Groupings

Section 006

Group 1: Adams to Ma: Lab Check in, then review quiz

Group 2: MacDonald, K. to Zheng: Review quiz, then lab check in

Tutorial in SSC 2032; lab check in Chemistry Rooms 10, 11, 12, 13 ground Floor

Significant Figures *(M&H, Ch. 1 p. 11)*

More importantly read lab manual p. 12 – 19; instructions for the calculator are also given there.

Significant figures indicate the reliability of experimental measurements, which usually contain some error or uncertainty

There are two types of experimental uncertainty:
Accuracy: how close is the result to the "true" value
Precision: How well do repeated measurements agree?

The number of sig. figs. used in reporting a result gives an estimate of the limits of error/uncertainty

The error is assumed to be ±1 in the last (least significant) digit; that is, the last digit on the right hand side.
Examples:

- 2.35 g has 3 sig. figs. and means 2.35 ±0.01 g
- 2.3568 g has 5 sig. figs. and means 2.3568 ± 0.0001 g

Are zeros significant? **Sometimes!**

2.350 has four sig. figs. Hence, 2.350 ± 0.001
i.e. trailing zeros are significant

However zeros are not significant when they are used to indicate the position of the decimal point; e.g. 0.0235

How about 235,000?

This is ambiguous and it is better to express it in exponential or scientific notation:

- 2.35×10^5 has an error of $0.01 \times 10^5 = 1000$
- 2.3500×10^5 has an error of $0.0001 \times 10^5 = 10$

Note for exact or defined quantities the concept of significant figures does not apply since there is no uncertainty

1 m = exactly 100 cm   1 L = exactly 1000 mL

In chemical equations, coefficients are exact numbers:

$$\boxed{1}P_4 + \boxed{6}Cl_2 \rightarrow \boxed{4}PCl_3 \text{ (exactly)}$$

The number of sig. figs. may change in a calculation, and this is detailed in the lab manual.
At the moment consider the following simpler method summarized below:

1. **Addition and Subtraction**

Find the measurement with the fewest figures to the right of the decimal point (dp).
Your final answer should also have this many figures to the right of the dp.

$$\begin{array}{r} 2.3468 \\ +\underline{0.026} \\ 2.373 \end{array} \qquad \begin{array}{r} 4.15 \\ -\underline{3.0815} \\ 1.07 \end{array}$$

2. Multiplication and Division

Find the measurements with the fewest number of sig. figs., and this is how many sig. figs. should be in your final answer

Examples:

- $\underbrace{17.15}_{4SF} \times \underbrace{0.0977}_{3SF} = \underbrace{1.68}_{3SF}$
- $\underbrace{4.383\ g}_{4SF} / \underbrace{2.72\ L}_{3SF} = \underbrace{1.61\ g\ L^{-1}}_{3SF}$

Even though your calculator will give you many more digits, always think about sig. figs. before writing them down.

Sample Standard Deviation (ssd)

= mathematical measure of precision

A large value of ssd relative to the average of your measurements (or mean) suggests that measurements were inconsistent

Formulas: $\bar{x} = \text{average value} = \text{mean}$

$$\bar{x} = \frac{x_1 + x_2 + \cdots + x_N}{N} = \frac{1}{N}\sum_{i=1}^{N} x_i$$

Symbol means summation

$$\text{ssd} = \sqrt{\frac{1}{N-1}\sum_{i=1}^{N}\left(x_i - \bar{x}\right)^2}$$

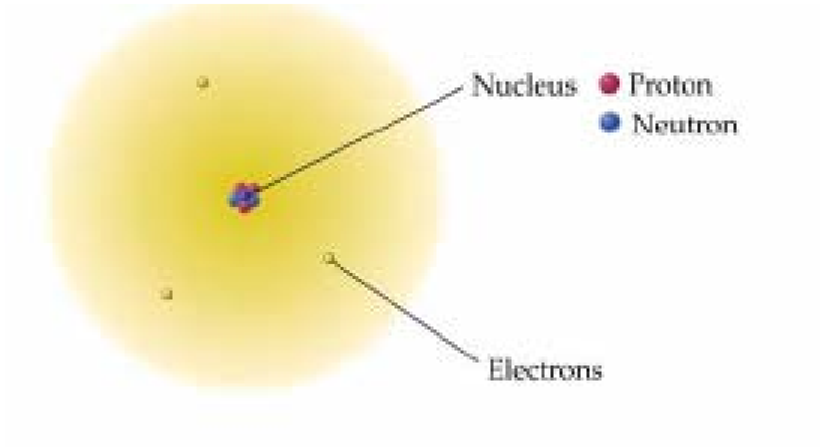
N = number of runs; x_i = the value of each run,

Don’t worry both mean and ssd can be found using the required calculator!

The first non-zero digit shows where measurements become unreliable. This determines the last digit reported for the mean.

Atoms, Molecules and Ions (M&H Ch. 2)

All matter is composed, ultimately, of **atoms** with a nucleus of protons and neutrons, an a surrounding shell (cloud) of electrons



Example: $^{197}_{79}\mathrm{Au}$

79 = **atomic number** = Z= number of protons = # electrons in the neutral atom

197 = **mass number** = number of protons and number of neutrons; that is, the number of neutrons in Au is 197 – 79 = 118

Protons carry +1 charge while neutrons carry 0 charge Electrons carry -1 charge.

Mass (proton and neutron) >>>> mass of electron

Virtually all the atom's mass lies in the nucleus

An element is a substance in which all the atoms have the same atomic number.

Most naturally occurring elements however are mixtures of **isotopes**
(= atoms which have the same atomic number but different mass numbers)

For example: $^{12}_{6}\mathrm{C}$ 98.90% natural abundance

$^{13}_{6}\mathrm{C}$ 1.100% natural abundance

$\therefore$ The average atomic mass of elemental carbon is

$$\frac{98.90}{100.0}(12.0)+\frac{1.100}{100.0}(13.0)=12.01\,\mathrm{amu}$$

amu ≡ atomic mass unit

There is another isotope of carbon ^{14}C which is radioactive, and is widely used in carbon dating to estimate the age of old plants and animal matter

Radioactivity

- Radioactivity is a term used to describe the decomposition of unstable isotopes.
Usually the atomic number Z changes in radioactive decay; that is,
a new element if formed

- In carbon dating the amount of ^{14}C present in an artefact is measured.
Since ^{14}C decays over time, the amount of this unstable isotope can be correlated to the age of the material

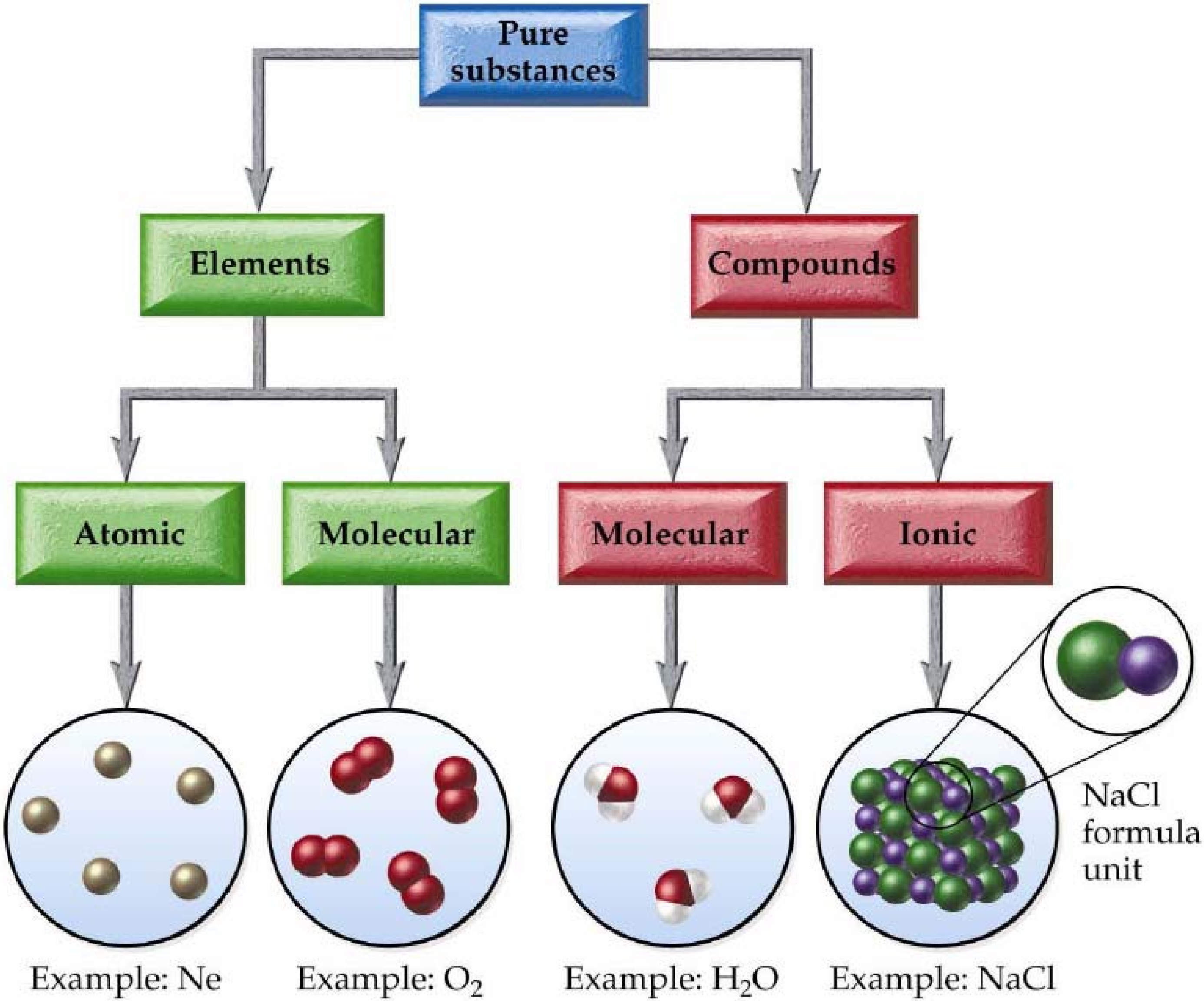
Molecules and Compounds

- A molecule consists of two or more atoms joined together by covalent bonds where the electrons of the atoms are shared

- A compound is a substance in which the atoms of more than one element are present, usually in a whole number ratio;
that is, a compound must have two or more different elements

- These can be classified as compounds and molecules:
  H_2O (one oxygen and two hydrogen atoms)
  CH_4 (one carbon and four hydrogen atoms)

- However, molecules need not be compounds:
  O_2 (two oxygen atoms), N_2, Cl_2, H_2, etc.
  By convention these are still called elements



Pure substances
Elements
Compounds
Atomic
Molecular
Molecular
Ionic
NaCl formula unit
Example: Ne
Example: O_2
Example: H_2O
Example: NaCl