Lecture 2

Outline

2.3 Scientific Notation

2.4 Significant figures and Accuracy

3.1 Structure and Building Principles of Atoms

3.2 Element Symbols

3.3 Masses and the Mole

3.4 Introduction of the Periodic Table

Scientific Notation

Mole

The number of atoms in 12 g of carbon:

602,200,000,000,000,000,000,000

In science **6.022×10^{23}**

The mass of a single carbon atom in grams:

0.0000000000000000000000199

In science **1.99×10^{-23}**

$\boxed{N \times 10^{n}}$

n is a positive or negative integer
N is a number between 1 and 10

Scientific Notation and Significant Figures

568.762

move decimal left

n > 0

568.762 = 5.68762 x 10^{2}

0.00000772

move decimal right

n < 0

0.00000772 = 7.72 x 10^{-6}

Uncertainty in Measurement

- Uncertainties always exist in measured quantities
- Measured quantities are generally reported in such a way that only the last digit is uncertain.
- All digits of a measured quantity, including the uncertain one, are called significant figures.

Sample Problem **Determining the Number of Significant Figures**

PROBLEM: For each of the following quantities, underline the zeros that are significant figures (sf), and determine the number of significant figures in each quantity. For **(d)** to **(f)** express each in exponential notation first.

(a) 0.0030 L **(b)** 0.1044 g **(c)** 53.069 mL

(d) 0.00004715 m **(e)** 57,600. s **(f)** 0.0000007160 cm^3

PLAN: Determine the number of sf by counting digits and paying attention to the placement of zeros.

SOLUTION:

(a) $0.00\underline{30}$ L 2sf **(b)** $0.\underline{1044}$ g 4sf **(c)** $\underline{53.069}$ mL 5sf

(d) 0.00004715 m **(e)** 57,600. s **(f)** 0.0000007160 cm^3

(d) $\underline{4.715}\text{x}10^{-5}$ m 4sf **(e)** $\underline{5.7600}\text{x}10^{4}$ s 5sf **(f)** $\underline{7.160}\text{x}10^{-7}$ cm^3 4sf

Rules for Significant Figures in Answers

- **For addition and subtraction.** The answer has the ***same number of decimal places as there are in the measurement with the fewest decimal places.***

Example: adding two volumes:

$$\begin{array}{r} 83.5\ \text{mL} \\ +\ 23.28\ \text{mL} \\ \hline 106.78\ \text{mL} \end{array} = 106.8\ \text{mL}$$

Example: subtracting two volumes:

$$\begin{array}{r} 865.9\ \text{mL} \\ -\ 2.8121\ \text{mL} \\ \hline 863.0879\ \text{mL} \end{array} = 863.1\ \text{mL}$$

Rules for Significant Figures in Answers

- **For multiplication and division.** The number with the least certainty limits the certainty of the result. Therefore, ***the answer contains the same number of significant figures as there are in the measurement with the fewest significant figures.***

Multiply the following numbers:

$\boxed{9.2}$ cm x 6.8 cm x 0.3744 cm = 23.4225 cm^3 = $\boxed{23}$ cm^3

Rules for Rounding Off Numbers

1. If the digit removed is *5 or more than 5,* the preceding number increases by 1. 5.379 rounds to 5.38 if three significant figures are retained and to 5.4 if two significant figures are retained.

2. If the digit removed is *less than 5,* the preceding number is unchanged. 0.2413 rounds to 0.241 if three significant figures are retained and to 0.24 if two significant figures are retained.

3. Be sure to carry two or more additional significant figures through a **multistep calculation** and round off only the *final* answer.

Sample Problem **Significant Figures and Rounding**

PROBLEM: Perform the following calculations and round the answer to the correct number of significant figures.

(a) $$\frac{16.3521\ \text{cm}^2 - 1.448\ \text{cm}^2}{7.085\ \text{cm}}$$

(b) $$\frac{4.80\text{x}10^4\ \text{mg}\left(\frac{1\ \text{g}}{1000\ \text{mg}}\right)}{11.55\ \text{cm}^3}$$

PLAN: In (a) we subtract before we divide; for (b) we are using an exact number.

SOLUTION:

(a) $$\frac{16.3521\ \text{cm}^2 - 1.448\ \text{cm}^2}{7.085\ \text{cm}} = \frac{14.904\ \text{cm}^2}{7.085\ \text{cm}} = 2.104\ \text{cm}$$

(b) $$\frac{4.80\text{x}10^4\ \text{mg}\left(\frac{1\ \text{g}}{1000\ \text{mg}}\right)}{11.55\ \text{cm}^3} = \frac{48.0\ \text{g}}{11.55\ \text{cm}^3} = 4.16\ \text{g/ cm}^3$$

Precision and Accuracy
Errors in Scientific Measurements

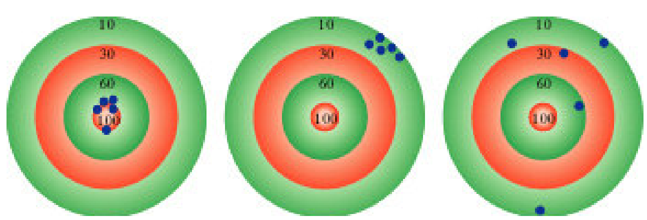
Precision -
Refers to *reproducibility* or how close the measurements are to each other.

Accuracy -
Refers to how close a measurement is to the real value.

Systematic error -
Values that are either all higher or all lower than the actual value.

Random Error -
In the absence of systematic error, some values that are higher and some that are lower than the actual value.

Precision and Accuracy



accurate & precise	precise but not accurate	not accurate & not precise

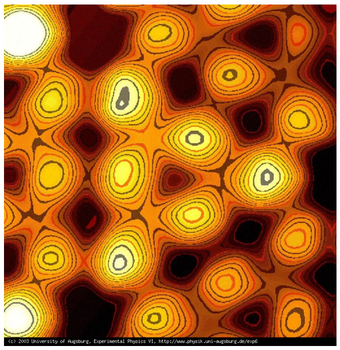
Atoms

- Atoms cannot be sub-divided by chemical reactions: available energies are not high enough
- Atoms can be subdivided, modified or transferred into new atoms by physical reactions (nuclear reactions)
  Reason: energies are sufficiently high
- Chemical properties of atoms are significantly determined by the elemental building units of atoms

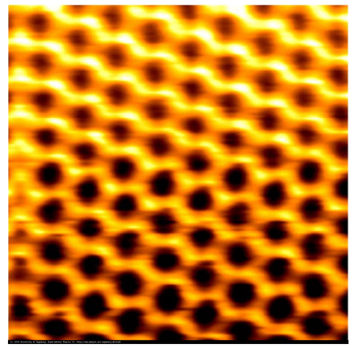
Structure of Atoms:

Atom:	built from elementary particle
elementary particle	Material particles which cannot be divided into smaller particles, but they can react to give other elementary particles
	Protons, neutron, electrons (valid for nearly all atoms: exception the hydrogen atom)
Different atoms:	Differ in the number and arrangement of elementary particles

Imaging Atoms



silicon



graphite

……. Atomic Force Microscopy or Scanning Tunnelling Microscopy

Building Principles of Atoms

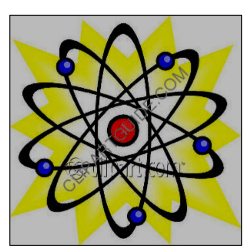
atom ⟶ elementary units

name (symbol)	mass (u)	charge (e)
Proton	1	+1
Neutron	1	-
Electron	0.0005	-1

Atomic mass unit u: $1u = 1.6606 \cdot 10^{-27}$ kg

Elementary charge e: $1e = 1.6022 \cdot 10^{-19}$ C

smallest charge, observed in nature



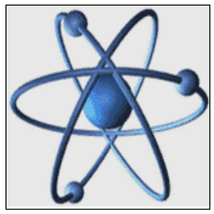
Building Principles of Atoms

protons + neutrons

nuclei

atomic core

radius ca. 10^{-15} m



electrons

electron shell

radius ca. 10^{-10} m

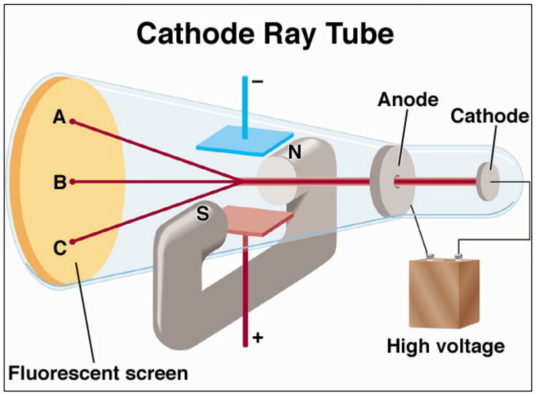
Proton Number + Neutron Number = Mass Number

….. Historical Aspects of these Discoveries.

Dalton's Atomic Theory (1805)

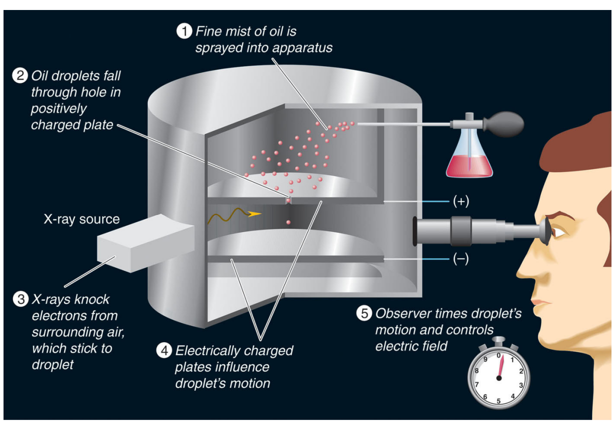
1. All matter consists of **atoms.**
2. All atoms of a given element are identical, having the same size, mass and chemical properties.
3. The atoms of one element are different from the atoms of all other elements.
4. Chemical reactions only involve the rearrangement of atoms. Atoms are not created or destroyed in chemical reactions.

J.J. Thomson (1906 Nobel Prize in Physics)



measured mass/charge of e⁻

Millikan's Oil-Drop Experiment for Measuring an Electron's Charge



Millikan

1923 Nobel Prize in Physics

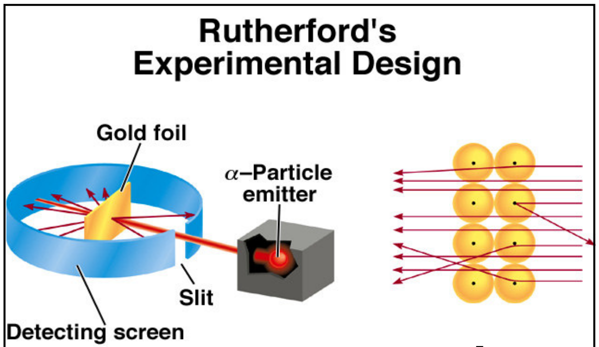
Measured mass of e⁻

e^- charge = -1.60×10^{-19} C

Thomson's charge/mass of e^- = -1.76×10^{8} C/g

e^- mass = 9.10×10^{-28} g

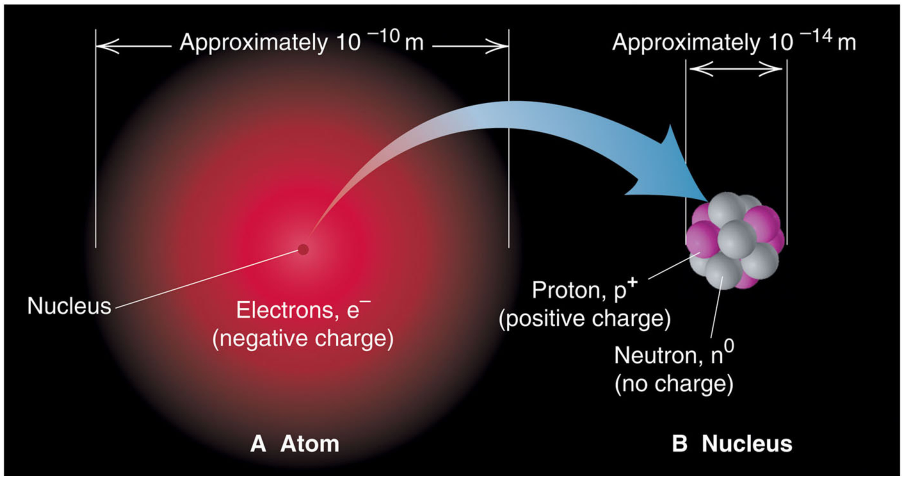
Rutherford (1908 Nobel Prize in Chemistry)



α particle velocity ~ 1.4 x 10^7 m/s
(~5% speed of light)

1. atoms positive charge is concentrated in the nucleus
2. proton (p) has opposite (+) charge of electron (-)
3. mass of p is 1840 x mass of e^- (1.67 x 10^{-24} g)

Model of the Atom



Substances - Definitions

Element: **Substance built from atoms with the same proton number**
Number of neutrons can vary

Isotopes of Carbon

Element: C $\quad\quad ^{12}C$, ^{13}C

$\downarrow$

Substance built from atoms with the same proton number and a defined number of neutrons: **Nuclide**

Compound: **Substance built from atoms with different proton numbers.**

Element - Symbols

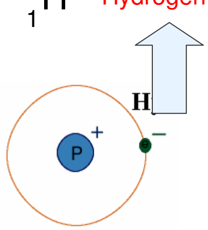
Characterisation of atoms/elements/compounds

Symbols consist of one to three letters; usually derived from the latin name of the element

H: Hydrogen (hydrogenium)	N: Nitrogen (nitrogenium)
C: Carbon (carbenium)	O: Oxygen (oxygenium)

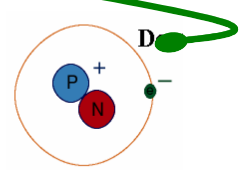
Isotopes have usually the same symbol. Exception: Hydrogen

$^{1}_{1}H$ Hydrogen $\quad\quad ^{2}_{1}D$ Deuterium $\quad\quad ^{3}_{1}T$ Tritium



Mass number $\longrightarrow$ m
Proton number (atom number) $\longrightarrow$ n

$^{m}_{n}E$



Element - Symbols

Information from the symbols

$^{12}_{6}C$ **12 – 6 = 6 neutrons** $^{16}_{8}O$ **16 – 8 = 8 neutrons**

Isotope nomenclature: $^{12}_{6}C$ **Called: Carbon-12 or C-12**

Isotope abundance:

$^{12}_{6}C$ 98.892% $^{13}_{6}C$ 1.108% $^{14}_{6}C$ traces

Generally the percentage contributions of heavier isotopes increases for the heavier elements

Masses of Atoms and Elementary Particles

1 atom of the carbon isotope $^{12}_{6}C$ weighs 12 u or $1.99264 \cdot 10^{-26}$ kg

that means $1u = 1.66053 \cdot 10^{-27}$ kg

Neutron: 1.00866 u; proton: 1.00727 u; electron $5.48593 \cdot 10^{-4}$ u

Masses of nuclides:

$^{2}_{1}D$ 2.014102 u $^{16}_{8}O$ 15.994915 u $^{19}_{9}F$ 18.998405 u

Atomic and molecular mass

average atomic mass: $$M_{Atom} = \frac{\Sigma(\text{nuclide mass} \cdot \text{abundance})}{100}$$

Example: Cl $$\overline{M}(Cl) = \frac{(34.9689 \cdot 75.53) + (36.9659 \cdot 24.47)}{100}$$

$$\overline{M}(Cl) = 35.453 \text{ u}$$

average molecular mass: $\overline{M}_{molecule}$

Example Cl_2 gas: $\overline{M}(Cl_2) = 35.453u \cdot 2 = 70.906$ u

Using relative atomic masses to count atoms – the mole

Counting unit on an atomic scale: The Mole *n* (symbol: mol; SI unit)

1 Mole of a substance contains as many particles (atoms, molecules, etc) as there are atoms in exactly 12 g of the nuclide 12-carbon (^{12}C).

The number N_A of particles in 1 Mole is $6.022 \cdot 10^{23}$!!!!

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1 mol Al contains $6.022 \cdot 10^{23}$ Al atoms.
1 mol H_2O contains $6.022 \cdot 10^{23}$ H_2O molecules.

Mass of 1 Mole of C $\rightarrow$ molar mass of C = M(C) = 12g/mol

Mass of 1 Mole of H_2O $\longrightarrow$ molar mass of H_2O =
$M(H_2O) = 18.015u \times 6.022 \cdot 10^{23} = 18.015$ g/mol

Formula to learn by heart !!!!:

$$\text{number of moles} = \frac{\text{grams of substance}}{\text{molar mass}}$$

$$\boxed{n = \frac{m}{M} \quad \left[\frac{g}{g/mol} = mol \right]}$$