# Last class

Matter – occupies space, has mass

- Composition – parts or components of matter
- Properties – qualities that distinguish matter
  - Physical – properties it exhibits with no change in composition
  - Chemical – way in which it can change from one type to another

Made up of atoms (will discuss this much further in Ch. 2)

Can be classified as "element" or "compound" (molecules)

Mixtures – more than one type of matter combined
- Homogeneous or heterogeneous
- Can be separated by physical methods

Pure substances – only one type of matter
- Homogenous
- Cannot be separated by physical methods

Three states of matter: Solid, liquid and gas

# Measurement of matter

Mass – amount of matter (Kg) - scale measures mass
Weight – force exerted by gravity on mass (g x m)
Time (s)
Temperature (K)

Seven "basic" properties and units

*Derived properties:* example: Volume ($m^3$, $cm^3$, L)

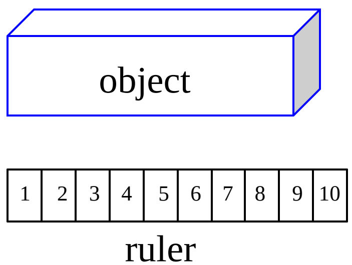
- Density = mass/volume ($g/cm^3$)

- % composition by mass = mass of component per 100g of substance

# Uncertainty in measurements

How good is your measurement of a property?

Property:
length




Ruler may not be well made:
*systematic error*

May not be able to get close enough to read the ruler well:
*random error*

1. Get the best possible ruler
2. Repeat the measurement several times and take the average

# Uncertainties in Scientific Measurements

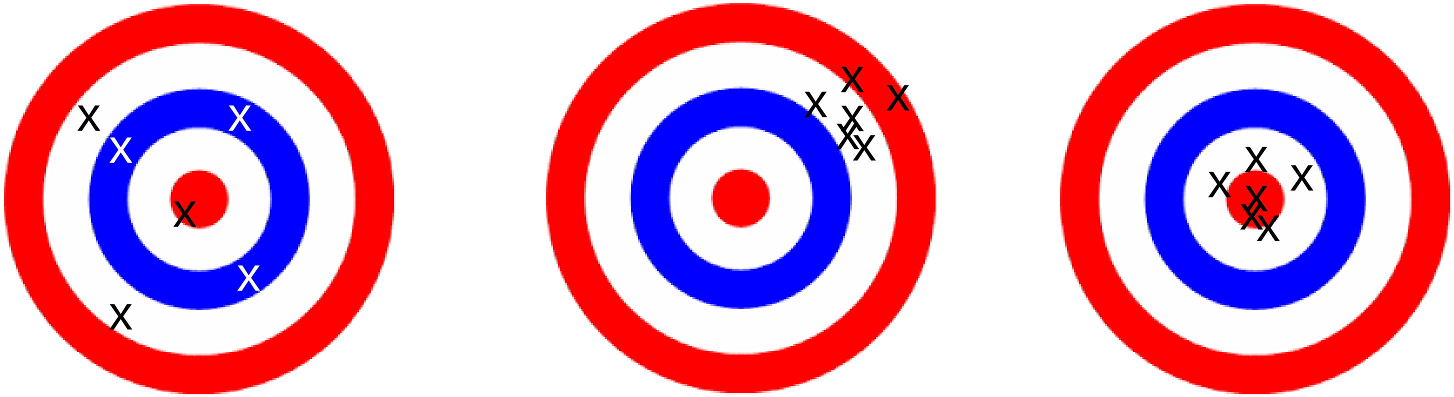
## *Precision*

Reproducibility of a measurement.

How close are multiple measurements to each other

## *Accuracy*

How close is the measured value to the real value.



Accuracy
Precision

# Precision


Meas. 1: 10.5
Meas. 2: 10.4
Meas. 3: 10.6


Meas. 1: 10.4979
Meas. 2: 10.4978
Meas. 3: 10.4977

| | | |
|---|---|---|
| Reproducibility | ~ 0.1 g | ~ 0.0001 g |
| Precision | low | high |

# Precision



Meas. 1: 10.5
Meas. 2: 10.4
Meas. 3: 10.6

Meas. 1: 10.4979
Meas. 2: 10.4978
Meas. 3: 10.4977

Average: 10.5 10.4978

Significant figures:

The significant figures (SF) are the numbers you know with certainty plus one which is an estimate

# Significant Figures

Rules:

1. All non-zero digits are significant
2. Zeros between non zero digits are significant

10.4979

3. Zeros preceding a decimal pt. are non significant

0.497

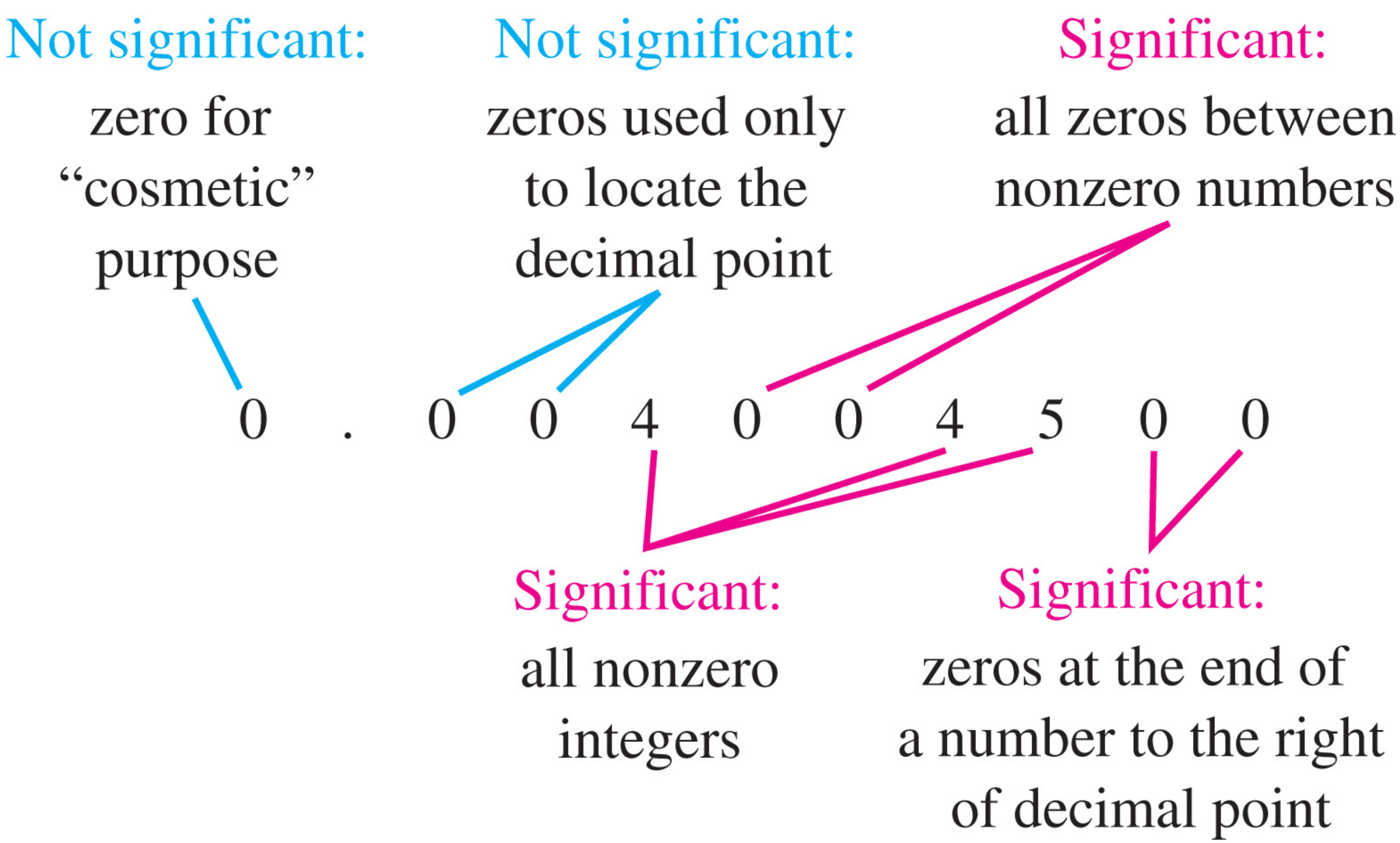
3. Zeros following a decimal pt. but before the first non zero digit are non significant

0.00497

4. Terminal zeros are ambiguous: need to write the decimal point

49700

# Significant Figures



# Significant figures (SF) in calculations

*Multiplication or division:*

Result can only have as many SF as the number with the least SF in the calculation

14.79 x 12.11 x 5.05 =

*Addition or subtraction*

Result must be expressed with the same number of digits after the decimal point as the quantity with the least number of digits after the decimal point

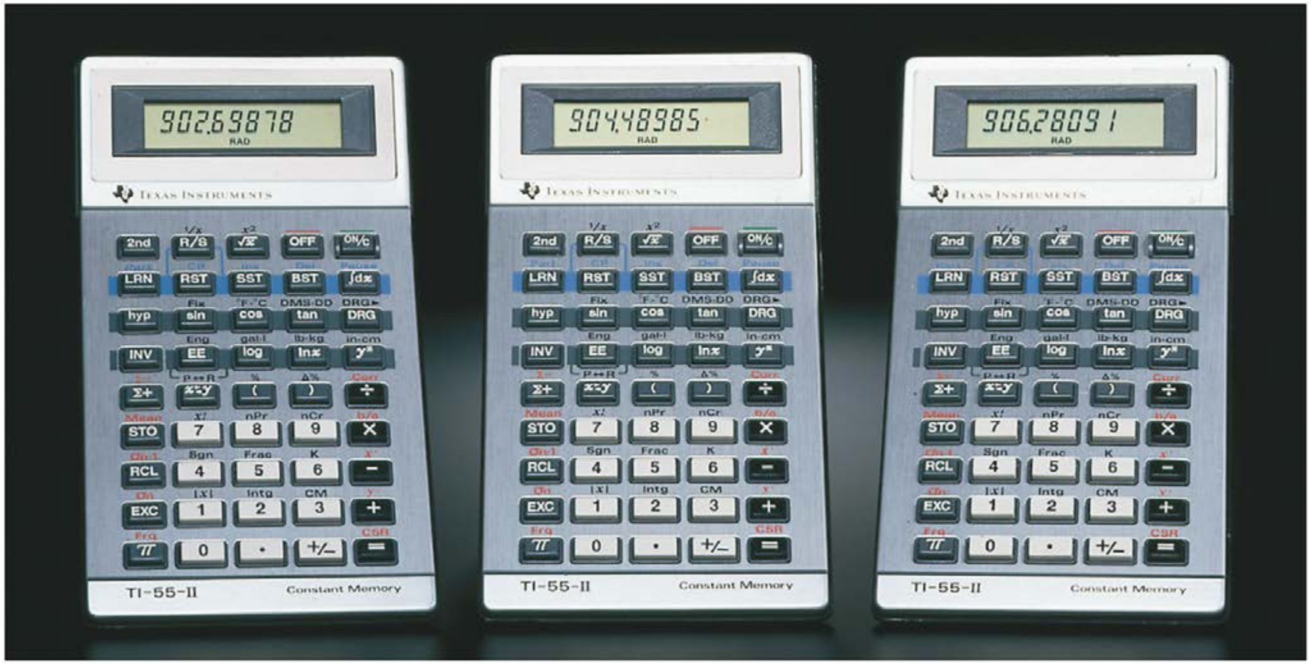
15.02 + 9986.0 + 3.538 =

**Need to "round off" to the right number of SF**

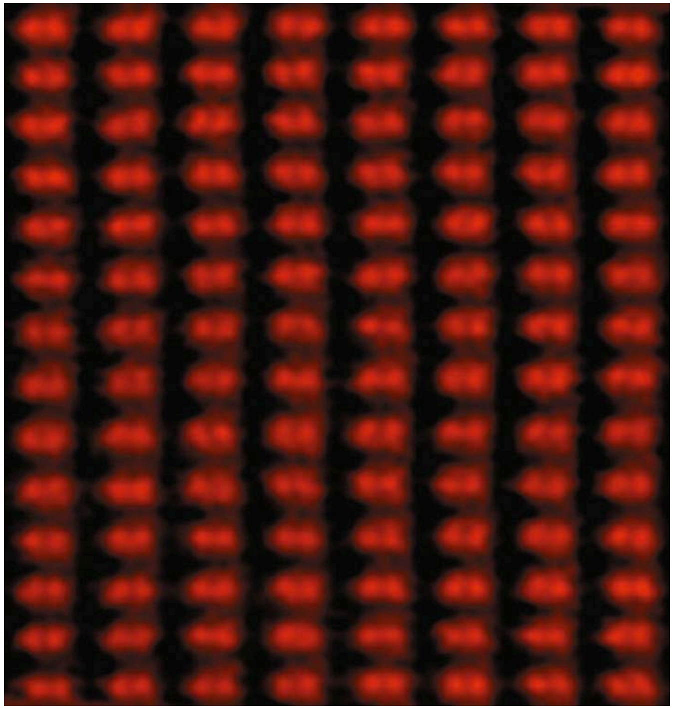
# Significant Figures

The calculators show the effect of the change in a low precision number (N) in a calculation $14.79 \times 12.11 \times N$

N = 5.04 5.05 5.06



# Chapter 2: Atoms and the Atomic Theory



CONTENTS


2-1 Early Chemical Discoveries and the Atomic Theory

2-2 Electrons and Other Discoveries in Atomic Physics

2-3 The Nuclear Atom

2-4 Chemical Elements

2-5 Atomic Mass

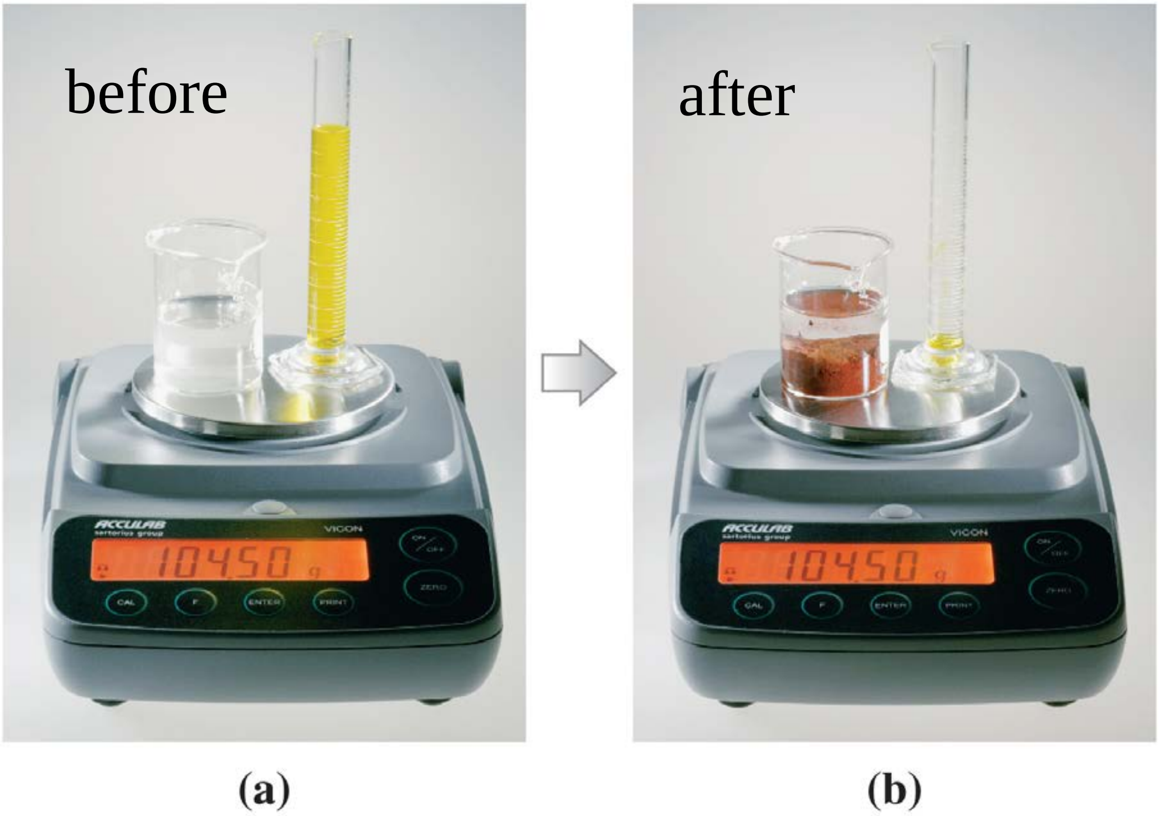
# Early Discoveries and the Atomic Theory

## Lavoisier 1774 Law of conservation of mass

Total mass of the substances present before a reaction are the same as those after the chemical reaction

Mass is not created or destroyed in a chemical reaction




Silver nitrate and potassium chromate ⇨ Silver chromate in a potassium nitrate soln.

▲ FIGURE 2-2

**Mass is conserved during a chemical reaction**

# Early Discoveries and the Atomic Theory

Lavoisier 1774 Law of conservation of mass

Proust 1799 Law of constant composition

All the samples of a compound have the same composition – the same proportions by mass of the constituent elements

Water always has 11.19% H and 88.81% O

Dalton 1803-1888 Atomic Theory

# Dalton's Atomic Theory

1. Each element is composed of small particles called **atoms**. Atoms are **neither created nor destroyed** in chemical reactions.

2. All atoms of a given element are **identical** and differ from all other elements

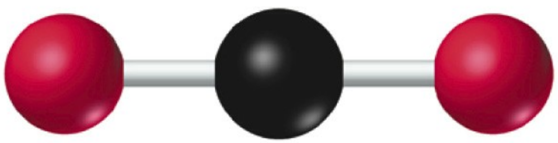
3. **Compounds** are formed when atoms of **more than one element** combine in simple numerical ratios.

Explained the Law of conservation of mass and the law of constant composition

Predicts the Law of multiple proportions

# Law of multiple proportions

If two elements form more than a single compound, the masses of one element combined with a fixed mass of the second are in the ratio of small whole numbers.


▲ Figure 2-3

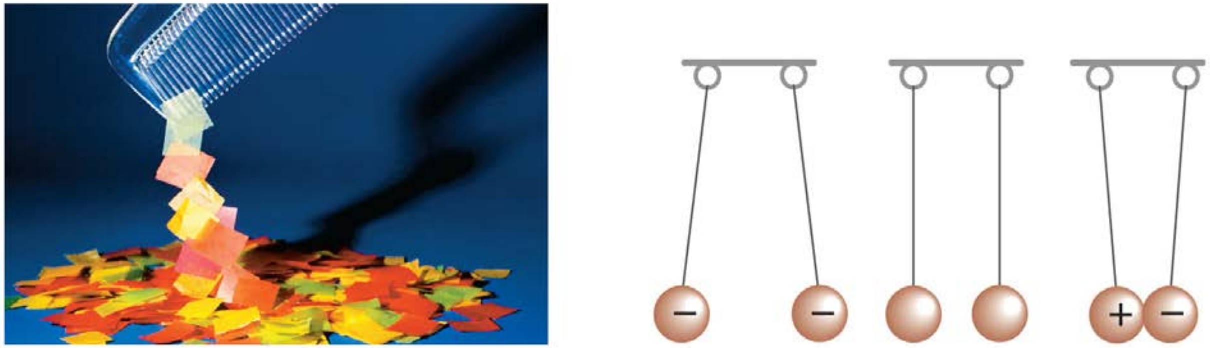
**Consequences of Dalton's theory**

- **In forming carbon monoxide, 1.0 g of carbon combines with 1.33 g of oxygen.**
- **In forming carbon dioxide, 1.0 g of carbon combines with 2.66 g of oxygen.**

# Nature of Atoms (Atomic Theory)

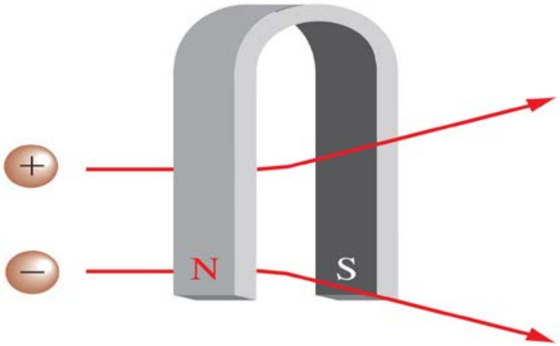
Two phenomena that were being investigated in physics were used in experiments that led to the discovery that atoms are complex units, made up of other particles

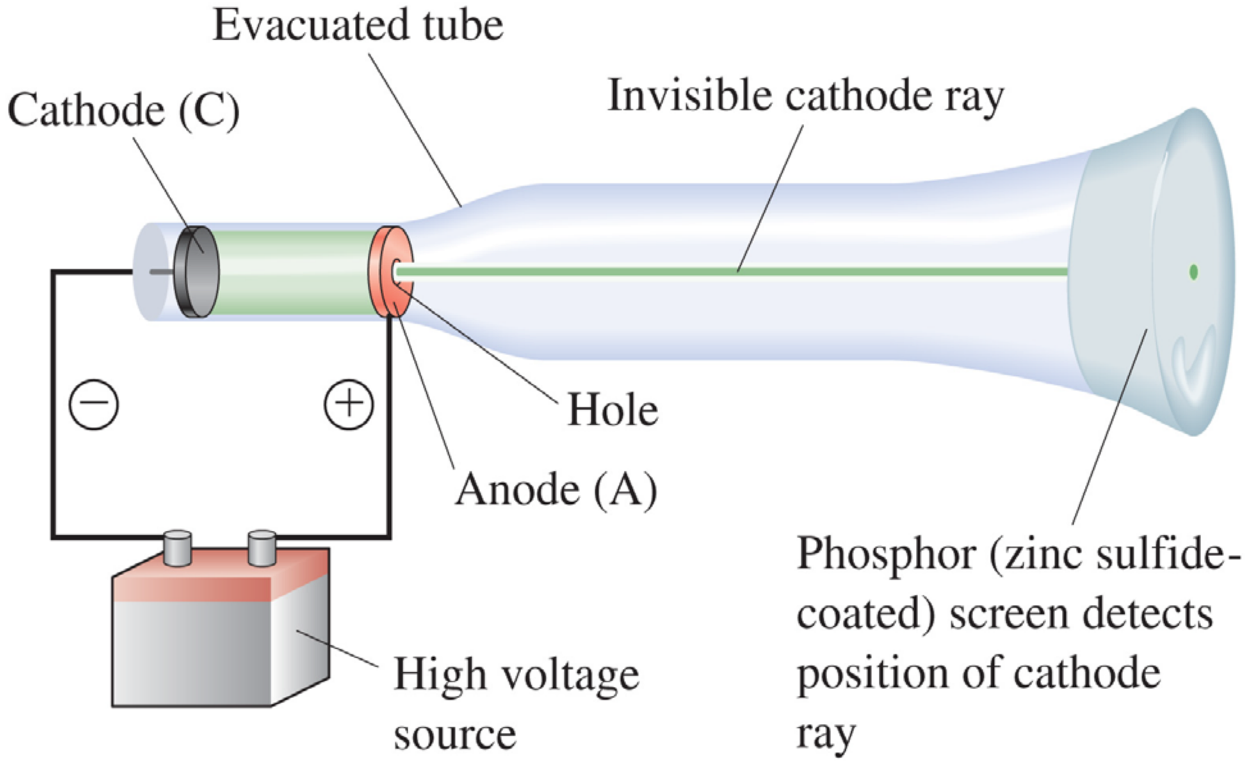
Electricity and charge



(a) (b)

Magnetism



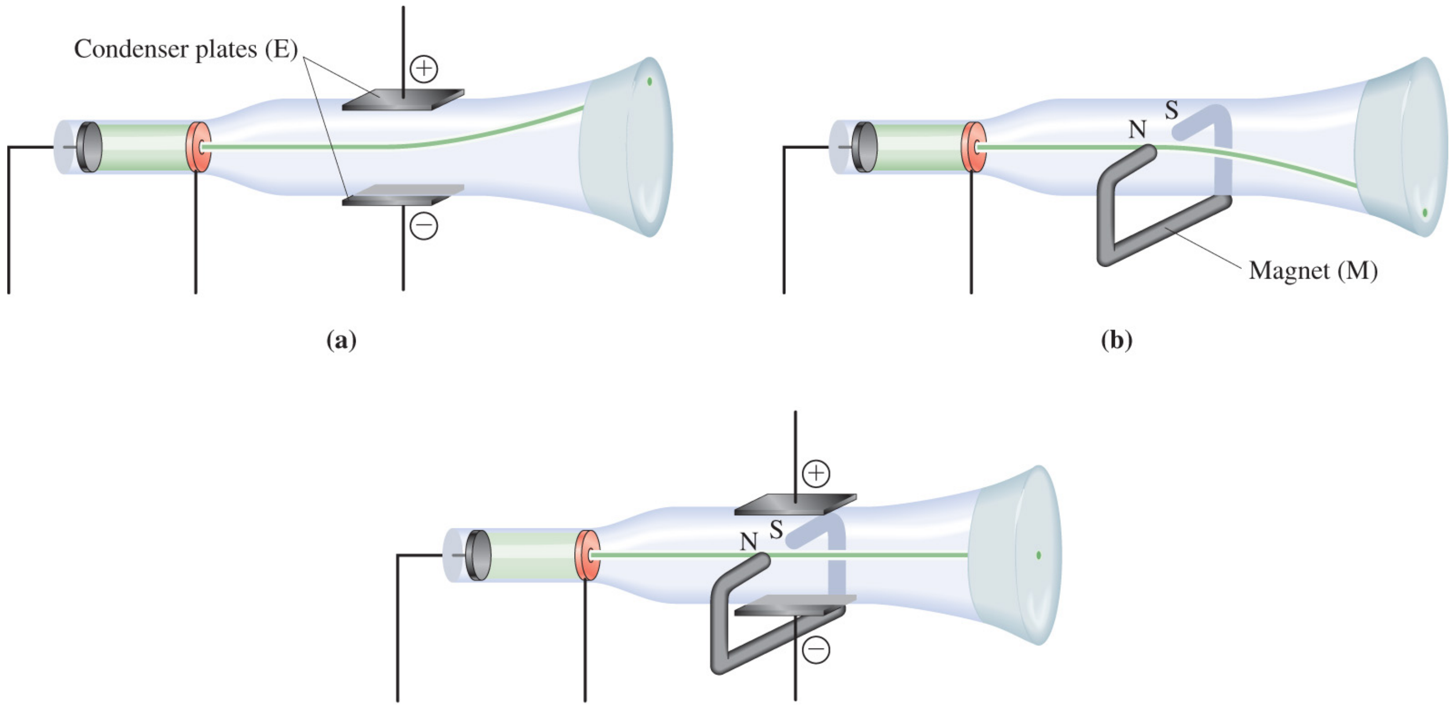



Cathode rays produce a spot on the phosphor: radiation from the cathode

Cathode Ray Tube (CRT) – Michael Faraday

# The radiation is deflected by magnetic fields and charge



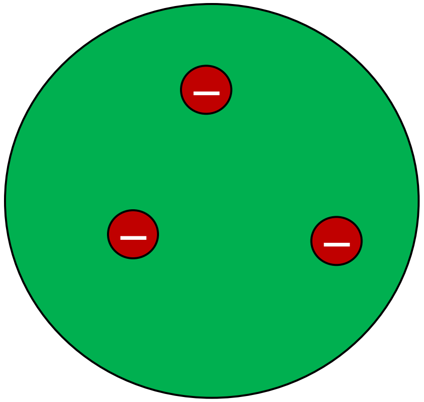

(a) (b) (c)

$$\text{Electron m/e} = -5.6857 \times 10^{-9}\ \text{g coulomb}^{-1}$$

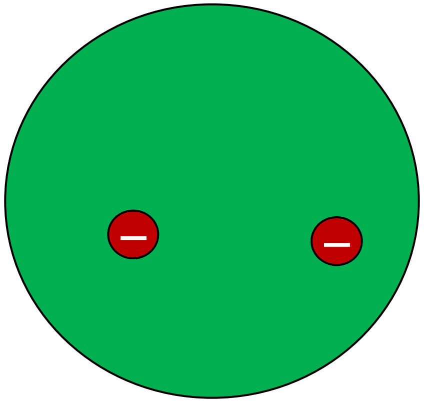
J. J. Thompson

# Early concept of the atom

"Plum pudding" model – J. J. Thompson




Lithium atom




Lithium (+) ion

# X-Rays and Radioactivity

While studying CRTs and electrons, other types of radiation (and radioactivity) were also discovered and investigated

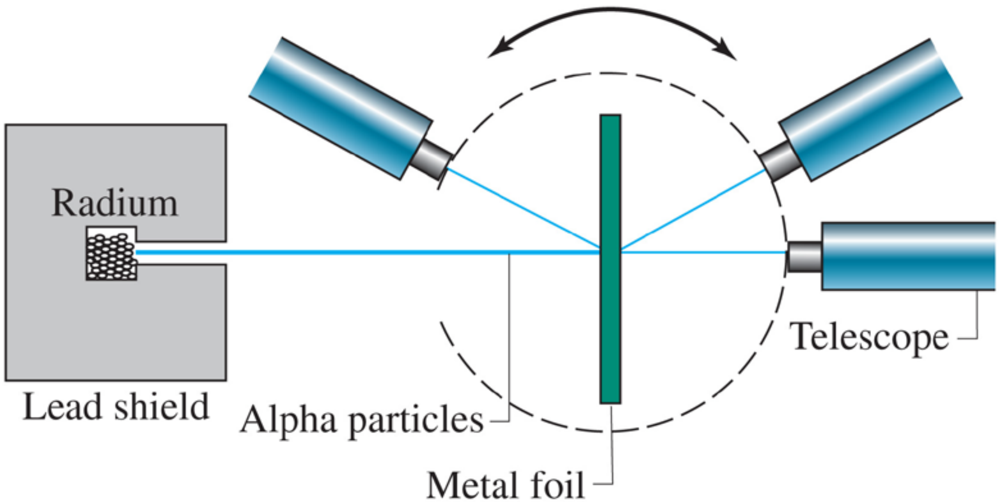
**Radioactivity is the spontaneous emission of radiation from a substance.**

- X-rays and $\gamma$-rays are high-energy light.
- $\alpha$-particles are a stream of helium nuclei, $He^{2+}$.
- $\beta$-particles are a stream of high speed electrons that originate in the nucleus.

# The Nuclear Atom

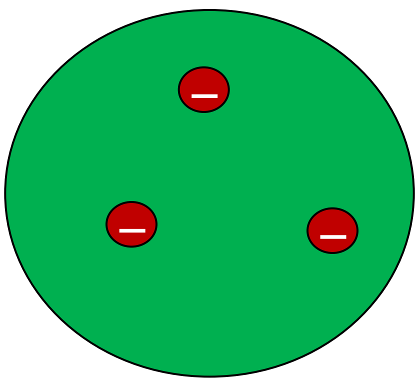
## Geiger and Rutherford 1909

$\alpha$-particles: $He^{2+}$




Observations:

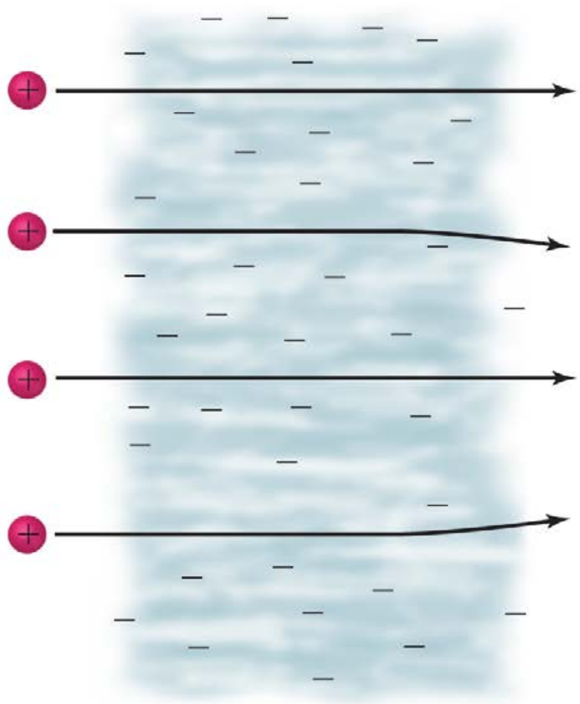
1. Most particles were undeflected
2. Some experienced a slight deflection
3. A few experienced a LARGE deflection
4. A few “BOUNCED” back



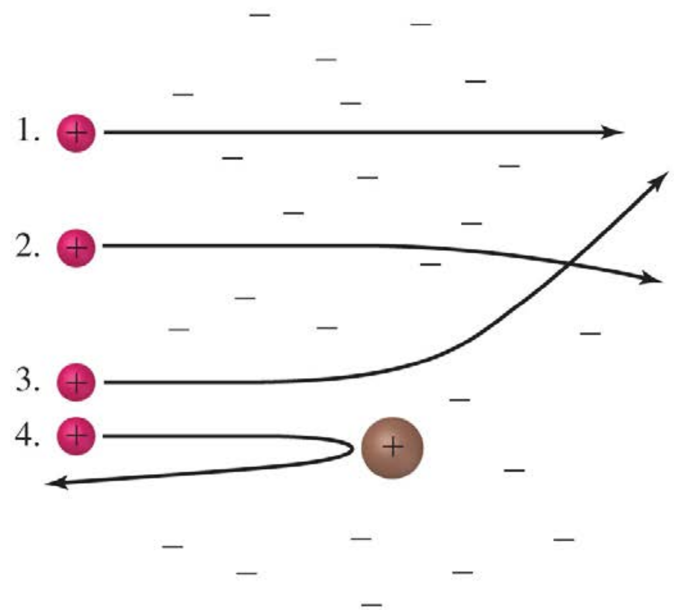

Plum pudding model

**The scattering of alpha particles by metal foil**

# Conclusions

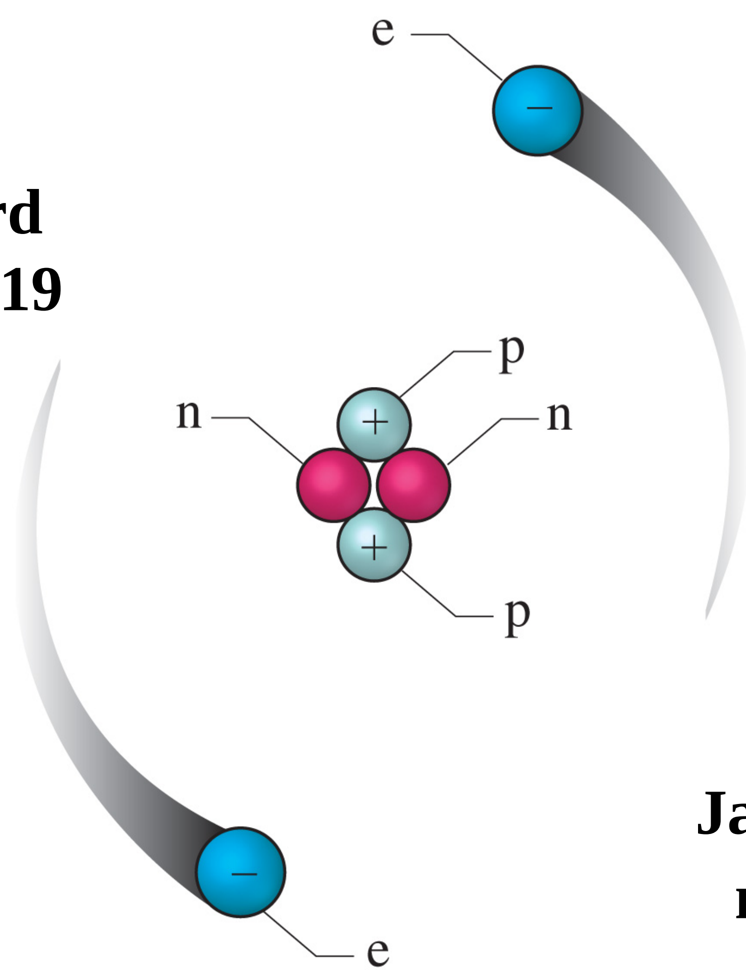


- Most of the mass and all of the positive charge is concentrated in a small region called the nucleus .




- There are as many electrons outside the nucleus as there are units of positive charge on the nucleus

**Rutherford**
**protons 1919**




- Magnitude of the positive charge is different for all elements and ~1/2 of the atomic mass

**James Chadwick**
**neutrons 1932**

▲ Figure 2-13

**The nuclear atom – illustrated by the helium atom**

# Mass of the Atom

Units: atomic mass units (amu) (SI symbol ‘u’ )

1 amu = 1/12 (mass of the atom of Carbon -12)

**TABLE 2.1 Properties of Three Fundamental Particles**

| | Electric Charge | | Mass | |
|---|---|---|---|---|
| | SI (C) | Atomic | SI (g) | Atomic (u)[a] |
| Proton | $+1.6022 \times 10^{-19}$ | +1 | $1.6726 \times 10^{-24}$ | 1.0073 |
| Neutron | 0 | 0 | $1.6749 \times 10^{-24}$ | 1.0087 |
| Electron | $-1.6022 \times 10^{-19}$ | −1 | $9.1094 \times 10^{-28}$ | 0.00054858 |

[a]u is the SI symbol for atomic mass unit (abbreviated as amu).

# Scale of Atoms

**The heaviest atom has a mass of only 4.8 x $10^{-22}$ g and a diameter of only 5 x $10^{-10}$ m.**

**<u>Useful units:</u>**

- **1 amu (atomic mass unit) = 1.66054 x $10^{-24}$ kg**
- **1 pm (picometer) = 1 x $10^{-12}$ m**
- **1 Å (Angstrom) = 1 x $10^{-10}$ m = 100 pm = 1 x $10^{-8}$ cm**

# Chemical Elements

**∗To represent a particular atom we use symbolism:**

number p + number n

$${}^{A}_{Z}\mathrm{E} \longleftarrow \text{symbol of the element}$$

number p

A= mass number Z = atomic number

number p + number n

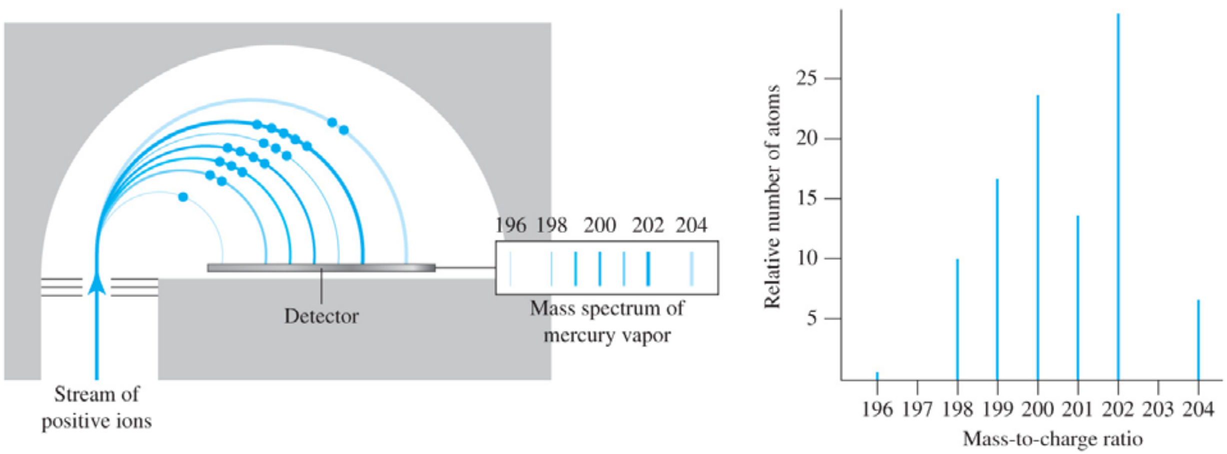
$${}^{A}_{Z}\mathrm{E}^{\pm ?} \longleftarrow \text{number p} - \text{number e}$$

number p

* Ion has net charge: p is not equal to e

# Isotopes

- Not all atoms of an element have the same mass
- They have the same Atomic number (or number of protons Z)
- Some atoms may have different numbers of neutrons, leading to different mass (isotopes).




▲ Figure 2-14

**A mass spectrometer and a mass spectrum**

# Atomic Mass

Weighted Average
Atomic Mass of an Element

Equation (2.3)

= fractional abundance of isotope 1 x atomic mass of isotope 1 + fractional abundance of isotope 2 x atomic mass of isotope 2 + ……

$$A_{ave} = \xi_1 \times A_1 + \xi_2 \times A_2 + \ldots\ldots \xi_n \times A_n$$

where $\xi_1 + \xi_2 + \square..+ \xi_n = 1.0$