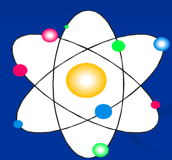
Atomic Structure



Section 4.1 Defining the Atom

- OBJECTIVES:
  - Describe Democritus's ideas about atoms.

Section 4.1 Defining the Atom

- OBJECTIVES:
  - Explain Dalton's atomic theory.

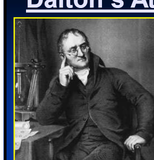
Section 4.1 Defining the Atom

- OBJECTIVES:
  - Identify what instrument is used to observe individual atoms.

Section 4.1 Defining the Atom

- The Greek philosopher Democritus (460 B.C. – 370 B.C.) was among the first to suggest the existence of atoms (from the Greek word "atomos")
  - He believed that atoms were *indivisible* and *indestructible*

Dalton's Atomic Theory (experiment based!)



John Dalton (1766 – 1844)

1) All elements are composed of tiny indivisible particles called atoms
2) Atoms of the same element are identical. Atoms of any one element are different from those of any other element.
3) Atoms of different elements combine in simple whole-number ratios to form chemical compounds
4) In chemical reactions, atoms are combined, separated, or rearranged – but never changed into atoms of another element.

Sizing up the Atom

- Elements are able to be subdivided into smaller and smaller particles – these are the *atoms*, and they still have properties of that element
  - If you could line up 100,000,000 copper atoms in a single file, they would be approximately *1 cm long*
  - Despite their small size, individual atoms **are** observable with instruments such as *scanning tunneling (electron) microscopes*

Section 4.2 Structure of the Nuclear Atom

- OBJECTIVES:
  - Identify three types of subatomic particles.

Section 4.2 Structure of the Nuclear Atom

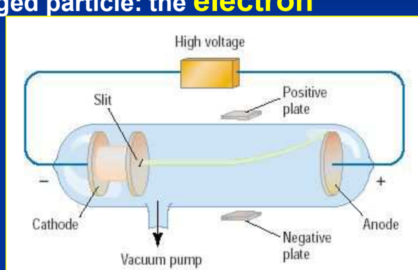
- OBJECTIVES:
  - Describe the structure of atoms, according to the Rutherford atomic model.

Section 4.2 Structure of the Nuclear Atom

- One change to Dalton's atomic theory is that **atoms are divisible** into subatomic particles:
  - Electrons, protons, and neutrons are examples of these fundamental particles
  - There are many other types of particles, but we will study these three

Discovery of the Electron

In 1897, J.J. Thomson used a *cathode ray tube* to deduce the presence of a negatively charged particle: the electron



Modern Cathode Ray Tubes

Television

Computer Monitor

- **Cathode ray tubes pass electricity through a gas that is contained at a very low pressure.**

Mass of the Electron

The oil drop apparatus

Mass of the electron is 9.11×10^{-28} g

1916 – Robert Millikan determines the mass of the electron: 1/1840 the mass of a hydrogen atom; has one unit of negative charge

Conclusions from the Study of the Electron:

a) Cathode rays have identical properties regardless of the element used to produce them. All elements must contain identically charged electrons.

b) Atoms are neutral, so there must be <u>positive particles</u> in the atom to balance the negative charge of the electrons

c) <u>Electrons have so little mass</u> that atoms must contain other particles that account for most of the mass

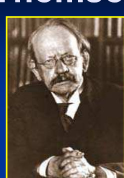
Conclusions from the Study of the Electron:

- **Eugen Goldstein** in 1886 observed what is now called the "<u>proton</u>" - particles with a positive charge, and a relative mass of 1 (or 1840 times that of an electron)
- 1932 – **James Chadwick** confirmed the existence of the "<u>neutron</u>" – a particle with no charge, but a mass nearly equal to a proton

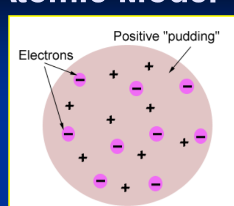
Subatomic Particles

Particle	Charge	Mass (g)	Location
Electron (e^-)	-1	9.11×10^{-28}	Electron cloud
Proton (p^+)	+1	1.67×10^{-24}	Nucleus
Neutron (n^o)	0	1.67×10^{-24}	Nucleus

Thomson's Atomic Model

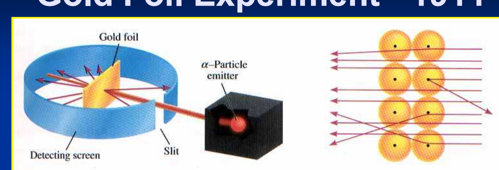


J. J. Thomson



Thomson believed that the electrons were like plums embedded in a positively charged "pudding," thus it was called the "**plum pudding**" model.

Ernest Rutherford's Gold Foil Experiment - 1911

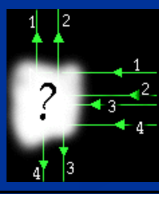


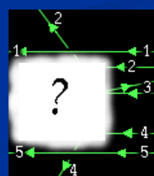
- Alpha particles are helium nuclei - The alpha particles were fired at a thin sheet of gold foil
- Particles that hit on the detecting screen (film) are recorded

Rutherford's problem:

In the following pictures, there is a target hidden by a cloud. To figure out the shape of the target, we shot some beams into the cloud and recorded where the beams came out. Can you figure out the <u>shape of the target</u>?

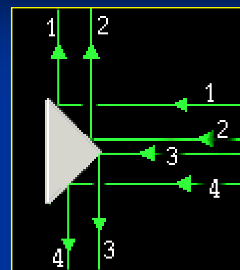
Target #1



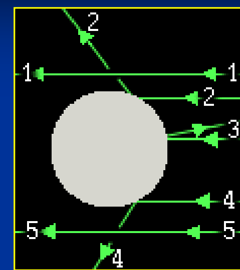


Target #2

The Answers:



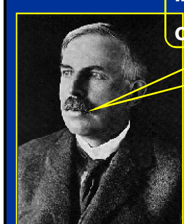
Target #1



Target #2

Rutherford's Findings

- **Most of the particles passed right through**
- **A few particles were deflected**
- **VERY FEW were greatly deflected**



"Like howitzer shells bouncing off of tissue paper!"

<u>Conclusions:</u>

a) The nucleus is small
b) The nucleus is dense
c) The nucleus is positively charged

The Rutherford Atomic Model

- Based on his experimental evidence:
  - The atom is mostly empty space
  - All the positive charge, and almost all the mass is concentrated in a small area in the center. He called this a "nucleus"
  - The nucleus is composed of protons and neutrons (they *make* the nucleus!)
  - The electrons distributed around the nucleus, and occupy most of the volume
  - His model was called a "nuclear model"

Section 4.3 Distinguishing Among Atoms

- OBJECTIVES:
  - <u>Explain</u> what makes *elements* and *isotopes* different from each other.

Section 4.3 Distinguishing Among Atoms

- OBJECTIVES:
  - <u>Calculate</u> the number of neutrons in an atom.

Section 4.3
Distinguishing Among Atoms

- OBJECTIVES:
  - Calculate the atomic mass of an element.

Section 4.3
Distinguishing Among Atoms

- OBJECTIVES:
  - Explain why chemists use the periodic table.

Atomic Number

- Atoms are composed of ***identical*** protons, neutrons, and electrons
  - How then are atoms of one element different from another element?
- Elements are different because they contain different numbers of PROTONS
- The "**atomic number**" of an element is the number of protons in the nucleus
- # protons in an atom = # electrons

Atomic Number

Atomic number (Z) of an element is the number of protons in the nucleus of each atom of that element.

Element	# of protons	Atomic # (Z)
Carbon	6	6
Phosphorus	15	15
Gold	79	79

Mass Number

Mass number is the number of protons and neutrons in the nucleus of an isotope: Mass # = p^+ + n^0

Nuclide	p^+	n^0	e^-	Mass #
Oxygen - 18	8	10	8	18
Arsenic - 75	33	42	33	75
Phosphorus - 31	15	16	15	31

Complete Symbols

- Contain the symbol of the element, the mass number and the atomic number.

Superscript → Mass number

Subscript → Atomic number

X

Symbols

- Find each of these:
  - a) number of protons
  - b) number of neutrons
  - c) number of electrons
  - d) Atomic number
  - e) Mass Number

$^{80}_{35}Br$

Symbols

- If an element has an atomic number of 34 and a mass number of 78, what is the:
  - a) number of protons
  - b) number of neutrons
  - c) number of electrons
  - d) complete symbol

Symbols

- If an element has 91 protons and 140 neutrons what is the
  - a) Atomic number
  - b) Mass number
  - c) number of electrons
  - d) complete symbol

Symbols

- If an element has 78 electrons and 117 neutrons what is the
  - a) Atomic number
  - b) Mass number
  - c) number of protons
  - d) complete symbol

Isotopes

- Dalton was wrong about all elements of the same type being identical
- Atoms of the same element *can* have different numbers of neutrons.
- Thus, different mass numbers.
- These are called **isotopes**.

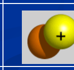
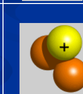
Isotopes

- Frederick Soddy (1877-1956) proposed the idea of isotopes in 1912
- **Isotopes are atoms of the same element having *different masses*, due to varying numbers of neutrons.**
- Soddy won the Nobel Prize in Chemistry in 1921 for his work with isotopes and radioactive materials.

Naming Isotopes

- We can also put the mass number *after* the name of the element:
  - carbon-12
  - carbon-14
  - uranium-235

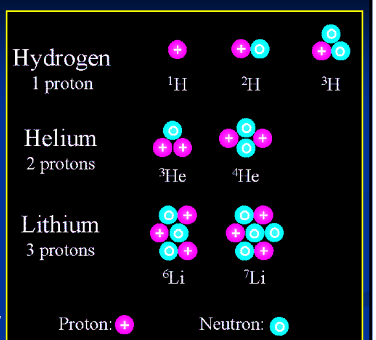
Isotopes are atoms of the same element having *different masses*, due to varying numbers of neutrons.

Isotope	Protons	Electrons	Neutrons	Nucleus
Hydrogen–1 (protium)	1	1	0	
Hydrogen-2 (deuterium)	1	1	1	
Hydrogen-3 (tritium)	1	1	2	

Isotopes

Elements occur in nature as *mixtures* of isotopes.

Isotopes are atoms of the same element that differ in the *number of neutrons.*



Atomic Mass

- How heavy is an atom of oxygen?
  - It depends, because there are different *kinds* of oxygen atoms.
- We are more concerned with the *average atomic mass.*
- This is based on the abundance (percentage) of each variety of that element in nature.
  - We don't use grams for this mass because the numbers would be too small.

Measuring Atomic Mass

- Instead of grams, the unit we use is the Atomic Mass Unit (amu)
- It is defined as one-twelfth the mass of a carbon-12 atom.
  - Carbon-12 chosen because of its isotope purity.
- Each isotope has its own atomic mass, thus we determine the average from percent abundance.

To calculate the average:

- Multiply the atomic mass of each isotope by it's abundance (expressed as a decimal), then add the results.
- If not told otherwise, the mass of the isotope is expressed in atomic mass units (amu)

Atomic Masses

Atomic mass is the average of all the naturally occurring isotopes of that element.

Isotope	Symbol	Composition of the nucleus	% in nature
Carbon-12	^{12}C	6 protons 6 neutrons	98.89%
Carbon-13	^{13}C	6 protons 7 neutrons	1.11%
Carbon-14	^{14}C	6 protons 8 neutrons	<0.01%

Carbon = 12.011

SAMPLE PROBLEM 4.2 - Page 117

Calculating Atomic Mass

Element X has two natural isotopes. The isotope with a mass of 10.012 amu (^{10}X) has a relative abundance of 19.91%. The isotope with a mass of 11.009 amu (^{11}X) has a relative abundance of 80.09%. Calculate the atomic mass of this element. ← **Question**

1 Analyze *List the knowns and the unknown.*

Knowns
- isotope ^{10}X:
  mass = 10.012 amu
  relative abundance = 19.91% = 0.1991
- isotope ^{11}X:
  mass = 11.009 amu
  relative abundance = 80.09% = 0.8009

Unknown
- atomic mass of element X = ?

The mass each isotope contributes to the element's atomic mass can be calculated by multiplying the isotope's mass by its relative abundance. The atomic mass of the element is the sum of these products.

} **Knowns and Unknown**

2 Calculate *Solve for the unknown.*

for ^{10}X: 10.012 amu × 0.1991 = 1.993 amu
for ^{11}X: 11.009 amu × 0.8009 = 8.817 amu ← **Solution**
for element X: atomic mass = 10.810 amu ← **Answer**

The Periodic Table: A Preview

- **A "*periodic table*" is an <u>arrangement of elements</u> in which the elements are separated into groups based on a set of repeating properties**
  - **The periodic table allows you to easily compare the properties of one element to another**

The Periodic Table: A Preview

- **Each *horizontal row* (there are 7 of them) is called a period**
- **Each *vertical column* is called a group, or family**
  - **Elements in a group have similar chemical and physical properties**
  - **Identified with a number and either an "A" or "B"**
  - **More presented in Chapter 6**

End of Chapter 4