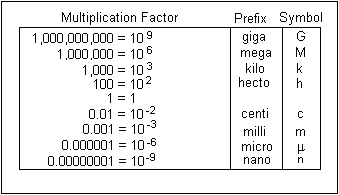
**Chapter 2. Measurements in Chemistry**

**2.1 Measurement Systems**

Measuring system is an integral part of any commerce or science 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 13th century. For example, until gallon was made the standardized measure of volume consolidating different types of gallons (ale, wine and corn).

In France and in 1799, a decimal system called metric system using centimeter, gram, and second (**CGS system**) for length, mass and time, respectively. All units are compared to a standard 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**



**Prefixes** allow simplifying the numbers with many zeros before and after the decimal place.

**Convert measurements to a unit that replaces the power of ten by a prefix:**

6.80 x 10-9 m (**6.08 nm**) 7.14 x 10-6 s (**7.14 m**) 2.88 x 10-3 g (**2.88 mg**)

2.54 x 10-2 m (**2.54 cm**) 4.56 x 103 g (**6.08 kg**) 7.14 x 106 s (**7.14 Ms**)

**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. SI units are based prefixes and upon **7 principal units**, 1 in each of 7 different categories measurements.

|  |  |  |
| --- | --- | --- |
| **Category** | **Name** | **Abbrev.** |
| 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 as needed. So one pound has 16.000000000000.... ounces. If you see any numbers you use in a calculation comes from a definition you could assume they are exact numbers.

**Inxxact 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 a balance and found that it has 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. 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. **Metric to metric system** conversion factors are **exact**

1. 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.)

**An example of an extensive property is**

a. color. b. freezing point. c. length. d. density.

Top of Form

|  |  |  |  |  |  |
| --- | --- | --- | --- | --- | --- |
|  | a | b | c | d | e |

Bottom of Form

**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 **uncertain digit**. Describe uncertainty and significant figures and how they are obtained for a measurement. The significant figure for this inexact number is obtained counting the other digits to the left from the uncertain digit. Therefore, 25.0125 g would have 6 significant figures: 25.013 **(5 SF)** , 25.01 g **(4 SF)** 25.0 **(3 SF)** g and 25 g **(2 SF).** Lower the SF, higher the uncertainty of the number and less accurate.

**GENERAL RULES FOR FIGURING WHICH NUMBERS are significant**

1. **ZEROS** used to "place" the decimal are **NOT significant figures**:

**0.0**15 g = 2 SF

1. **LEADING ZEROS BEFORE** **all the digits** are **NOT significant**:

000340 = 3 SF and 0.000216 g = 3 SF

1. **TRAILING ZEROS after all the digits** are **SIGNIFICANT**:

1.5**00** g= 4 SF

1. **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/in2 **(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 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 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**   1. **"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.** 0.0**67 (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 |
| **34,50**0.00 | **3.450** x 10 4 |
| 0.0000**345**0000 | **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 1067 x -**8.333** x 10-54 = (**2.56** x -**8.333**)(1067 x10-54) =

(-**21.3**3248)(1067-54) = (-**21.3**)(1013)

= (-**2.13**)(1014) = -**2.13** x 1014

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

(**3.46** x 1057 ) / (**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 (1057/10-75)) = ((**3.46** / (**9.431**)) x (1057x10-(-75))  
= (**3.46** /**9.431**)) x 10 57+75 = (**3.46** /**9.431**)) x 10 57+75

= 0.**366**875199 x 10137 = (**3.66**875199 x 101 ) x 10137

= (**3.67**) x ( (101 ) x 10137) = (**3.67** x (10138) = **3.67** x 10138

**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 =4.048x103  m2  1 acre =4840 yd2  1 mile2=2.589x106 m2 1 mile2= 640 acres | 1 gal = 4 qt  1 qt = 2 pt  1 gal = 3.785 L  1 L = 103  mL  1 mL = 1 cm3 |
| **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 | 1 cm | Align the conversion factor to a cancel cm |
| 10-2 m |

|  |  |  |
| --- | --- | --- |
| 4 ~~cm~~ | 10-2 m | = 4 x 10-2 m |
|  | 1 ~~cm~~ |

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

|  |  |  |  |  |
| --- | --- | --- | --- | --- |
| **1.00**  ~~m~~ | **100** ~~cm~~ | **1** in | = **39.3**7008 m | = **39.3** m |
|  | **1** ~~m~~ | **2.54** ~~cm~~ |

or

|  |  |  |  |  |  |  |  |
| --- | --- | --- | --- | --- | --- | --- | --- |
| **1.00**  ~~m~~ | x | **100** ~~cm~~ | x | **1** in |  | = **39.3**7008 m | = **39.3** m |
|  | **1** ~~m~~ | **2.54** ~~cm~~ |

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 102 cm)

**Mass**

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

|  |  |  |  |  |
| --- | --- | --- | --- | --- |
| **150** ~~lb~~ | **0.454** ~~kg~~ | **1000** g | = **681**00 g | = **6.81**x 104 g |
|  | **1** ~~lb~~ | **1** ~~kg~~ |

**Scientific notation allows to show SF clearly.**

**Convert** 65.5 centigrams to milligrams (Ans: 655 mg)

**Convert** 5 liters to cubic decimeters (Ans: 5 dm3)

The density of a substance is 2.7 g/cm3. **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 cm2****are in a m2 (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 cm2 = (10-2)2 m2

**1 cm2 = 10-4 m2**

**Volume unit:**

**How many cm3 are in m3 (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 cm3 = (10-2) 3 m3

**1 cm3 = 10-6 m3**

**How many m3 are in m3 (base unit) of volume?**

1 m = 10-6 m (Ans: **1 m3 = 10-18 m3**)

**How many m3 (base unit) of volume are in cm3.**

1 m = 106 m (Ans: **1 m3 = 1018 m3**)

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

**2.8 Density**

Density is one of thephysical 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 and liquid to the gas. The **unit of mass** is grams, **g.** The **unit of volume** of **sold** is cubic centimeters is **cm3**, **liquids** milliliters is **mL,** and for gases is liters **L or** cubic meters **m3**.

**Solids: d  = grams (g) ÷ cubic centimeters (cm3) = g/cm3**

**Lliquids: d = grams (g) ÷ milliliters (mL) = g/mL**

**Gases: d = grams (g) ÷ milliliters (L or m3) = g/L or g/m3**

**A student determines that a piece of an unknown solid material has a mass of 5.854 g and a volume of 7.57 cm3. 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.

|  |  |  |  |
| --- | --- | --- | --- |
| D = | M | = | ? |
| V |

M= 5.854 g, V = 7.57 cm3

Second: Substitute the known values in the formula

|  |  |  |  |  |  |
| --- | --- | --- | --- | --- | --- |
| D = | M | = | M= **5.854** g | = 0.**773**36 **g/cm3** | = **7.73** x 101 **g/cm3** |
| V | V = **7.57** cm3 |

|  |  |  |
| --- | --- | --- |
|  |  |  |
|  |

**Calculation using a factor label method:**

|  |  |  |  |
| --- | --- | --- | --- |
| **5.854** g | **1** | = 0.**773**36 | = **7.73** x 101 **g/cm3** |
|  | **7.57** 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 = | | M | ; V = | | M | = | M= **14.2** g | = | **5.25**93 **g/cm3** |
| V | D | D= **2.70 g/1 cm-3** |
| = | **5.26** g/cm3 | | |
|

**Calculation using a factor label method:**

|  |  |  |  |
| --- | --- | --- | --- |
| **14.2** ~~g~~ | **1 cm-3** | **= 5.25**93 **cm3** | **= 5.26** |
|  | **2.70 ~~g~~** |

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**)**:**

oC -->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 oF. Convert this****temperature to**

1. **oC and b) K scale.**

**a)****?** °C = (5/9) ((**x**)°F -32) = (5/9) ((**98.6** )°F -32) = 5/9 (66.6) = **37.0** °C

1. **?** 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 102 K

Describe heat energy and how they are measured in calories, dietary calories and joules.  
**Chemical Connections**: Describe Body Density and Percent Body Fat and Normal Human Body Temperature

Describe elements