Chapter 2. Measurements in Chemistry

2.1 Measurement Systems

Measuring system is an integral part of any commerce or science for people to agree on the quantities of in their business and practices. In England a measuring system for the trading volumes was not properly standardized until the end of 13th century. For example, until gallon was made the standardized measure of volume consolidating different types of gallons (ale, wine and corn) were widely used. Currently, commonly accepted **US retail market unit system** uses foot, pound, gallons and hrs while **European Union (EU)** uses meters, kg, liters and hrs.

In France and in 1799 for sciences, a decimal system called metric system using centimeter, gram, and second (**CGS system**) was created for length, mass and time, respectively. In **CGS system**, units are compared to a standardized measure: meter was defined as being one ten-millionth part of a quarter of the earth's circumference and 100 centimeters equals a meter.

Prefixes used in abbreviating measurements

Multiplication Factor	Prefix	Symbol
1,000,000,000 = 10^9	giga	G
1,000,000 = 10^6	mega	M
1,000 = 10^3	kilo	k
100 = 10^2	hecto	h
1 = 1		
0.01 = 10^{-2}	centi	c
0.001 = 10^{-3}	milli	m
0.000001 = 10^{-6}	micro	µ
0.00000001 = 10^{-9}	nano	n

A unit system also uses **prefix** allowing the numbers with many zeros before and after the decimal place to be simplified.

For example, converting measurements to a unit that replaces the power of ten by a prefix:

Small Numbers:

0.00000000**680** m 6.80 x **10^{-9}** m (**6.08 nm**)	0.00000**714** s 7.14 x **10^{-6}** s (**7.14 µm**)	0.00**288** 2.88 x **10^{-3}** g (**2.88 mg**)

Large Numbers:

4560 g 4.56 x 10^3 g (**6.08 kg**)	7140000 s 7.14 x **10^6** s (**7.14 Ms**)	**360**0000000000 bites 3.60 x **10^{12}** bites (b) (**3.60 Gb**)

2.2 Metric System Units-SI system

Later a comprehensive Le Systeme international d'Unites (**SI unit system**) was created in **1960** and has been officially adopted by nearly all countries for mainly sciences. SI units are based prefixes and upon **7 principal units**, 1 in each of 7 different categories measurements.

Category	Name	Abbreviation
Length	metre	**m**
Mass	kilogram	**kg**
Time	second	**s**
Electric current	ampere	**A**
Temperature	Kelvin	**K**
Amount of	mole	**Mol**
Luminous intensity	candela	**cd**

2.3 Exact and Inexact Numbers

Exact Measurements: For certain types of number associated with **conversion factors** are considered "**exact**." For example, there are exactly 16 ounces in one pound. The number 16 would have as many decimal places, or significant figures (discussed later) as needed. So one pound has 16.000000000000.... ounces. If you see that any numbers you use in a calculation comes from a definition or conversion factor you could safely assume they are **exact numbers**.

Inexact Measurements: In most of measurements the number associated with them are considered "**inexact**." For example, if you measure the mass of a certain object on various balances and found that it could have masses: 25.0125 g, 25.013, 25.01 g 25.0 g and 25 g. Notice that the rightmost decimal place is rounded off to a place at right depending on the decreasing accuracy (quality or how good they are) of the balance used. If you use a number in a calculation which comes from an experimental measurement you could safely assume they are **inexact numbers**.

Which types of numbers are considered "exact?" Below are the general rules.

1. **Metric to metric system** conversion factors are **exact** e.g. **1** m = **100** cm or **1** m/**100** cm (In this conversion, 1 and 100 are both **exact.)**

Conversions **between English and Metric system** are **generally _NOT_ exact**. Exceptions will be pointed out to you.

2. e.g. **1** in = **2.54** cm exactly (**1** and **2.54** are both **exact**.)
3. e.g. **454** g = **1** lb or **454** g/1 lb (**454** has 3 sig. fig., but **1** is exact.)

2.4 Uncertainty in Measurement and Significant Figures

An **inexact numbers** always comes out of an experimental measurement. They have been rounded off to show the uncertainty of the measurement. For example, if you measure the mass of a certain object on a balance and found that it gives a mass of **25.0125** g, **25.013**, **25.01** g **25.0** g or **25** g with a decimal place rounded off to a place at right depending on the decreasing accuracy of the balances used. A measurement is always written down after considering the instrumental uncertainties. The uncertainly of a measurement could be conveniently expressed as a **significant figure (SF)**. The **right most digit** in 25.012**5** g which is **5** at the 4th decimal place is considered the **uncertain digit**.

Significant Figure (SF).

The significant figure for an inexact number is obtained by counting the other digits to the left from the uncertain digit. Therefore, 25.012**5** g would have 6 significant figures **(6 SF)**: 25.01**3** **(5 SF)** , 25.0**1** g **(4 SF)** 25.**0** **(3 SF)** g and 2**5** g **(2 SF)**. Lower the SF, higher the uncertainty of the measurement and less the accuracy of the instrument.

GENERAL RULES FOR FIGURING WHICH NUMBERS are significant

1. **ZEROS** used to "place" the decimal are **NOT significant figures**:
   0.015 g = 2 SF
2. **LEADING ZEROS BEFORE all the digits** are **NOT significant**:
   000340 = 3 SF and 0.000216 g = 3 SF
3. **TRAILING ZEROS after all the digits in the decimal places** are **SIGNIFICANT**:
   1.5**00** g= 4 SF
   TRAILING ZEROS in whole digits may or may not be **SIGNIFICANT**
   46**00.** g= 4 SF 4600 g= 2 SF or written like this is the best 4.6 x 10^3
4. **SANDWICH ZERO** WITHIN a number are **SIGNIFICANT**:
   0.01**0**5 g = 3 SF and 1**0**.5 g = 3 SF
   0.027 g = **2 SF** (LEADING zeros BEFORE all the digits are NOT significant)
   2.600 m = **4 SF** (TRAILING zeros after all the digits are SIGNIFICANT)
   21**0.0**5 s = **5 SF** (SANDWICH zeros WITHIN a number are SIGNIFICANT)
   0.03**0**6 Kcal = **3 SF** (LEADING zeros BEFORE as well as SANDWICH zeros WITHIN are SIGNIFICANT)

How many significant figures are in the following numbers?
a) 0.0945 **(3 SF)** b) 83.22 **(4 SF)** c) 106 **(3 SF)** d) 0.000130 **(3 SF)**

Deduce the number of significant figures contained in the following:
a) 16.0 cm **(3 SF)** b) 0.0063 m **(2 SF)** c) 100 km **(3 SF)**
d) 2.9374 g **(5 SF)** e) 1.07 lb/in^2 **(3SF)**

How many significant figures are in the following measurements?
a) 25.9000g **(6 SF)** b) 102 cm **(3 SF)** c) 0.002 m **(1 SF)** d) 2001 kg **(5 SF)** e) 0.0605 s **(3 SF)** f) 21.2 m **(3 SF)** g) 0.023 kg **(2 SF)**
h) 46.94 cm **(4 SF)** i) 453.59 g **(5 SF)** j) 1.6030 km **(5 SF)**

Uncertainty, Error, Accuracy, and Precision of measurements

Uncertainty is expressed in terms of the rounding off to a significant figure.
e.g. 25.013 **(5 SF)** , 25.01 g **(4 SF)** 25.0 **(3 SF)** g and 25 g**(2 SF).** Note that greater the number of significant figures, the greater the precision.

Precision versus Accuracy:
Precision = How close series measurements agree: If you take more reading of the same measurement how close they are.
e.g. 25.0125 g, 25.0124 g and 25.0126 g are more precise than 25.0225 g, 25.10127 g and 25.0326.

Accuracy = how close measurement is to the true value: something could be precise but inaccurate if the value is off by a calibration error.

e.g. Measurements 25.0125 g, 25.0124 g and 25.0126 and the true value 25.0125 g are **both precise and accurate.**

e.g. Measurements 25.0125 g, 25.0124 g and 25.0126 and the true value 25.0125 g are **both precise and accurate.**

e.g. 25.0225 g, 25.10127 g and 25.0326 g and the true value 26.0125 g are **neither precise nor accurate.**

2.5 Significant Figures and Mathematical Operations

Most of the experiments involve calculation of a answer (derived quantity) from basic measurements with various units to a complex quantity with derived units. A simple example would be calculation of velocity of an object (car) travelling certain distance 356.5 miles in a given time in 4.8 hours. Question is how we obtain a velocity with correct uncertainties (SF) corresponding to uncertainties of distance and the time. Some of derived quantities involve additions/subtractions and/or multiplication/division.

General rules for significant figure in addition/subtraction:

1. When **adding or subtracting numbers**, all numbers must have the same units.
2. The answer can have no more decimals than the measurement with the **fewest DECIMALS**.

254 mL **- 54.1 mL** ~~**208.1 mL (4 SF)**~~ **208 mL (3 SF)**	**125.4 g** **2.54 g** ~~**127.94 g**~~ **127.9 g** **(4 SF)**	Rounding Off Numbers 1. **"extra" digit is LESS than 5-drop it.** 2. **"extra" digit is MORE than 5-ADD 1.** 3. **"extra" digit is 5 "Odd rule"** **e.g. 2.535 is rounded as "2.54** 4. **"extra" digit is 5 "Even rule"** **e.g. 2.525 is rounded as "2.52"**

Significant figures in multiplication/division calculations

1. When multiplications or divisions of numbers, all numbers must have the same units.
2. The answer can have no more significant figures than the measurement with the **fewest SIGNIFICANT FIGURES.**

(**231.54** * **43**)/**433.4** = 22.972358 (231.54 ~~**(5 SF)**~~ 43**(2 SF)** 433.4 ~~**(4 SF)**~~
= **22**.972358 = 21 **(2 SF)**
= 21 **(2 SF)**

Calculate **3.21** cm x 15.091 cm =**18.3**01 cm = **Ans. 18.3 cm (3 SF)**

Calculate 3.82 x **1.1** x 2.003 = **8.4**16606 = **Ans. 8.4 (2 SF)**

Calculate 13.87 ÷ **1.23** = **11.2**7642276 = **Ans. 11.3 (3 SF)**

Calculate 0.0**95** ÷ 1.427 = 0.0**66**573231 = **Ans. 0.067 (2 SF)**

In a long calculation involving mixed operations, carry as many digits as possible through the entire set of calculations and then round the final result appropriately. For example,

(**5.00** / **1.235**) + **3.000** + (**6.35** / **4.0**)

=**4.04**858... + **3.000** + **1.5**875=**8.6**30829... 5=**8.6**

2.6 Scientific Notation

Scientific notation uses power-of-10 to express an extremely large or small numbers. Scientific notation has a **regular number** with correct significant figure with a value between 1 to 10, and a **power** of 10 by which the regular number is multiplied. E.g. 0.0**67** is converted to scientific notation: **6.7** x 10^{-2}

The table shows several examples of numbers written in standard decimal notation (left-hand column) and in scientific notation (right-hand column).

Number in decimal form with significant digits color	**Scientific notation with correct significant digits**
1,222,000.00	**1.222** x 10^{6}
34,500.00	**3.450** x 10^{4}
0.0000**3450**000	**3.45** x 10^{-5}
-0.0000000**165**	-**1.65** x 10^{-8}

Scientific notation makes it easy to multiply and divide gigantic and/or minuscule numbers. To obtain the product of these two numbers (the coefficients) are multiplied, and the powers of 10 are added. This produces the following result:

2.56 x 10^{67} x -**8.333** x 10^{-54} = (**2.56** x -**8.333**)(10^{67} x10^{-54}) =
(-**21.3**3248)(10^{67-54}) = (-**21.3**)(10^{13})
= (-**2.13**)(10^{14}) = -**2.13** x 10^{14}

Now consider the quotient of the two numbers multiplied in the previous example:

(**3.46** x 10^{57}) / (**9.431** x 10^{-75})

To obtain the quotient, the coefficients are divided, and the powers of 10 are subtracted. This gives the following:

= ((**3.46** / (**9.431**)) x (10^{57}/$10^{-75)}$) = ((**3.46** / (**9.431**)) x (10^{57}x$10^{-(-75)}$)
= (**3.46** /**9.431**)) x 10^{57+75} = (**3.46** /**9.431**)) x 10^{57+75}

= 0.**366**875199 x 10^{137} = (**3.66**875199 x 10^{1}) x 10^{137}
= (**3.67**) x ((10^{1}) x 10^{137}) = (**3.67** x (10^{138}) = **3.67** x 10^{138}

2.7 Conversion Factors and Dimensional Analysis

Conversion Factors are used to convert a measurement to another with different units. They are used for the length, mass, area, volume, temperature, energy, force and time conversions as listed below:

Conversion Factors

Length	**Mass**	**Area**	**Volume**
1 ft =12 in 1 yd = 3 ft 1 mi 5280 ft = 1mi = 1.609km 1 in =2.54 cm 1 m =3.281 ft	1 lb = 16 oz 1 ton = 2000 lbs 1 lb = 453.59 g	1 acre =4.048x10^{3} m^{2} 1 acre =4840 yd^{2} 1 $mile^{2}$=2.589x10^{6} m^{2} 1 $mile^{2}$= 640 acres	1 gal = 4 qt 1 qt = 2 pt 1 gal = 3.785 L 1 L = 10^{3} mL 1 mL = 1 cm^{3}
Energy	**Pressure**	**Time**	**Temperature**
1 cal = 4.18681 J Btu=1.05506E3 J 1 food cal = 1 kcal	1 atm = 760 torr 1 atm = 1.01325E5 Pa 1 mmHg = 1 torr 1 mmHg = 1.333E2 Pa 1 Psi = 6.89476E3 Pa	60 min = 1 hr 24 hr = 1 day 365.25 days = 1 year	**?** K = (**x**) °C +273.15) **?** °C = (**x**) K - 273.15 **?** °C = (5/9) ((**x**)°F -32) **?** °F = (9/5)(**x**)°C +32

Simple unit conversations using factor label method

Dimensional Analysis (also called Factor-Label Method or the Unit Factor Method) is a problem-solving method that uses conversion factors to convert unit to get the answer with correct units.

- First write the measurement need to be converted.
- Select the conversion factors from conversion tables.
- Line up conversion factors so the units of the desired answers are obtained.
- Unit of bottom (denominator) must cancel when factor is multiplied by given number.

Length

How many meters are in a 4 cm?
Conversion factor: 1 cm = 10^{-2} m or

1 cm = 10^{-2} m or $\frac{1\ \text{cm}}{10^{-2}\ \text{m}}$ Align the conversion factor to a cancel cm

$$4\ \cancel{\text{cm}} \times \frac{10^{-2}\ \text{m}}{1\ \cancel{\text{cm}}} = 4 \times 10^{-2}\ \text{m}$$

How many inches are in 1 meter? Given the conversion factors 1 inch = 2.54 cm and 1 meter = 100 cm

$$\frac{1.00\ \cancel{\text{m}}}{} \Big| \frac{100\ \cancel{\text{cm}}}{1\ \cancel{\text{m}}} \Big| \frac{1\ \text{in}}{2.54\ \cancel{\text{cm}}} = 39.37008\ \text{m} = 39.3\ \text{m}$$

or

$$1.00\ \cancel{\text{m}} \times \frac{100\ \cancel{\text{cm}}}{1\ \cancel{\text{m}}} \times \frac{1\ \text{in}}{2.54\ \cancel{\text{cm}}} = 39.37008\ \text{m} = 39.3\ \text{m}$$

Note digits in **blue** are exact and was not considered for **SF**

Convert 2.4 meters to centimeters (Ans: 240 cm or 2.4 x 10^2 cm)

Mass

How many grams (g) in 150 pounds (lb) given the equalities 1 lb = 0.454 kg and 1 kg = 1000 g?

$$150\ \cancel{\text{lb}} \Big| \frac{0.454\ \cancel{\text{kg}}}{1\ \cancel{\text{lb}}} \Big| \frac{1000\ \text{g}}{1\ \cancel{\text{kg}}} = 68100\ \text{g} = 6.81 \times 10^4\ \text{g}$$

Scientific notation allows to show SF clearly.

Convert 65.5 centigrams to milligrams (Ans: 655 mg)

Convert 5 liters to cubic decimeters (Ans: 5 dm^3)

The density of a substance is 2.7 g/cm^3. **What is the density of the substance in kilograms per liter?** (Ans: 2.7 kg/L)

A car is traveling 65 miles per hour. How many feet does the car travel in one second? (Ans: **95 ft/sec**)

How many basketballs can be carried by 8 buses? Given 1 bus = 12 cars, 3 cars = 1 truck, 1000 basketballs = 1 truck (Ans: 32 000 basketball)

Area units:

How many cm^2 are in a m^2 (base unit) of area?
Square each **number and unit** in the conversion factor
1 cm = 10^{-2} m ; $(1\ cm)^2 = (10^{-2}\ m)^2$; $(1)^2\ cm^2 = (10^{-2})^2\ m^2$
$1\ cm^2 = 10^{-4}\ m^2$

Volume unit:
How many cm^3 are in m^3 (base unit) of volume?.
1 cm = 10^{-2} m
Cube each **number and unit** in the conversion factor
1 cm = 10^{-2} m; $(1\ cm)^3 = (10^{-2}\ m)^3$; $(1)^3\ cm^3 = (10^{-2})^3\ m^3$
$1\ cm^3 = 10^{-6}\ m^3$

How many m^3 are in μm^3 (base unit) of volume?
1 μm = 10^{-6} m (Ans: **$1\ \mu m^3 = 10^{-18}\ m^3$**)

How many m^3 (base unit) of volume are in μcm^3?
1 m = 10^{6} μm (Ans: **$1\ m^3 = 10^{18}\ \mu m^3$**)

Chemistry at a Glance: Practice unit conversion Factors using factor labeled method.

2.8 Density

Density is one of the physical characteristics of a substance that help to identify the substance. **Density** (**d**), is defined as **mass per unit volume**. Density is calculated by dividing the mass of an object by its volume. This is shown in equation form, as follows:

Density = mass ÷ volume

We can calculate the density of a solid, liquid, or gas. Note the difference in units in the formulas of the density of a solid & liquid to the gas. The **unit of mass** is grams, **g.** The **unit of volume** of **sold** is cubic centimeters is **cm^3**, **liquids** milliliters is **mL,** and for gases is liters **L or** cubic meters **m^3**.

Solids: d = grams (g) ÷ cubic centimeters (cm^3) = g/cm^3

Liquids: d = grams (g) ÷ milliliters (mL) = g/mL

Gases: d = grams (g) ÷ milliliters (L or m^3) = g/L or g/m^3

A student determines that a piece of an unknown solid material has a mass of 5.854 g and a volume of 7.57 cm^3. What is the density of the material, rounded to the correct number of significant digits?

Calculation using a formula:
First: Write the correct formula at the top of your page, and list the knowns and the unknowns.

M= 5.854 g, **V** = 7.57 cm^3
Second: Substitute the known values in the formula

$$D = \frac{M}{V} = \frac{M = \mathbf{5.854}\ g}{V = \mathbf{7.57}\ cm^3} = 0.77336\ \mathbf{g/cm^3} \qquad = \mathbf{7.73} \times 10^1\ \mathbf{g/cm^3}$$

Calculation using a factor label method:

5.854 g	1
	7.57 cm^3

= 0.77336 = **7.73** x 10^1 **g/cm³**

Aluminum block weighs 14.2 g and has a density of 2.70 g cm^{-3}. Calculate the volume of the block.

Calculation using a formula:
First: Write the correct formula at the top of your page, and list the knowns and the unknowns.

M= **14.2 g** , **2.70 g cm^{-3}**; or = **2.70 g/1 cm^{-3} ; V= ?**
Second: Substitute the known values in the problem

$$D = \frac{M}{V}\ ; \quad V = \frac{M}{D} = \frac{M = \mathbf{14.2}\ g}{D = \mathbf{2.70\ g/1\ cm^{-3}}} = \mathbf{5.25}93\ \mathbf{g/cm^3}$$

= **5.26** g/cm^3

Calculation using a factor label method:

14.2 ~~g~~	1 cm^{-3}
	2.70 ~~g~~

= 5.2593 **cm^3** **= 5.26**

$$\mathbf{D} = \frac{\mathbf{M}}{\mathbf{V}} = \ ?$$

The density of water is one gram per cubic centimeter. **What is the density of water in pounds per liter?** (Ans: **0.45 lb/L**)

2.9 Temperature Scales and Heat Energy

Temperature Scales

Astronomers and other scientists like to use a temperature scale called "**Kelvin**." On a Kelvin thermometer water freezes at **273 degrees** and boils at **373 degrees**. Zero degrees Kelvin is called "**absolute zero**." It is the lowest possible temperature of matter.

Temperature Conversions

? K = (**x**) °C + 273.15)
? °C = (**x**) K - 273.15
? °C = (5/9) ((**x**)°F -32)
? °F = (9/5)(**x**)°C +32

To convert degrees Celsius (°C) to Kelvin (K):

^{o}C -->K ; **?** K = (**x**) °C -273.15)

$K = C + 273.15$

Simply add 273. (Example, 0 (**x**) C° in K = **?=**)

? K = (**0**) °C +273.15) = **273.15** K

To convert Kelvin (K) to degrees Celsius (°C):

? °C = (**x**) K - 273.15

Simply subtract 273 degrees. (Example, 273 K = 0 deg. C)

To convert degrees Celsius (°C) to Fahrenheit (°F)

? °F = (9/5)(**x**)°C +32

Begin by multiplying the degrees Celsius (°C) temperature by 9, then divide the answer by 5. Finally add 32.

To convert degrees Fahrenheit (°F) to Celsius (°C):

? °C = (5/9) ((**x**)°F -32)

Begin by subtracting 32 from Fahrenheit (°F) temperature, then multiply by 5 and divide the answer by 9.

Human body temperature is 98.6 ^{o}F. Convert this temperature to

a) ^{o}C and b) K scale.

a) **?** °C = (5/9) ((**x**)°F -32) = (5/9) ((**98.6**)°F -32) = 5/9 (66.6) = **37.0** °C
b) **?** K = (**x**) °C + 273.15) ; **98.6** °F = **37.0** °C from a.
? K = (**37.0**) °C + 273.15) = **310**.2 K = **310**. K = **3.10** x 10^2 K

Describe heat energy and how they are measured in calories, dietary calories and joules.

Energy conversions: Thermal Energy Units

Calorie (cal)

The original definition of the **calorie** was the amount of thermal energy required to raise the temperature of a gram of water by one degree Celsius. This amount of energy turns out to be equal to **4.186 joules** (**J**).

1 **cal =** 4.186 **J** (**SI system units use joules**)

It is still a rather small amount of energy. The familiar "food **calorie**," used to measure chemical energy that our bodies can extract from food, is actually a kilocalorie (**kcal**)

1 **kcal** = 1000 **cal**,
1 **kcal** is 1000 calories—enough energy to raise the temperature of a kilogram of water by one degree Celsius. The thermal energy unit kcal is also know as "**food calories**" or simply a **calorie**.

British thermal unit (Btu), defined as the amount of thermal energy required to raise the temperature of one pound of water by one degree Fahrenheit.

1 **Btu** = 1055.06 J = 0.252 kcal = 1055.06 J
1 **Btu** = = 2.931×10^{-4} **kWh** (household electrical energy)

kcals in food

1 gram of each:
1 g carbohydrate = 4 calories (kcals)
1 g protein = 4 calories (kcals)
1 g fat = 9 calories (kcals)

How many kcals you need to be burned to lose 1 lb of body fat?
1 lb = 453.59 g

$$\frac{1.00\ \cancel{lb}}{} \left| \frac{453.59\ \cancel{g}}{1\ \cancel{lb}} \right| \frac{9\ \text{kacl}}{1\ \cancel{\text{g of fat}}} = 4982.31\text{kcal} = 4980\ \text{kcal}$$

An active adult needs as many as 2,500 to 3,000 calories per day.

Chemical Connections: Describe Body Density and Percent Body Fat and Normal Human Body Temperature.

Calories for Different Foods

Food	Calories	Protein (g)	Fat (g)	Carbohydrate (g)	Water (g)
Milk	65	3.3	4	5	87
Butter	740	-	82	-	15
Cream	210	2	21	3	72
Cheese	310	22	25	-	44
Ice Cream	170	4	7	25	64
Margarine	740	-	81	-	16
Eggs	150	12	11	-	75
Pork (Grilled)	340	29	24	-	36
Chicken (Roast)	150	25	5	-	55
Fish (eg. Cod)	220	20	10	8	60
Beans (Boiled)	20	2	-	3	90
Cabbage (Boiled)	10	1	-	1	96
Carrot (Boiled)	20	0.6	-	4	91
Cauliflower (Boiled)	10	1.5	-	1	93
Cucumber (Raw)	10	0.6	-	2	96
Peas (Boiled)	50	5	-	8	80
Potatoes (Boiled)	80	1	-	22	77
Tomatoes	15	1	-	3	93
Apples	45	0.3	-	12	84
Bananas	80	1	-	20	70

Calories for Different Activities

Activity	**Calories (125-174 lbs)Wt.**	**Activity**	**Calories (125-174 lbs)Wt.**
Sleeping	10	Auto Repair	35
Sitting & Watching Television	10	Carpentry	32
Sitting and Talking	15	Bricklaying	28
Dressing or Washing	26	Farming Chores	32
Standing	12	House Painting	29
Pick and Shovel Work	56	Walking Downstairs	56

Chopping Wood	60	Walking Upstairs	146
Dragging Logs	158	Walking at 2 miles per hour	29
Making Beds	32	Walking at 4 miles per hour	52
Washing Floors	38	Running at 5.5 miles per hour	90
Washing Windows	35	Running at 7 miles per hour	118
Dusting	22	Running at 12 miles per hour	164
Preparing a Meal	32	Cycling at 5.5 miles per hour	42
Shoveling Snow	65	Cycling at 13 miles per hour	89
Light Gardening	30	Sitting Writing	15
Weeding Garden	49	Light Office Work	25
Mowing Grass with Power Mower	34	Standing with Light Activity	20
Mowing Grass with Manual Mower	38	Typing with Computer	19
Assembly Line	20	Badminton	43
Horseback Riding	56	Baseball	39
Ping-Pong	32	Basketball	58