac{8.0 \text{ lb blood}}{100 \text{ lb body}} \quad \text{or} \quad \frac{8.0 \text{ kg blood}}{100 \text{ kg body}}$$

- Because we see from the ratios above that two different things are mentioned, blood and body, it is important to identify the 145 lb as **145 lb body**.

Example 2.11: Your body is about 8.0% blood. If you weigh 145 pounds, what is the mass of your blood in kilograms?

- We want kg blood, which has a single unit, so we set it equal to the value given with a single unit. (Remember that the percentage hides a ratio of two units.)
- Next, we write the skeleton of the first conversion factor.

$$? \text{ kg blood} = 145 \text{ lb body} \left( \frac{\quad}{\text{lb body}} \right)$$



Example 2.11: Your body is about 8.0% blood. If you weigh 145 pounds, what is the mass of your blood in kilograms?

- We can either use the first conversion factor below to convert from pounds of body to pounds of blood or we can first convert pounds to kilograms and then use the second conversion factor below to convert to kilograms of blood. .

$$\frac{8.0 \text{ lb blood}}{100 \text{ lb body}} \quad \text{or} \quad \frac{8.0 \text{ kg blood}}{100 \text{ kg body}}$$

$$? \text{ kg blood} = 145 \text{ lb body} \left( \frac{8.0 \text{ lb blood}}{100 \text{ lb body}} \right) \left( \frac{453.6 \text{ g}}{1 \text{ lb}} \right) \left( \frac{1 \text{ kg}}{10^3 \text{ g}} \right) = 5.3 \text{ kg blood}$$

$$? \text{ kg blood} = 145 \text{ lb body} \left( \frac{453.6 \text{ g}}{1 \text{ lb}} \right) \left( \frac{1 \text{ kg}}{10^3 \text{ g}} \right) \left( \frac{8.0 \text{ kg blood}}{100 \text{ kg body}} \right) = 5.3 \text{ kg blood}$$

# Steps for Calculations Using Unit Analysis



- **Step 1** State your question in an expression that sets the unknown unit(s) equal to one or more of the values given.
  - To the left of the equals sign, show the unit(s) you want in your answer.

# Steps for Calculations Using Unit Analysis



- **Step 1 (cont.)**
  - To the right of the equals sign, start with an expression composed of the given unit(s) that parallels in kind and placement the units you want in your answer.
    - If you want a single unit in your answer, start with a value that has a single unit.
    - If you want a ratio of two units in your answer, start with a value that has a ratio of two units, or start with a ratio of two values, each of which has one unit.
    - Put each type of unit in the position you want it to have in the answer.

# Steps for Calculations Using Unit Analysis



- **Step 2** Multiply the expression to the right of the equals sign by conversion factors that cancel unwanted units and generate the desired units.
  - If you are not certain which conversion factor to use, ask yourself, “What is the fundamental conversion the problem requires, and what conversion factor do I need to make that type of conversion?” Figure 2.3 (atoms-first) provides a guide to useful conversion factors.

# Conversion Types

Metric unit	$\frac{\text{Conversion factors from prefixes on Table 1.2}}{\text{(same type of measurement)}}$	Another metric unit
English length unit	$\frac{2.54 \text{ cm}}{1 \text{ in.}}$ or conversion factors in Table 2.1 & 8.1	Metric length unit
English mass unit	$\frac{453.6 \text{ g}}{1 \text{ lb}}$ or conversion factors in Table 2.1 & 8.1	Metric mass unit
English volume unit	$\frac{3.785 \text{ L}}{1 \text{ gal}}$ or conversion factors in Table 2.1 & 8.1	Metric volume unit
English unit	$\frac{\text{Conversion factors in Table A.5 Appendix A}}{\text{(same type of measurement)}}$	Another English unit
Mass	$\frac{\text{Density as a conversion factor}}{\text{Density as a conversion factor}}$	Volume
Mass of a part	$\text{For X\% by mass } \frac{X \text{ (any mass unit) part}}{100 \text{ (same mass unit) whole}}$	Mass of a whole
Volume of a part	$\text{For X\% by volume } \frac{X \text{ (any volume unit) part}}{100 \text{ (same volume unit) whole}}$	Volume of a whole

# Steps for Calculations Using Unit Analysis



- **Step 3** Do a quick check to be sure you used correct conversion factors and that your units cancel to yield the desired unit(s).
- **Step 4** Do the calculation, rounding your answer to the correct number of significant figures and combining it with the correct unit.

# Example 2.12: Convert 4567.36 micrograms to kilograms.

- First, we set the unit that we want equation to the value given.

$$? \text{ kg} = 4567.36 \text{ } \mu\text{g}$$

- Next, we set up the skeleton of the next conversion factor.

$$? \text{ kg} = 4567.36 \text{ } \cancel{\mu\text{g}} \left( \frac{\quad}{\cancel{\mu\text{g}}} \right)$$

- When converting from one SI unit to another, it is most reliable to convert from the unit you have to the base unit and then from the base unit to the unit you want.

$$? \text{ kg} = 4567.36 \text{ } \cancel{\mu\text{g}} \left( \frac{\cancel{\text{g}}}{\cancel{\mu\text{g}}} \right) \left( \frac{\text{kg}}{\cancel{\text{g}}} \right)$$

# Example 2.12: Convert 4567.36 micrograms to kilograms.

$$? \text{ kg} = 4567.36 \cancel{\mu\text{g}} \left( \frac{1 \text{ g}}{10^6 \cancel{\mu\text{g}}} \right)$$

$$? \text{ kg} = 4567.36 \cancel{\mu\text{g}} \left( \frac{1 \cancel{\text{g}}}{10^6 \cancel{\mu\text{g}}} \right) \left( \frac{\quad}{\cancel{\text{g}}} \right)$$

$$? \text{ kg} = 4567.36 \cancel{\mu\text{g}} \left( \frac{1 \cancel{\text{g}}}{10^6 \cancel{\mu\text{g}}} \right) \left( \frac{1 \text{ kg}}{10^3 \cancel{\text{g}}} \right)$$

$$? \text{ kg} = 4567.36 \cancel{\mu\text{g}} \left( \frac{1 \cancel{\text{g}}}{10^6 \cancel{\mu\text{g}}} \right) \left( \frac{1 \text{ kg}}{10^3 \cancel{\text{g}}} \right) = 4.56736 \times 10^{-6} \text{ kg}$$



# Example 2.12: Convert 4567.36 micrograms to kilograms.

Desired unit

Converts given metric unit to metric base unit

$$? \text{ kg} = 4567.36 \text{ } \cancel{\mu\text{g}} \left( \frac{1 \cancel{\text{g}}}{10^6 \cancel{\mu\text{g}}} \right) \left( \frac{1 \text{ kg}}{10^3 \cancel{\text{g}}} \right) = 4.56736 \times 10^{-6} \text{ kg}$$

Given value

Converts metric base unit to desired metric unit

# Example 2.13: Convert 475 miles to kilometers.

$$? \text{ km} = 475 \text{ mi}$$

$$? \text{ km} = 475 \text{ mi} \left( \frac{\quad}{\text{mi}} \right)$$

- We are converting from English length to SI length, and 2.54 cm per inch is a good conversion factor to use for this because it is exact.
- Our problem becomes a three-step problem.

$$\text{mi} \rightarrow \text{in.} \rightarrow \text{cm} \rightarrow \text{km}$$

$$? \text{ km} = 475 \text{ mi} \left( \frac{5280 \text{ ft}}{1 \text{ mi}} \right)$$

# Example 2.13: Convert 475 miles to kilometers.

$$? \text{ km} = 475 \text{ mi} \left( \frac{5280 \text{ ft}}{1 \text{ mi}} \right) \left( \frac{\quad}{\text{ft}} \right)$$

$$? \text{ km} = 475 \text{ mi} \left( \frac{5280 \text{ ft}}{1 \text{ mi}} \right) \left( \frac{12 \text{ in.}}{1 \text{ ft}} \right)$$

$$? \text{ km} = 475 \text{ mi} \left( \frac{5280 \text{ ft}}{1 \text{ mi}} \right) \left( \frac{12 \text{ in.}}{1 \text{ ft}} \right) \left( \frac{\quad}{\text{in.}} \right)$$

$$? \text{ km} = 475 \text{ mi} \left( \frac{5280 \text{ ft}}{1 \text{ mi}} \right) \left( \frac{12 \text{ in.}}{1 \text{ ft}} \right) \left( \frac{2.54 \text{ cm}}{1 \text{ in.}} \right)$$

$$? \text{ km} = 475 \text{ mi} \left( \frac{5280 \text{ ft}}{1 \text{ mi}} \right) \left( \frac{12 \text{ in.}}{1 \text{ ft}} \right) \left( \frac{2.54 \text{ cm}}{1 \text{ in.}} \right) \left( \frac{\quad}{\text{cm}} \right)$$

# Example 2.13: Convert 475 miles to kilometers.

$$? \text{ km} = 475 \text{ mi} \left( \frac{5280 \text{ ft}}{1 \text{ mi}} \right) \left( \frac{12 \text{ in.}}{1 \text{ ft}} \right) \left( \frac{2.54 \text{ cm}}{1 \text{ in.}} \right) \left( \frac{1 \text{ m}}{10^2 \text{ cm}} \right)$$

$$? \text{ km} = 475 \text{ mi} \left( \frac{5280 \text{ ft}}{1 \text{ mi}} \right) \left( \frac{12 \text{ in.}}{1 \text{ ft}} \right) \left( \frac{2.54 \text{ cm}}{1 \text{ in.}} \right) \left( \frac{1 \text{ m}}{10^2 \text{ cm}} \right) \left( \frac{\quad}{\text{m}} \right)$$

$$? \text{ km} = 475 \text{ mi} \left( \frac{5280 \text{ ft}}{1 \text{ mi}} \right) \left( \frac{12 \text{ in.}}{1 \text{ ft}} \right) \left( \frac{2.54 \text{ cm}}{1 \text{ in.}} \right) \left( \frac{1 \text{ m}}{10^2 \text{ cm}} \right) \left( \frac{1 \text{ km}}{10^3 \text{ m}} \right)$$

# Example 2.13: Convert 475 miles to kilometers.

- When you do the calculation, the calculator shows 764.4384.

$$? \text{ km} = 475 \text{ mi} \left( \frac{5280 \text{ ft}}{1 \text{ mi}} \right) \left( \frac{12 \text{ in.}}{1 \text{ ft}} \right) \left( \frac{2.54 \text{ cm}}{1 \text{ in.}} \right) \left( \frac{1 \text{ m}}{10^2 \text{ cm}} \right) \left( \frac{1 \text{ km}}{10^3 \text{ m}} \right) = 764 \text{ km}$$

English-English unit conversions      English-metric conversion factor      metric-metric unit conversions

- Memorizing other English-metric conversion factors will save you time and effort. For example, if you know that 1.609 km = 1 mi, the problem becomes much easier.

$$? \text{ km} = 475 \text{ mi} \left( \frac{1.609 \text{ km}}{1 \text{ mi}} \right) = 764 \text{ km}$$

# Example 2.14: What is the volume in liters of 64.567 pounds of ethanol at 20 °C?

$$? \text{ L} = 64.567 \text{ lb}$$

- Pound is a mass unit, and we want volume in liters. Density provides a conversion factor that converts between mass and volume. You can find the density of ethanol on a table such as Table 2.2. It is 0.7893 g/mL at 20 °C.
- Our problem becomes a three-step problem.

$$\text{lb} \rightarrow \text{g} \rightarrow \text{mL} \rightarrow \text{L}$$

$$? \text{ L} = 64.567 \cancel{\text{ lb}} \left( \frac{\quad}{\cancel{\text{ lb}}} \right)$$

$$? \text{ L} = 64.567 \cancel{\text{ lb}} \left( \frac{453.6 \text{ g}}{1 \cancel{\text{ lb}}} \right)$$

Example 2.14: What is the volume in liters of 64.567 pounds of ethanol at 20 °C?

$$? \text{ L} = 64.567 \cancel{\text{ lb}} \left( \frac{453.6 \cancel{\text{ g}}}{1 \cancel{\text{ lb}}} \right) \left( \frac{\quad}{\cancel{\text{ g}}} \right)$$

$$? \text{ L} = 64.567 \cancel{\text{ lb}} \left( \frac{453.6 \cancel{\text{ g}}}{1 \cancel{\text{ lb}}} \right) \left( \frac{1 \text{ mL}}{0.7893 \cancel{\text{ g}}} \right)$$

$$? \text{ L} = 64.567 \cancel{\text{ lb}} \left( \frac{453.6 \cancel{\text{ g}}}{1 \cancel{\text{ lb}}} \right) \left( \frac{1 \cancel{\text{ mL}}}{0.7893 \cancel{\text{ g}}} \right) \left( \frac{\quad}{\cancel{\text{ mL}}} \right)$$

$$? \text{ L} = 64.567 \cancel{\text{ lb}} \left( \frac{453.6 \cancel{\text{ g}}}{1 \cancel{\text{ lb}}} \right) \left( \frac{1 \cancel{\text{ mL}}}{0.7893 \cancel{\text{ g}}} \right) \left( \frac{1 \text{ L}}{10^3 \cancel{\text{ mL}}} \right)$$

$$? \text{ L} = 64.567 \cancel{\text{ lb}} \left( \frac{453.6 \cancel{\text{ g}}}{1 \cancel{\text{ lb}}} \right) \left( \frac{1 \cancel{\text{ mL}}}{0.7893 \cancel{\text{ g}}} \right) \left( \frac{1 \text{ L}}{10^3 \cancel{\text{ mL}}} \right) = 37.11 \text{ L}$$

# Conversion Factors from “Something per Something” Statements

- Anything that can be read as “something per something” can be used as a unit analysis conversion factor. For example, if a car is moving at 55 miles *per* hour, we could use the following conversion factor to convert from distance traveled in miles to time in hours or time in hours to distance traveled in miles.

$$\left( \frac{55 \text{ mi}}{1 \text{ hr}} \right)$$

- If you are building a fence, and plan to use four nails *per* board, the following conversion factor allows you to calculate the number of nails necessary to nail up 94 fence boards.

$$\left( \frac{4 \text{ nails}}{1 \text{ fence board}} \right)$$



Example 2.15: The label on a can of cat food tells you there are 0.94 lb of cat food per can with 0.15% calcium. If there are three servings per can, how many grams of calcium are in each serving?

- It is a good strategy write down the values you are given in the problem, and if things you are given can be written as ratios, it's a good idea to write them as conversion factors.
- Note that two phrases in this question can be read as “something per something” so they can be used as unit analysis conversion factors. The phrase “three servings per can” leads to the first conversion factor below, and “0.94 lb of cat food per can” leads to the second.
- Assuming that 0.15% is a mass percent, we can build the third conversion factor below.

$$\left( \frac{3 \text{ serv.}}{1 \text{ can}} \right) \quad \left( \frac{0.94 \text{ lb food}}{1 \text{ can}} \right) \quad \left( \frac{0.15 \text{ lb Ca}}{100 \text{ lb food}} \right)$$

Example 2.15: The label on a can of cat food tells you there are 0.94 lb of cat food per can with 0.15% calcium. If there are three servings per can, how many grams of calcium are in each serving?

$$? \text{ g Ca} = \text{One serv.}$$

$$? \text{ g Ca} = \text{One ~~serv.~~} \left( \frac{\quad}{\text{serv.}} \right)$$

$$? \text{ g Ca} = \text{One ~~serv.~~} \left( \frac{1 \text{ can}}{3 \text{ serv.}} \right)$$

$$? \text{ g Ca} = \text{One ~~serv.~~} \left( \frac{1 \text{ ~~can}serv.}~~} \right) \left( \frac{\quad}{\text{can}} \right)$$

$$? \text{ g Ca} = \text{One ~~serv.~~} \left( \frac{1 \text{ ~~can}serv.}~~} \right) \left( \frac{0.94 \text{ lb food}}{1 \text{ ~~can}~~} \right)$$

Example 2.15: The label on a can of cat food tells you there are 0.94 lb of cat food per can with 0.15% calcium. If there are three servings per can, how many grams of calcium are in each serving?

$$? \text{ g Ca} = \text{One serv.} \left( \frac{1 \text{ can}}{3 \text{ serv.}} \right) \left( \frac{0.94 \text{ lb food}}{1 \text{ can}} \right) \left( \frac{\quad}{\text{lb food}} \right)$$

$$? \text{ g Ca} = \text{One serv.} \left( \frac{1 \text{ can}}{3 \text{ serv.}} \right) \left( \frac{0.94 \text{ lb food}}{1 \text{ can}} \right) \left( \frac{0.15 \text{ lb Ca}}{100 \text{ lb food}} \right)$$

$$? \text{ g Ca} = \text{One serv.} \left( \frac{1 \text{ can}}{3 \text{ serv.}} \right) \left( \frac{0.94 \text{ lb food}}{1 \text{ can}} \right) \left( \frac{0.15 \text{ lb Ca}}{100 \text{ lb food}} \right) \left( \frac{\quad}{\text{lb}} \right)$$

$$? \text{ g Ca} = \text{One serv.} \left( \frac{1 \text{ can}}{3 \text{ serv.}} \right) \left( \frac{0.94 \text{ lb food}}{1 \text{ can}} \right) \left( \frac{0.15 \text{ lb Ca}}{100 \text{ lb food}} \right) \left( \frac{453.6 \text{ g}}{1 \text{ lb}} \right)$$

$$? \text{ g Ca} = \text{One serv.} \left( \frac{1 \text{ can}}{3 \text{ serv.}} \right) \left( \frac{0.94 \text{ lb food}}{1 \text{ can}} \right) \left( \frac{0.15 \text{ lb Ca}}{100 \text{ lb food}} \right) \left( \frac{453.6 \text{ g}}{1 \text{ lb}} \right) = 0.21 \text{ g Ca}$$

Example 2.16: When 2.3942 kg of the sugar glucose are burned (combusted), 37,230 kJ of heat are evolved. What is the heat of combustion of glucose in J/g?

- When the answer you want is a ratio of two units, make it clear which unit you want on the top and which unit you want on the bottom

$$\frac{? \text{ J}}{\text{g glucose}} =$$

- If you want two units, start your unit analysis setup with a ratio of two units. Put the correct type of unit in the correct position in the ratio. For this problem, we put the heat unit on the top and the mass unit on the bottom. (Heat evolved is described with a negative sign.)

$$\frac{? \text{ J}}{\text{g glucose}} = \frac{-37,230 \text{ kJ}}{2.3942 \text{ kg glucose}}$$

Example 2.16: When 2.3942 kg of the sugar glucose are burned (combusted), 37,230 kJ of heat are evolved. What is the heat of combustion of glucose in J/g?

$$\frac{? \text{ J}}{\text{g glucose}} = \frac{-37,230 \text{ kJ}}{2.3942 \text{ kg glucose}} \left( \frac{\text{---}}{\text{kJ}} \right)$$

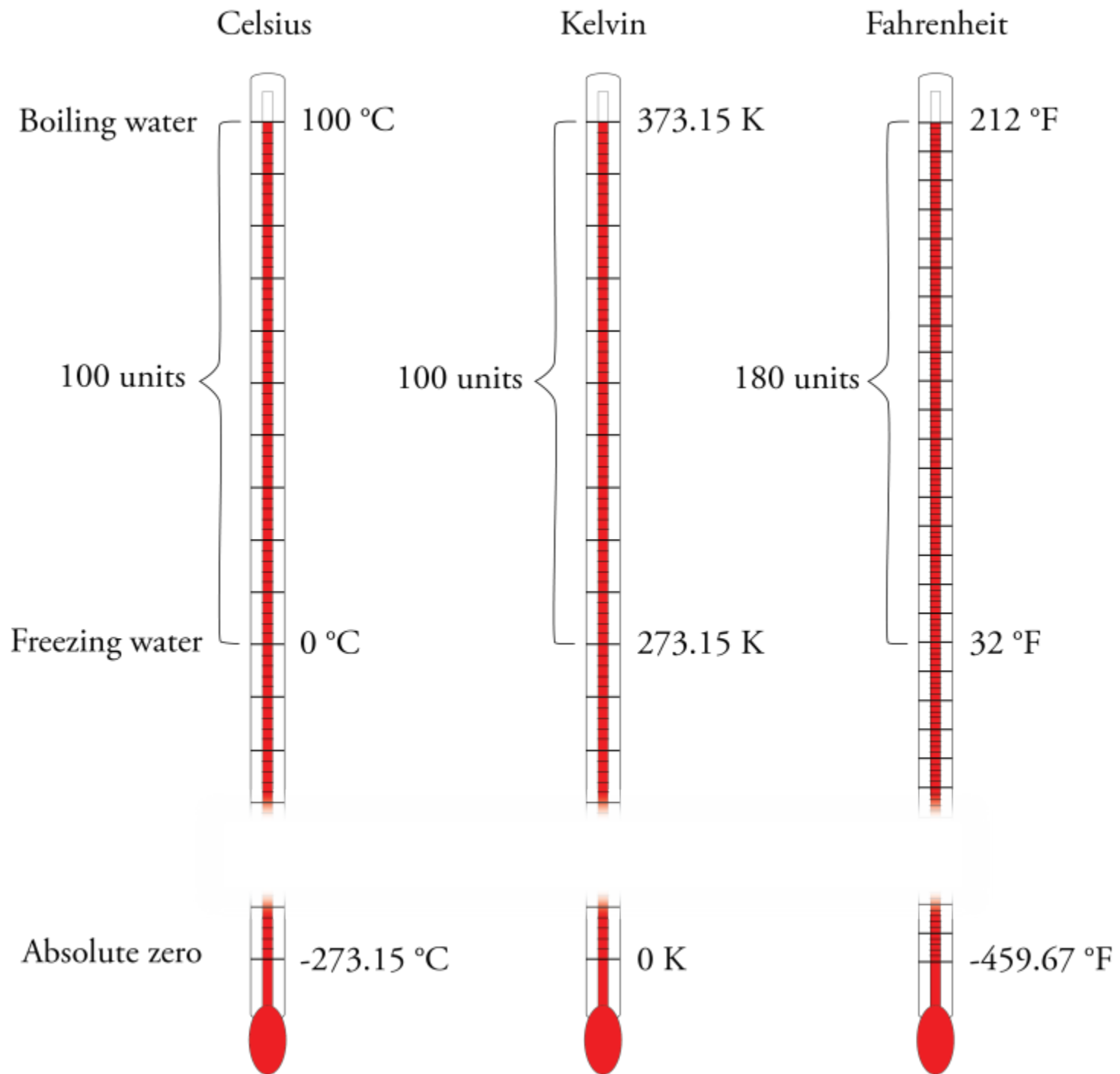
$$\frac{? \text{ J}}{\text{g glucose}} = \frac{-37,230 \text{ kJ}}{2.3942 \text{ kg glucose}} \left( \frac{10^3 \text{ J}}{1 \text{ kJ}} \right)$$

$$\frac{? \text{ J}}{\text{g glucose}} = \frac{-37,230 \text{ kJ}}{2.3942 \text{ kg glucose}} \left( \frac{10^3 \text{ J}}{1 \text{ kJ}} \right) \left( \frac{\text{---}}{\text{kg}} \right)$$

$$\frac{? \text{ J}}{\text{g glucose}} = \frac{-37,230 \text{ kJ}}{2.3942 \text{ kg glucose}} \left( \frac{10^3 \text{ J}}{1 \text{ kJ}} \right) \left( \frac{1 \text{ kg}}{10^3 \text{ g}} \right)$$

$$\frac{? \text{ J}}{\text{g glucose}} = \frac{-37,230 \text{ kJ}}{2.3942 \text{ kg glucose}} \left( \frac{10^3 \text{ J}}{1 \text{ kJ}} \right) \left( \frac{1 \text{ kg}}{10^3 \text{ g}} \right) = -15,550 \text{ J/g}$$

# Temperature Scales



# Temperature Conversions

- We will use the following equations to convert a temperature reported in one of these systems to the equivalent temperature in another.

$$? \text{ } ^\circ\text{F} = \text{number of } ^\circ\text{C} \left( \frac{1.8 \text{ } ^\circ\text{F}}{1 \text{ } ^\circ\text{C}} \right) + 32 \text{ } ^\circ\text{F}$$

$$? \text{ } ^\circ\text{C} = (\text{number of } ^\circ\text{F} - 32 \text{ } ^\circ\text{F}) \left( \frac{1 \text{ } ^\circ\text{C}}{1.8 \text{ } ^\circ\text{F}} \right)$$

$$? \text{ K} = \text{number of } ^\circ\text{C} + 273.15$$

$$? \text{ } ^\circ\text{C} = \text{number of K} - 273.15$$

- Note that the numbers 1.8, 32, and 273.15 in these equations all come from definitions, so they are all exact.

Example 2.17: “Heavy” water contains the heavy form of hydrogen called deuterium, whose atoms each have one proton, one neutron, and one electron. Heavy water freezes at 38.9 °F. What is this temperature in °C?

- We use the equation for converting Fahrenheit temperatures to Celsius:

$$? \text{ } ^\circ\text{C} = (38.9 \text{ } ^\circ\text{F} - 32 \text{ } ^\circ\text{F}) \left( \frac{1 \text{ } ^\circ\text{C}}{1.8 \text{ } ^\circ\text{F}} \right)$$

- Rounding off the answer can be tricky here. When you subtract 32 from 38.9, you get 6.9. The 32 is exact, so it is ignored when considering how to round off the answer. The 38.9 is precise to the first number after the decimal point, so the answer to the subtraction is reported to the tenths place. There are two significant figures in 6.9, so when we divide by the exact value of 1.8 °F, we round our answer to two significant figures.

$$? \text{ } ^\circ\text{C} = (38.9 \text{ } ^\circ\text{F} - 32 \text{ } ^\circ\text{F}) \left( \frac{1 \text{ } ^\circ\text{C}}{1.8 \text{ } ^\circ\text{F}} \right) = (6.9 \text{ } ^\circ\text{F}) \left( \frac{1 \text{ } ^\circ\text{C}}{1.8 \text{ } ^\circ\text{F}} \right) = 3.8 \text{ } ^\circ\text{C}$$



Example 2.18: The compound 1-chloropropane,  $\text{CH}_3\text{CH}_2\text{CH}_2\text{Cl}$ , melts at  $46.6\text{ }^\circ\text{C}$ . What is this temperature in  $^\circ\text{F}$ ?


- We use the equation for converting Celsius temperatures to Fahrenheit:

$$? \text{ }^\circ\text{F} = 46.6 \text{ }^\circ\text{C} \left( \frac{1.8 \text{ }^\circ\text{F}}{1 \text{ }^\circ\text{C}} \right) + 32 \text{ }^\circ\text{F}$$

Example 2.18: The compound 1-chloropropane,  $\text{CH}_3\text{CH}_2\text{CH}_2\text{Cl}$ , melts at  $46.6\text{ }^\circ\text{C}$ . What is this temperature in  $^\circ\text{F}$ ?

$$? \text{ }^\circ\text{F} = 46.6 \text{ }^\circ\text{C} \left( \frac{1.8 \text{ }^\circ\text{F}}{1 \text{ }^\circ\text{C}} \right) + 32 \text{ }^\circ\text{F} = 83.9 \text{ }^\circ\text{F} + 32 \text{ }^\circ\text{F} = 115.9 \text{ }^\circ\text{F}$$

- Because the calculation involves multiplication and division as well as addition, you need to apply two different rules for rounding off your answer.
- When you multiply 46.6, which has three significant figures, by the exact value of  $1.8\text{ }^\circ\text{F}$ , your answer should have three significant figures.
- The answer on the display of the calculator, 83.88, would therefore be rounded off to 83.9.
- You then add the exact value of  $32\text{ }^\circ\text{F}$  and report the answer to the tenths place.



Example 2.19:  
Silver melts at 961 °C.  
What is this temperature in K?

$$? \text{ K} = 961 \text{ }^\circ\text{C} + 273.15 = 1234 \text{ K}$$

- For rounding off our answer, we assumed that 961 °C came from a measurement, so it is not exact. It is precise to the unit position.
- On the other hand, 273.15 is exact, and has no effect on the uncertainty of our answer.
- We therefore report the answer for our addition to the unit position, rounding off 1234.15 to 1234.

Example 2.20:

Tin(II) sulfide, SnS, melts at 1155 K.

What is this temperature in °C?

$$? \text{ } ^\circ\text{C} = 1155 \text{ K} - 273.15 = 882 \text{ } ^\circ\text{C}$$

- Because 1155 is precise to the ones place, and 273.15 is exact, we report the answer for our subtraction to the ones place.