General Chemistry
(Chem 1012)

Chemistry Department
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Unit 1

1. Chemistry and Measurement

- Chemistry is a science studying the composition, characteristics, properties, and transformations of all matter in the known universe.
- Measurement is an important part of most experiments involves the determination (often, the estimation) of quantity, volume, dimensions, capacity, or extent of something – these determinations are measurements.
- A measurement gives a number to that property. Measurements are always made using an instrument of some kind. Rulers, stopwatches, weighing scales, and thermometers are all measuring instruments.

Base Units of the SI System

Physical Quantity	Name of Unit	Abbreviation
Mass	Kilogram	kg
Length	Meter	m
Time	Second	s[a]
Temperature	Kelvin	K
Amount of substance	Mole	mol
Electric current	Ampere	A
Luminous intensity	Candela	cd

Common Unit Prefixes

Prefix	Abbreviation	Meaning	Example
Giga	G	10^9	1 gigameter (Gm) = 1×10^9 m
Mega	M	10^6	1 megameter (Mm) = 1×10^6 m
Kilo	k	10^3	1 kilometer (km) = 1×10^3 m
Deci	d	10^{-1}	1 decimeter (dm) = 0.1 m
Centi	c	10^{-2}	1 centimeter (cm) = 0.01 m
Milli	m	10^{-3}	1 millimeter (mm) = 0.001 m
Micro	μ[a]	10^{-6}	1 micrometer (μm) = 1×10^{-6} m
Nano	n	10^{-9}	1 nanometer (nm) = 1×10^{-9} m
Pico	p	10^{-12}	1 picometer (pm) = 1×10^{-12} m
Femto	f	10^{-15}	1 femtometer (fm) = 1×10^{-15} m

1.2. Measurement Uncertainty

- **Uncertainty** of measurement is the doubt that exists about the result of any measurement.
- Counting is the only type of measurement that is free from uncertainty and gives exact number.
- The uncertainty of a measurement tells us something about its quality.
- Every measurement is subject to some uncertainty. A measurement result is only complete if it is accompanied by a statement of the uncertainty in the measurement.
- Uncertainty is a quantification of the doubt about the measurement result.

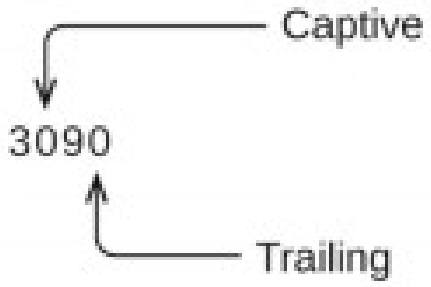
- It is important not to confuse the terms ‘error’ and ‘uncertainty’.
- Error is the difference between the measured value and the ‘true value’ of the thing being measured.

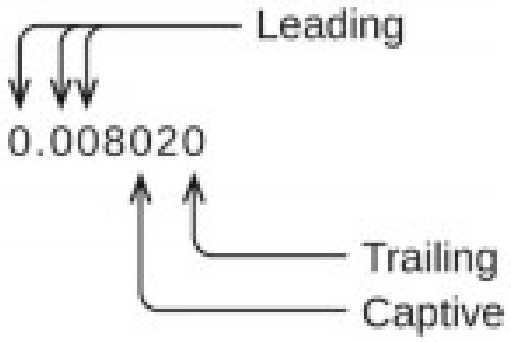
1.1.2 Significant Figures in Measurement

- Every measurement has a degree of uncertainty associated with it.
- The uncertainty derives from the measuring device and from the skill of the person doing the measuring.
- The **significant figures** in a measurement consist of all the certain digits in that measurement plus one uncertain or estimated digit.
- Significant figures or significant digits are all of the digits in a measurement, including the uncertain last digit.
- The term number of significant figures refers to the number of digits reported for the value of a measured or calculated quantity, indicating the precision of the value.

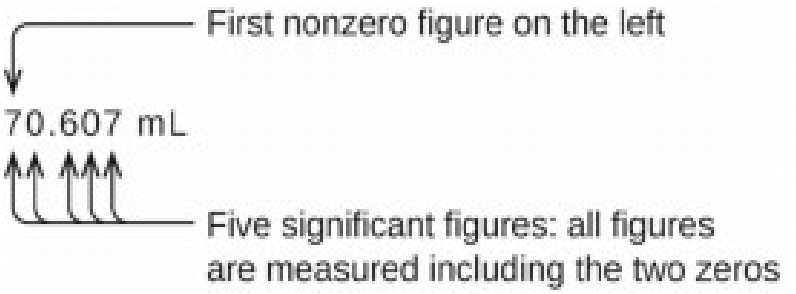
Scientific notation

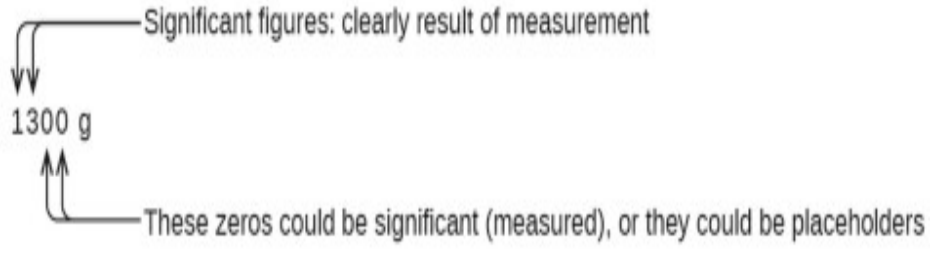
- Scientific notation is the representation of a number in the form $Ax10^n$, where A is a number with a single nonzero digit to the left of the decimal point and n is an integer, or whole number.
- In scientific notation, the measurement 900 cm precise to two significant figures is written 9.0×10^2cm.
- All nonzero digits are significant, and it is only zeros that require some thought. We will use the terms “leading,” “trailing,” and “captive” for the zeros and will consider how to deal with them.

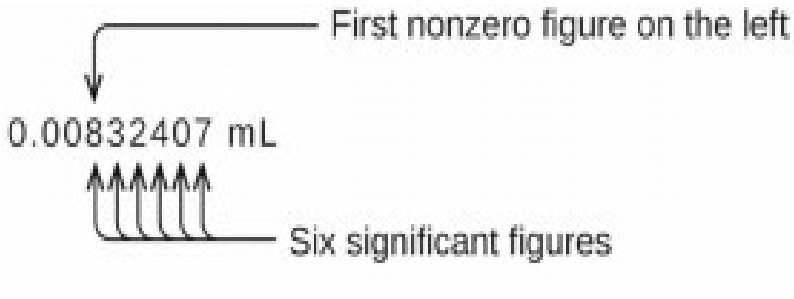




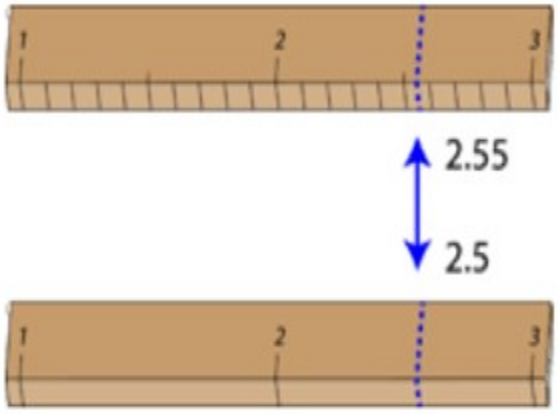
- Starting with the first nonzero digit on the left, count this digit and all remaining digits to the right.
- This is the number of significant figures in the measurement unless the last digit is a trailing zero lying to the left of the decimal point.
- Captive zeros result from measurement and are therefore always significant.
- Leading zeros, however, are never significant—they merely tell us where the decimal point is located







Examples:



Ruler A **2.55** ← plus one estimate

two certain digits

Ruler B **2.5** ← plus one estimate

one certain digit

Significant figures rules

Rule 1: All nonzero digits in a measurement are significant.

237 has three significant figures.

1.897 has four significant figures.

Rule 2: Zeros that appear between other nonzero digits (i.e., "middle zeros") are always significant.

39,004 has five significant figures.

5.02 has three significant figures.

Rule 3: Zeros that appear in front of all of the nonzero digits are called leading zeros. Leading zeros are never significant.

0.008 has one significant figure.

0.000416 has three significant figures.

Rule 4: Zeros that appear after all nonzero digits are called trailing zeros. A number with trailing zeros that lacks a decimal point may or may not be significant.

1400 is ambiguous.

1.4×10^3 has two significant figures.

1.400×10^3 has four significant figures.

Rule 5: Trailing zeros in a number with a decimal point are significant. This is true whether the zeros occur before or after the decimal point.

620.0 has four significant figures and **19.000** has five significant figures.

Exercises

Give the number of significant figures in each. Identify the rule for each.

a.5.87 b. 0.031 c.52.90 d.00.2001 e.500 f.6 atoms

Ans:

a) All three numbers are significant (rule 1)
b) The leading zeros are not significant (rule 3). The 3 and the 1 are significant (rule 1)
c) The 5, the 2 and the 9 are significant (rule 1). The trailing zero is also significant (rule 5)
d) The leading zeros are not significant (rule 3). The 2 and the 1 are significant (rule 1) and the middle zeros are also significant (rule 2
e) The number is ambiguous. It could have one, two or three significant figures
f) The 6 is a counting number. A counting number is an exact number.

1.1.2 Significant figures in calculations

Rounding

- Rounding is the procedure of dropping non-significant digits in a calculation result and adjusting the last digit reported.
- The general procedure is as follows: Look at the leftmost digit to be dropped. Then . . .
- If this digit is 5 or greater, add 1 to the last digit to be retained and drop all digits farther to the right.
- Thus, rounding 1.2151 to three significant figures gives 1.22.
- If this digit is less than 5, simply drop it and all digits farther to the right.
- Rounding 1.2143 to three significant figures gives 1.21.

1.1.3 Accuracy versus Precision

- **Accuracy** refers to the nearness or proximity of a measurement value to the true value of a quantity.
- **Accuracy** refers to how closely the measured value of a quantity corresponds to its “true” value.
- **Precision** refers to the closeness of several measurements to each other (Precision relates to the uncertainty of a measurement).
- **Precision** expresses the degree of reproducibility or agreement between repeated measurements.
- The more measurements you make and the better the **precision**, the smaller the error will be.
- For a measured quantity, we can generally improve its accuracy by making more measurements



1.2 Conversion factors and dimensional analysis

- A conversion factor is the number or formula you need to convert a measurement in one set of units to the same measurement in another set of units.
- A ratio of two equivalent quantities expressed with different measurement units can be used as a unit conversion factor.
- Dimensional analysis (or the factor-label method) is the method of calculation in which one carries along the units for quantities.
- A conversion factor is a factor equal to one that converts a quantity expressed in one unit to a quantity expressed in another unit.
- Suppose

$$1\ \text{L} = 10^3 \text{cm}^3$$

If you divide both sides of the equality by the right-hand quantity, you get

$$\frac{1\ \text{L}}{10^3\ \text{cm}^3} = \frac{10^3\ \text{cm}^3}{10^3\ \text{cm}^3} = 1$$

- Because it is always possible to multiply any quantity by 1 without changing that quantity, you can multiply the previous expression for the volume by the factor 1 L/10^3cm^3 without changing the actual volume. You are changing only the way you express this volume

$$V = 5.00^3\ \text{cm}^3 \times \underbrace{\frac{1\ \text{L}}{10^3\ \text{cm}^3}}_{\substack{\text{converts} \\ \text{cm}^3\ \text{to L}}} = 125 \times 10^{-3}\ \text{L} = 0.125\ \text{L}$$

- The ratio 1 L/10^3cm^3 is called a conversion factor because it is a factor equal to 1 that converts a quantity expressed in one unit to a quantity expressed in another unit.

Common Conversion Factors in Chemistry

Quantity	Conversion factor
Volume	1 gallon = 3.78541 liters
	1 liter = 10 deciliters
	1 liter = 1000 milliliters
	1 mL = 1 cm^3
Mass	1 pound = 16 ounces
	1 kilogram = 1,000 grams
	1 gram = 10^6 microgram
	1 pound = 453.592 grams
Time	1 minute = 60000 milliseconds
	1 minute =60 seconds
	1 hour = 60 minutes
Temperature	1 Kelvin = ºC + 273.15

Conversion of Temperature Units

- Temperature is a measure of “hotness”.
- A thermometer is a device for measuring temperature.
- The Celsius scale (formerly the centigrade scale) is the temperature scale in general scientific use.
- However, the SI base unit of temperature is the kelvin (K), a unit on an absolute temperature scale.
- The Celsius and the Kelvin scales have equal-size units (that is, a change of 1°C is equivalent to a change of 1 K), where 0°C is a temperature equivalent to 273.15 K.
- Thus, it is easy to convert from one scale to the other, using the formula

$$T_K = \left(t_C \times \frac{1\,K}{1^\circ C}\right) + 273.15\,K$$

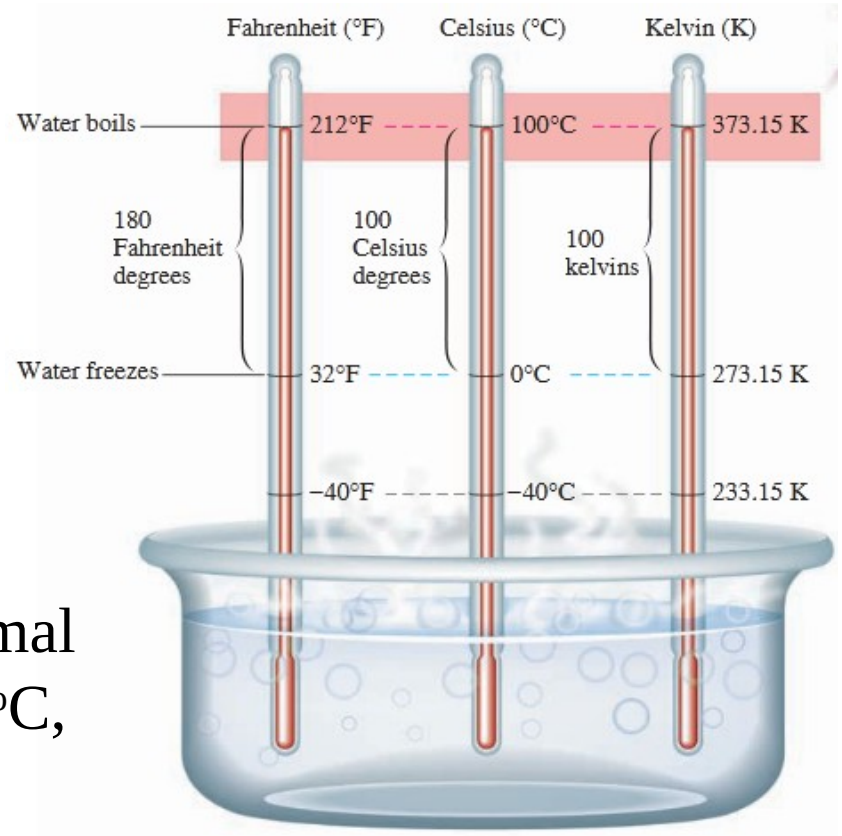
The Fahrenheit scale is at present the common temperature scale in the United States.

$0^{o}C$ is the same as $32^{o}F$ (both exact), and $100^{o}C$ corresponds to $212^{o}F$

$$t_F = \left(t_C \times \frac{9°F}{5°C}\right) + 32°F$$

And we can obtain a formula for converting degrees Fahrenheit to degrees Celsius:

$$t_C = \frac{5°C}{9°F} \times (t_F - 32°F)$$



Water boils under normal pressure at $212^{o}F$, $100^{o}C$, and 373.15 K.

Example:

The hottest place on record in North America is Death Valley in California.

It reached a temperature of 134F in 1913. What is this temperature reading in degrees Celsius? in kelvins?

Solution Substituting, we find that

$$t_C = \frac{5°C}{9°F} \times (t_F - 32°F) = \frac{5°C}{9°F} \times (134°F - 32°F) = \mathbf{56.7°C}$$

In kelvins,

$$T_K = \left(t_C \times \frac{1\ K}{1°C}\right) + 273.15\ K = \left(56.7°C \times \frac{1\ K}{1°C}\right) + 273.15\ K = \mathbf{329.9\ K}$$

Problems

1. How many significant figures are there in each of the following measurements?

a) 73.0000 g
b) 0.0503 kg
c) 6.300 cm
d) 0.80090 m
e) 5.10 107m
f) 2.010 s

2. Convert:

g) 125 Å to micrometers
h) 32.4 kg to milligrams
i) 16.8 cm to millimeters
j) 2.2 ns to microseconds