Chapter 1 – Chemistry and Measurements

Section 1-1: The Numerical Value of a Measurement

1) Sec 1-1.1 – Chemistry – Elements, Compounds, Atoms, and Molecules
Chemistry is the study of matter and the changes it undergoes.

Matter is anything that has mass and occupies space. The most common ways to look at matter in Chemistry are as **elements** and **compounds**.

All of the matter that you see around you and out in space is composed of fewer than 90 unique substances called **elements**. Most of the elements on earth are not present in their free form but are chemically combined with other elements to form **compounds.**

Each element is made up of tiny basic particles called **atoms**. Most compounds are made of **molecules** which are combinations of atoms of two or more different elements. We will later be even more specific about these names.

We will spend more time with this starting in Chapter 2.

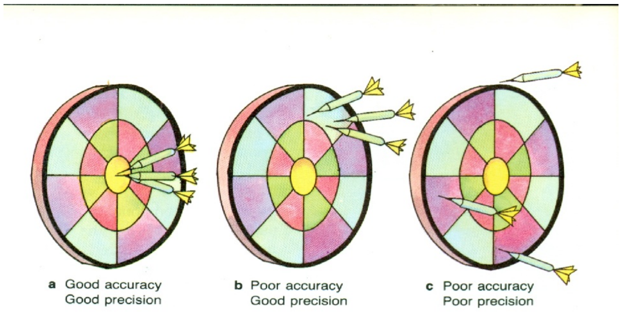
2) The Importance of Measurement
a) Qualitative Measurements give results in a descriptive, non-numeric form.
b) Quantitative Measurements give results in definite form, usually as numbers.

Example – Classify as qualitative or quantitative
1. By feeling your forehead, you say you have a fever.
2. By using a thermometer, your temperature is measured as 102.5°F.

3) Sec 1-1.2 – Measurement of the Mass of an Element and Significant Figures
Each atom has a measured mass. Inside the front cover of your book you have the elements listed alphabetically each with its own specific atomic mass. The unit for atomic mass is called the **atomic mass unit**. We will talk more about this in detail later.

You will note that some of the atomic masses are pretty detailed (hydrogen at 1.00794 or fluorine at 18.9984032) and others are not (zinc at 65.39 or curium at 247). The difference comes from how the masses are measured. We will used these masses for all kinds of things as we go forward. Right now we want to define a measurement and the meanings of the values the measurements express.

A **measurement** determines the quantity, dimensions, or extent of something, usually in comparison to a specific unit. A **unit** is a definite quantity adopted as a standard of measurement. In a measurement, a **significant figure (or digit)** is a digit that is either reliably known or closely estimated.
a) Accuracy is how close a measurement comes to the actual dimension or true value of whatever is measured.
b) Precision is concerned with the reproducibility or uncertainty of whatever is measured. It is how close together separate measurements are to each other.



Example – Comment on the accuracy and precision of the following experimental results for three different trials each. The accepted value is 22.4.
1. Abbey obtained results of 18.6, 19.0, and 18.9.
2. Nikki obtained results of 18.6, 20.6, and 28.0.
3. Julie obtained results of 22.2, 22.3, and 22.6.

Significant figures help us to express the precision of our measurements and calculations from those measurements in a way that others can understand what we did and meant.

4) Significant Figures in Measurements

a) We often make measurements that we then use to calculate other values. The value obtained in the calculation is obviously related to the accuracy of the measurements, but it is also related to how much or where we "round off" our calculations.

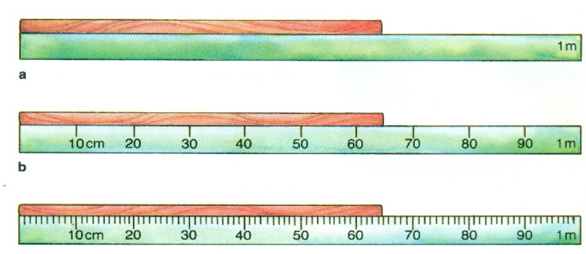
Example – Divide 365 by 2240. Then multiply the answer by 215. What value do you get if you round the quotient to:
1. one place past the decimal point before multiplying by 215 Ans. =
2. two places past the decimal point before multiplying by 215 Ans. =
3. three places past the decimal point before multiplying by 215 Ans. =
4. leave the numbers in your calculator and multiply by 215 Ans. =

b) The accuracy of any measurement depends upon the instrument used and upon the observer. The data that should be recorded consist of the definitely known digits plus one estimated digit. The digits that are known plus the first digit that is uncertain are known as the significant digits or significant figures.

Three different meter sticks could be used to measure the length of a board.

What measurement do you get in each case?

The greater the number of divisions on the meter stick, the greater the number of significant figures in the measurement.



c) Not all the numbers we use come from measurements. Sometimes we use numbers that come from a direct count of objects or that come from definitions. Such numbers are called exact numbers and are said to have an "infinite" number of significant figures. How many are in this room? Eggs in a dozen?

5) Sec 1-1.3 – Zero as a Significant Figure

In a recorded measurement, we will assume the following:

1. All non-zero digits are significant. Ex.: 123578 or 3.1416

This is the easy part. The hard part is what to do with zeros. A zero can be a real or reliable value, or it can be just a place holder to tell us where to place the decimal point. But there is no easy way to distinguish between the two so we come up with some rules to help us.

2. "Imbedded" zeros are significant (zeros between other digits). Ex.: 709 or 1.008

3. Zeros at the end of a number and to the right of the decimal point are significant. Ex.: 2.540 or 0.200

4. Zeros used to place the decimal
   a. are not significant if they are "leading" zeros (zeros to the right of the decimal with no non-zero digits in front.) Ex.: 0.0023
   It should be noted, however, that although the leading zeros are not significant that does not mean that they are not ***important*** *because they do give us an indication of the magnitude of the number.*

   b. may or may not be significant if they are "trailing" zeros (zeros to the right of non-zero digits but to the left of the decimal point). Ex.: 12500 or 890

   One simple way to solve the confusion with 4b is to place a decimal point after the zero at the end if it/they is/are significant. But what happens if not all the zeroes are significant? There is another way that we will get to soon.

Example How many significant digits are in the following measurements?

1. 3658 m 2. 0.0260 L 3. 1.206 4. 0.0003 Km 5. 12 people

Section 1-2: Significant Figures in Chemical Calculations

6) Sec 1-2.1 – The Mass of a Compound

We can use the atomic masses of the elements involved in a compound to determine the mass of the compound. Sometimes this is referred to as the **molecular weight** but, more appropriately, as the **molecular mass** or these days, often, the **molar mass.** (We will get more on this later.) This is accomplished by simply adding up the atomic masses of the atoms present. However, when we get done the final answer has to be properly expressed.

7) Sec 1-2.2 – Molecular Weight and the rules for Rounding Off and Addition and Subtraction

Most calculations involve numbers measured by someone. Calculations should indicate only the degree of accuracy justified, never a higher accuracy, never the limit of the calculator. An answer cannot be more precise than the least precise measurement.

Calculators will just give numbers but are not able to determine significant figures. Like it or not, that is something we have to do. (Note: Calculators also know nothing of units!

Rounding Off Rule – If the digit following the last significant digit is less than 5, all the digits after the last significant digit are dropped. If the digit following the last significant digit is 5 or greater, the value of the digit in the last significant place is increased by 1.

Example – Round off each measurement to three significant figures.

1. 87.073 = 2. 4.3621×10^8 = 3. 0.01552 =

In **addition or subtraction**, the sum or difference cannot be stated to more places after the decimal point than the number with the fewest number of places after the decimal point.

That is, the answer cannot be more precise (more significant figures) than the least precise (fewest decimal places) number in the calculation.

Examples (Think decimal places!)

1. 15.7 + 6.40 + 14.8968 = 2. 150.00 - 2.1 + 1.030 =

Say we have the compound cuprous chloride, CuCl. Copper has an atomic mass of 63.546 and chlorine is 35.4527. With one copper and one chlorine we get the mass by just adding:

$$\begin{array}{l} 63.546 \text{ amu} \\ \underline{35.4527 \text{ amu}} \\ 98.9987 \text{ amu} \end{array}$$

Since copper has three decimal places and chlorine has four, copper has the fewest and we would round the answer to three decimal places, or 98.999 amu.

Example

1. What is the mass of hydrogen chloride (HCl)? It contains 1 hydrogen (atomic mass = 1.00974) and 1 chlorine.

8) Sec 1-2.3 – The Rules for Multiplication and Division

In multiplication and division, we consider the number of significant figures in the answer rather than the number of decimal places. The answer is expressed with the same number of significant figures as the multiplier, dividend, or divisor with the least number of significant figures.

Examples (Think significant figures!)

1. $2.38 \times 2.1 =$

2. $\dfrac{2.460 \times 25.6}{0.0064} =$

If the calculation involves an exact number or a count, then the calculation is like a short-cut to addition and the answer is expressed to the appropriate number of decimal places.

Further, if the calculation is mixed, that is that it involves addition/subtraction and multiplication/division, the rules for multiplication and division take precedence.

9) Sec 1-2.4 – Calculating Percent Error

When we do a measurement or calculation sometimes we already know what the accepted or real value is. In this case we can determine how far our measurement or calculation is away from that value. This determination is referred to as the **error**.

However, the error may not actually tell us what we want to know, so we can go a step farther and compare the error to the accepted value by calculating the **percent error**.

Error = |Accepted Value - Experimental Value| (always positive)

$$\text{Percentage Error} = \frac{|\text{Error}|}{\text{Accepted Value}} \times 100\%$$

Examples: What is the error and percentage error in each of the following?

a) A student sold \$10 worth of candy, but only collected \$6.

b) A clerk sold \$10,000 worth of goods, but collected only \$9996.

Section 1-3: Expressing the Large and Small Numbers Used in Chemistry

10) Sec 1-3.1 – Chemistry and Scientific Notation

Chemical calculations often involve very large or very small numbers. Dealing with those number can be difficult.

Very large and very small numbers are easier to work with when they are put into exponential form. A number is in scientific notation when it is written as $A \times 10^n$, where $1 \le A < 10$ and n is an integer.

$602{,}200{,}000{,}000{,}000{,}000{,}000{,}000 = 6.02 \times 10^{23}$ or
$0.00000000000000000218 = 2.18 \times 10^{-18}$

Note: $10^0 = 1$, $10^1 = 10$, $10^2 = 100$, $10^{-1} = 0.1$, $10^{-2} = 0.01$

Examples

a) Write the following numbers in scientific notation:

a) 5450 = b) 0.0000786 =

b) Write the following numbers as regular numbers:

a) 5.68×10^5 = b) 3.46×10^{-5} =

11) Sec 1-3.2 – Mathematical Manipulation and Scientific Notation

We can do all the calculations we mentioned earlier using scientific notation. Your calculator will allow you to do these calculations relatively easily.

One important note is that scientific notation allows us to consider significant zeros and not be concerned by those that are not significant. Thus, for the number 12000 where the zeros may/may not be significant:

1.2×10^4 has two significant figures
1.20×10^4 has three significant figures
1.200×10^4 has four significant figures
1.2000×10^4 has five significant figures

So, in our earlier example (page 3) 12500 could be expressed with the proper amount of precision by using scientific notation, while we do not know with standard notation.

Section 1-4: The Units Used in Chemistry

We have talked about the numerical part of measurements and now we will look at the other part - the units.

12) Sec 1-4.1 – The International System of Units, SI (Metric System)

The Metric system, developed in France near the end of the 18th century, is used in scientific work throughout the world. It is in general use in practically all countries except the United States. Many American industries that participate in foreign markets have adopted the metric system. By the Metric Conversion Act of 1975, the United States was formally committed to encourage, but not to require, the change to metric measurements.

The metric system is a decimal system that has a single, simple numerical relationship between units. Computations with measurements are easily performed. A standard set of prefixes is used to expand the basic metric system. A revised version, the International System of Units (SI) was adopted by international agreement in 1960. The SI system has seven basic units from which other SI units of measurement such as volume, density, and pressure are derived. Although it is possible to report all measurements in SI units, we will not do so in this class. And there are some common units that most scientists commonly accept.

13) Sec 1-4.2 – Fundamental versus Derived Units

Of the possible measurements that can be made, only seven are considered to be

fundamental, that means that they cannot be described in terms of anything else. We are going to be principally concerned with five of the fundamental units.

Measurement	Unit	Symbol
Mass	kilogram	kg
Length	meter	m
Time	second	s
Temperature	Kelvin	K
Quantity	mole	mol

By combining these five units in many useful ways by multiplication and division we can get other units that are referred to as **derived** units. Several possibilities are:

Measurement	Unit	Symbol	Derivation
Volume	cubic meter	m^3	$Length^3$
Energy	joule	J	Mass x $Distance^2/Time^2$
Pressure	pascal	Pa	Mass/($Time^2$ x Distance)

14) Units of Length – The meter (m) is the basic unit of length
1 m = 39.4 in, 1 in = 2.54 cm, 1 mile = 1.61 km

Examples: What metric unit of length would be the best to describe the following?
a) The length of your foot
c) The diameter of a bolt

b) The distance between Saint Louis and Chicago

15) Units of Volume – The space occupied by matter is called volume. It is a derived unit. Volume is calculated from linear measurements of length, width, and height. The SI unit of volume is the cubic meter (m^3). A more convenient metric unit is the liter (L). A liter is the volume of a cube that is 10 cm (1 dm) on each edge. There are 1000 liters in a cubic meter. (1 m^3 = 1000 L)

1 L = 1000 cm^3 (cc) and 1 mL = 1 cc 1 L = 1.06 qt, 1 qt = 946 mL

Examples: What metric unit of volume would be used for the following?
1) A tankful of gasoline
2. Amount of milk added to a recipe

16) Units of Mass
a) Mass is a measure of the total quantity of matter in an object. Weight is a measure of the force exerted on an object by the pull of gravity. The mass of an object is always the same, but the weight of an object is a variable depending upon the geographical location of the object.

Note: This is one reason why many chemists refer to the mass of a compound as molecular mass instead of molecular weight , sections 1-2.1 and 2.2.

b) The mass of an object is measured by comparing it to a standard mass of 1 Kg, which is the basic SI unit of mass. A kilogram is the mass of 1 L of water at 4°C.

Examples: What metric unit of mass would be used for the following?
1. Your mass
2. The mass of a penny

17) Sec 1-4.3 – Metric System Prefixes

The metric system uses prefixes to change between values for units. These prefixes are convenient powers of 10 and work well with scientific notation. Your text lists a number of the prefixes. In this class we will use the following prefixes:

kilo (k) = 1000 = 10^3 deci (d) = 0.1 = 10^{-1} centi (c) = 0.01 = 10^{-2}

milli (m) = 0.001 = 10^{-3} micro (μ) = 10^{-6} nano (*n*) = 10^{-9}

18) Sec 1-4.4 – The English System of Measurements

The English system used in daily activities in the United States presents some problems in scientific use. In a sense, it is a system that just grew. The units have practical size for common use. However, the chief disadvantage of the English system is that no single, simple numerical relationship exits between different units of measure. For example, 12 inches = 1 foot, 3 feet = 1 yard, 5280 feet = 1760 yards = 1 mile, 1 lb = 16 oz, 2 pints = 1 quart, 4 quarts = 1 gallon.

19) Sec 1-4.5 – Relationships Between Metric and English Units

Table 1-5 in your text lists common conversion factors between metric and english units that you might encounter. There is also a table inside the back cover.

I have given you a sheet of common conversion units. On it I have indicated the metric conversions you must know. The others are for reference. When we do English-metric conversions (on tests) I will give factors that can get you to the answer.

Section 1-5: A Technique of Problem Solving – Dimensional Analysis

In dimensional analysis we use the units (dimensions) that are part of measurements to help solve (analyze) the problem. Many problems in chemistry are merely conversions from one kind of units to another.

20) Sec 1-5.1 through 5.3 – Single Step and Multistep Conversions

Units cancel just as numbers do. If you cancel the units FIRST, then the numbers will have to give you the correct answer.

Steps

a) Identify the unknown as well as the units of the quantity to be determined.

b) Identify what is Known or Given (both a numerical value and its units).

c) Plan a solution. You will need conversion factors that will get you from the given units to the unknown's units. Obtain these conversion factors from what you know from what is available to you.

d) Multiply by the conversion factors in a manner such that the unwanted (original) units are cancelled out, leaving only the desired units. At this point you will be able to tell if you have everything that you need, or if you are missing something or have made an error.

e) Do the calculations as indicated by the conversion factors.

f) Finish up. The answer should always be expressed to the correct number of significant figures. Also, check your work. Does your answer make sense?

Conversion Factors – In a conversion factor, the measurement in the numerator is equivalent to the measurement in the denominator. When a measurement is multiplied by a conversion factor, the value of the measurement remains the same. Although the numerical value of the measurement is changed, the change in the units compensates for this. Many conversion factors are defined quantities and thus have an unlimited number of significant figures.

Note: a x 1 = a; if 1 ft = 12 in so $\frac{1\text{ ft}}{12\text{ in}} = 1$ and $\frac{12\text{ in}}{1\text{ ft}} = 1$

Converting Between Units – Examples

a) 500 g = Kg

b) 250 cc = L

c) 53.5 cm = Km

d) 4.65 Kcal/min = cal/sec

Multi-Step Problems and Converting Complex Units – Examples

a) 100 yards = meters

b) An automobile can get 15 Km per liter. What is this in miles per gallon?

c) How many square meters are there in a lot that is 80 feet by 120 feet?

d) How many pounds of water can be held by a cylindrical tank with a radius of 2 feet and a height of 6 feet? (Assume the density of water is 1.0 g/mL.)

e) How large a container, in liters, would you need to hold 10.0 Kg of alcohol? (Assume a density of 0.790 g/cc for alcohol.)

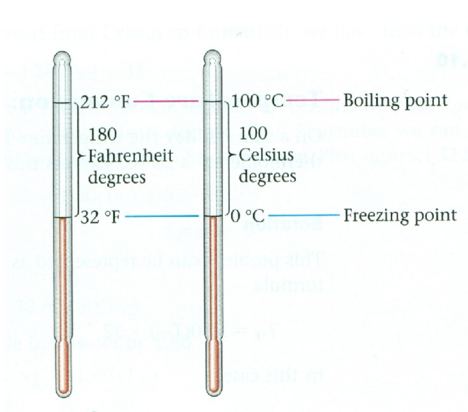
Section 1-6: Measurement of Temperature

Temperature is a measure of the hotness or coldness of an object. Almost all substances expand with an increase in temperature and contract as the temperature decreases. This property is the basis for most common thermometers.

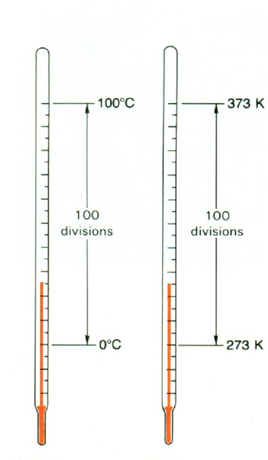
21) Sec 1-6.1 – Thermometer Scales

Common thermometers will give temperatures in one of two scales: the Celsius scale or the Fahrenheit scale.

A third temperature scale that we will use in calculations in chemistry is the Kelvin or absolute scale. *(Remember that Kelvin is the fundamental unit of temperature in the metric system.)* Within thermometers the fluids expand linearly by the same amount no matter what the temperature is.



Comparison of the Celsius and Fahrenheit scales.



0°K (-273°C, -459°F) is the predicted lowest temperature possible. For calculations, there can be no negative temperatures.

22) Sec 1-6.2 – Relationships Between Scales

In looking at the adjustments and the differences in degrees between Celcius and Fahrenheit one can determine how to convert from one scale to another.

Thus: °F = 1.8 x °C + 32 °C = (°F - 32) / 1.8 °K = °C + 273

23) *I have another method for converting between °C and °F. I learned it from Dr. Werner Bromund at Oberlin College. It is based on the fact that at – 40 both °C and °F scales read the same.*

°F = [(°C + 40) x 1.8] – 40 and °C = [(°F + 40) / 1.8] – 40

Examples:

1. What is normal body temperature (98.6°F) in °C and °K?

2. a) What is the °F temperature outside if your thermometer reads -15°C?

   b) What is that temperature in K?

24) Density is another derived measurement and is the ratio of the mass of an object to its volume.

$$\text{density} = \frac{\text{mass}}{\text{volume}}$$ units are g/mL or g/cc

We have already seen above how it might be a conversion factor to go from the volume of a liquid to the mass of that liquid.

Densities of some common substances at 25°C (in g/cc)
Au = 19.3, Hg = 13.6, Pb = 11.3, Al = 2.70, H_2O = 0.997,
ethanol = 0.789

Sample Calculations

1. A stone removed from the gall bladder of a patient has a volume of 2.2 cc and a mass of 1.89 g. What is its density? Will it float in water?

2. When 49.33 g of lead were placed in a graduated cylinder containing 15.0 mL of water, the water level increased to 19.4 mL. What is the density of the lead? (This is a method of measuring "volume by difference" which is especially helpful for irregular shapes, etc.)

3. Calculate the density of a piece of metal that has a mass of 38.4 g and is 5.2 cm long, 2.3 cm wide, and 1.1 cm high.