CHAPTER 4
Atomic Structure

4.1 Early Theories of Matter

- Earth, Water, Air, Fire
- Matter was thought to be infinitely divisible
- No method was available to test theories

Democritus (460 B.C. - 370 B.C.)

- First suggested the idea of atoms
- Proposed that matter was not infinitely divisible
- Atoms could not be created, divided, or destroyed
- Different atoms had different sizes and shapes
- Atoms have different properties bases on their sizes and shapes.
- Could not explain what held atoms together

Aristotle (384 B.C. - 322 B.C.)

- Philosopher with his own ideas of nature
- Disagreed with Democritus because his theories could not be proven.
- The existence of atoms was denied for nearly two thousand years.

Dalton (1766 - 1844)

- Revived Democritus's ideas
- Experiments could now be performed to refine and verify theories made by Democritus
- Began the development of the atomic theory
- A huge step toward the current model, but his entire theory was not accurate.

Dalton's Atomic Theory

1. Matter is composed of small particles called atoms
2. Atoms of an element are identical and different from the atoms of other elements
3. Atoms are **not divisible** and cannot be created or destroyed
4. Atoms combine in whole-number ratios to form compounds
5. In chemical reactions, atoms can be separated, combined, or rearranged as long as the atom remains unchanged.

What was wrong with Dalton's Theory?

1. **All atoms of a given element are not identical, is is possible for a given element to have different properties based on varying masses *(different isotopes)***

2. **Atoms are divisible *(subatomic particles)***

Defining the Atom

- **The smallest particle of an element that retains the properties of that element.**
- **Have a size ranging from: (5 x 10^{-11}m) – (2 x 10^{-10}m)**
- **A solid copper penny would have: 2.9 x 10^{22} atoms**
- **6 billion copper atoms lined up is less than 1 meter**

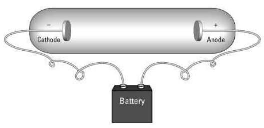
4.2 Subatomic Particles and the Nuclear Atom

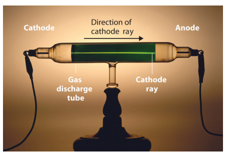
- **Discovering subatomic particles was an accidental discovery.**
- **The actual protons, electrons, and neutrons are not unique and are the same in every atom.**
- **The number of protons, electrons, and neutrons is what makes each atom unique.**

Discovering the Electron

- **Understanding the relationship between matter and electric charge was very important**
- **The invention of the vacuum pump was required.**
- **Sir William Crookes discovered the Cathode Ray**

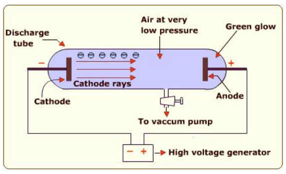
CRT





Cathode Ray Tube and Electrons

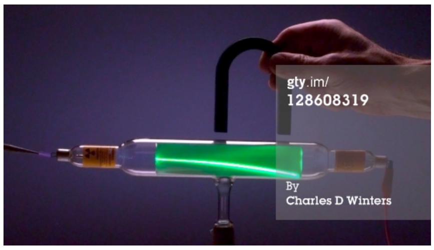
- **The cathode ray was a stream of charged particles with a negative charege**



J.J. Thomson and Electrons

- Discovered the mass of an electron and disproved part of Dalton's theory that atoms are indivisible
- The lightest atom was hydrogen, and the mass of the particle was far less.
- He identified the first subatomic particle

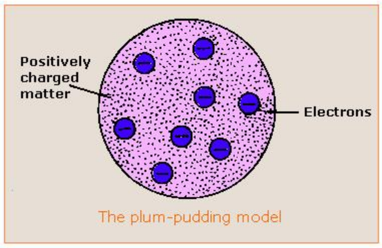
Electrons being repelled by magnet...



Millikan and Electrons

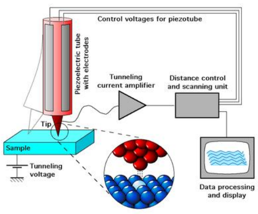
- Determined the charge of an electron was 1-.
- Also determined the mass of an electron occupied 1/1840 of a hydrogen atom.
- Led scientist to think about what occupies the remaining mass of an atom.
- Thomson developed the Plum Pudding Model.

Notice
This model has electrons floating in evenly distributed positive matter

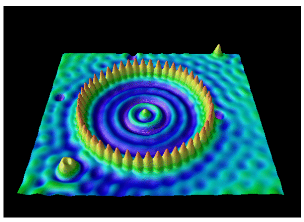


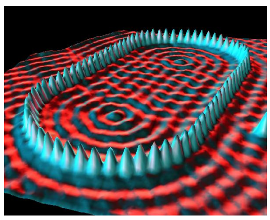
Scanning Tunneling Microscope

Visualizes individual atoms and makes it possible to see them as a 3D image on a monitor.



Atoms of Copper and Iron





Electrons

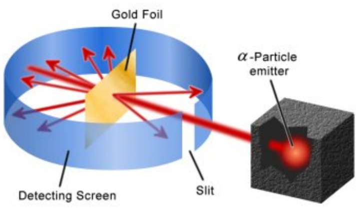
- Negatively charged subatomic particle
- 1/1840 the mass of 1 proton
- Millikan: discovered that anelectron carries one unit of negative charge

Rutherford Model *(1871 - 1937)*

- **Known as the nuclear atom.**
- **Concluded the plum pudding model was wrong**
- **The Gold Foil experiment proved that a dense region existed within the atom and that electrons move through empty space.**
- **A dense positive region in atoms must be defelcting the positive beam of particles**

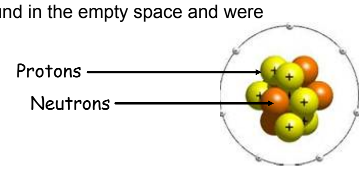
Rutherford's Gold Foil Experiment

Rutherford expected most (+) particles to travel straight through the gold atoms, however, some particles were deflected backwards due to a dense region



The Nuclear Atom

- A tiny dense region centrally located in an atom
- Contained the atom's positive charge and virtually all of the atom's mass.
- The electrons floated around in the empty space and were attracted to the nucleus.



Size and Density of Atoms

If an atom had the diameter of 2 football fields, the nucleus would be a nickel in the middle.

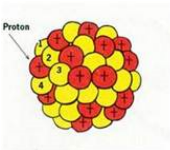
If the nucleus were the size of a (.) on a page, it would have the mass of 70 cars.

Discovering Protons & Neutrons

- Rutherford proved that atoms are neutral and that most of the mass was in the nucleus, however, it took another 20 years to determine where all the mass of an atom was coming from.
- Rutherfod determined there were positively charged particles that made an atom neutral in charge.
- Rutherford's coworker, James Chadwick, showed that a neutral particle was located in the nucleus.

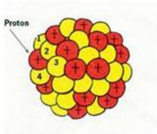
Protons

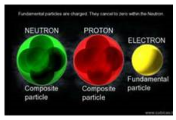
- Positively charged subatomic particles
- Carries a one positive (+) charge
- Located in the nucelus.

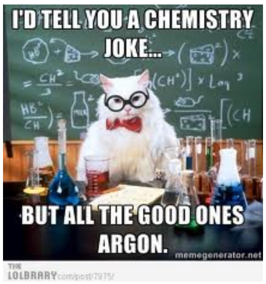


Neutrons

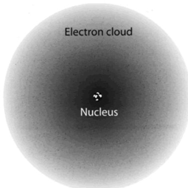
- **Subatomic particle with no electrical charge**
- **Discovered by Chadwick in 1932**
- **Mass equals that of a proton**







The electron cloud consists of fast moving electrons held within the atom by their attraction to the protons



3 parts of an atom

1. Protons
2. Electrons
3. Neutrons

Summary of the Nucleus

- 99.97% of the mass of an atom
- Contains Protons and Neutrons
- Overall positive charge
- Very little volume of the atom
- Proton Number defines the atom

Summary of the Electron Cloud

- 0% of the mass of an atom
- Contains electrons only
- Overall negative charge
- Most of the volume of the atom
- Determines the chemical properties of an atom

4.3 How do atoms differ?

- There are 110 elements, which means there are 110 different types of atoms.
- How are atoms of each element different?
- The atomic number of each atom is different

Atomic Number

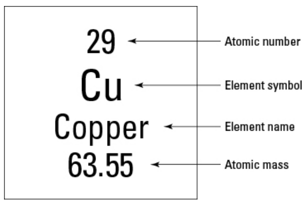
- Differences among elements result from differences in the numbers of protons in their atoms.
- Equal to the number of protons in an element
- The periodic table provides the atomic number of each element, therefore, the number of protons

Proton & Electron Number

- When atoms are electrically neutral, the proton number must equal the electron number.
- If you know the atomic number of an element, you know the number of protons and electrons in an atom

Since it's atomic number is 29, Copper has 29 protons and 29 electrons

What about the number of neutrons?



Mass Number

- The number of protons and neutrons in an atom
- Remember, electrons have virtually NO mass.
- **Protons + Neutrons = Mass Number**
- So which isotope do you use?

Bromine	Tin
Atomic Number =	Atomic Number =
Mass Number =	Mass Number =
Proton Number =	Proton Number =
Electron Number =	Electron Number=
Neutron Number=	Neutron Number =
Mass of Electrons =	Mass of Electrons =

Mass Number

- The number of protons and neutrons in an atom
- Remember, electrons have virtually NO mass.
- **Protons + Neutrons = Mass Number**
- So which isotope do you use?

Why is the atomic mass a decimal?

- The atomic mass is not the actual mass of any atom
- It is the weighted average of the different versions of a given element.
- There are different versions of each element, each version has a different mass.
- Each version exists in nature in different amounts.

Isotopes

- Different versions of an atom are called isotopes
- Each isotope of a given atom has the same number of protons, but a different number of neutrons
- They usually have the same chemical properties
- So which isotope do you use?

3 Isotopes of Oxygen

- **Oxygen-15 has 8 protons & 7 neutrons**
- **Oxygen-16 has 8 protons & 8 neutrons**
- **Oxygen-17 has 8 protons & 9 neutrons.**

$\boxed{15.999}$
Average

Which isotope is the most common?

3 Isotopes of Hydrogen

- Hydrogen-1 (one proton, no neutrons)
  - Protium
- Hydrogen-2 (one proton, one neutron)
  - Deuterium
- Hydrogen-3 (one proton, two neutrons)
  - Tritium

1.0078

Aluminum - 27	Palladium - 106
Atomic Number =	Atomic Number =
Mass Number =	Mass Number =
Proton Number =	Proton Number =
Electron Number =	Electron Number=
Neutron Number=	Neutron Number =
Mass of Electrons =	Mass of Electrons =

Shorthand Notation

- The atomic number is written as a subscript.
- The mass number is written as a superscript.

- Example: $^{9}_{4}Be$ $^{27}_{13}Al$

Atomic Mass Units

- The masses of atoms are too small to work with so we use Carbon as the comparative atom.
- An Atomic Mass Unit (amu) is defined as one-twelfth the mass of a carbon-12 atom.
- Proton=1-amu, Neutron=1-amu, electrons=0-amu

Atomic Mass vs. Mass Number

- **There is a difference!!**
- **Average Atomic Mass is the mass on the periodic Table**
- **Mass Number is the actual mass of a given atom.**

Calculating the Atomic Mass of an Element

You must know:
- The number of stable isotopes of that element
- The mass of each isotope
- The natural percent abundance of each isotope

Carbon-12 (91.7%)
Carbon-13 (6.9%)
Carbon-14 (1.4%)

Average Atomic Mass of Carbon

- Carbon-12: 12 x 0.917 = 11.004
- Carbon-13: 13 x 0.069 = 0.897
- Carbon-14: 14 x 0.014 = 0.196

Add together products together
12.097 (Avg. atomic mass)

Determining Average Atomic Mass of a "New" element

Isotope	Percentage	Show Work
Fontainium - 41	87.3 %	
Fontainium - 43	9.8 %	
Fontainium - 44	2.9 %	
Average Atomic Mass		

What if you aren't given the percentage?

You found 400 Germanium atoms:

- 260 atoms are Germanium-73
- 140 atoms are Germanium-72

What is the average atomic mass?

The weighted average mass reflects both the mass and the relative abundance of the isotopes as they occur in nature.

Chlorine-35	Isotopes	Chlorine-36
(54%)	Relative Abundance	(46%)

Write the following in shorthand notation...

Oxygen - 15		Isotope or Common
Oxygen - 16		Isotope or Common
Oxygen - 17		Isotope or Common

What can change in an atom?

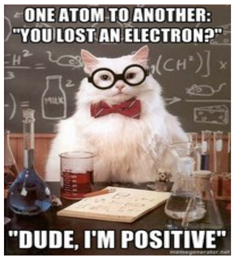
- Protons: can never change, or the atom changes
- Electrons: if the number changes, then an ion forms.
- Neutrons: if the number changes, then an isotope forms.

If an atom gains electrons, then….

- Then a negatively charged version of the same atom forms. P=8 E=10
- If oxygen gains 2 electrons, it has 10 electrons and 8 protons.
- It now has a "2-" charge

If an atom loses electrons, then….

- Then a positively charged version of the same atom forms. P=20 E=18
- If calcium lost 2 electrons, it has 18 electrons and 20 protons.
- It now has a "2+" charge.



Examples of Ions

	Sulfur	Magnesium	Lithium	Chlorine
Neutral Version	P=16	P=12	P=3	P=17
	E=16	E=12	E=3	E=17
Gain or Loss	Gain 2 electrons	Lose 2 electrons	Lose 1 electron	Gain 1 electron
New electron number	E=18	E=10	E=2	E=18
Charge	S^{2-}	Mg^{2+}	Li^{+}	Cl^{-}

Cation vs. Anion

- Cation: Positive Ion
  More protons than electrons
  I am "Positive" I love my Cat.
- Anion: Negative Ion
  More electrons than protons

	Atomic Number	Protons	Electrons	Neutrons	Mass Number
Al^{3+}					
Hydrogen – 1					
Hydrogen – 3					
H^{+}					
Oxygen – 16					
Oxygen – 14					
O^{2-}					

	Atomic Number	Protons	Electrons	Neutrons	Mass Number
Al^{3+}	13	13	10	14	27
Hydrogen – 1	1	1	1	0	1
Hydrogen – 3	1	1	1	2	3
H^{+}	1	1	0	0	1
Oxygen – 16	8	8	8	8	16
Oxygen – 14	8	8	8	6	14
O^{2-}	8	8	10	8	16

Neutral = Periodic Table Version

Element	Atomic #	Protons Number	Electron Number	Neutron Number	Mass Number	Ion, Isotope, or Common Version
Fluorine	9	9	9	10	19	
Chromium	24	24	24	32	56	
Calcium	20	20	18	20	40	
Iodine	53	53	54	74	127	
Polonium	84	84	84	131	215	

Element	Atomic #	Protons Number	Electron Number	Neutron Number	Mass Number	Ion, Isotope, Common
Fluorine	9	9	9	10	19	Neutral & Common
Chromium	24	24	24	32	56	Isotope
Calcium	20	20	18	20	40	Cation
Iodine	53	53	54	74	127	Anion
Polonium	84	84	84	131	215	Isotope

Isotopes vs. Allotropes

Isotopes: atoms of the same element with different numbers of neutrons O^{16}_{8} O^{17}_{8}
Allotropes: different forms of an element, usually when the same atoms are joined as compound.
The way atoms are arranged

Carbon exhibits both…
- Isotopes: Carbon-12, Carbon-13, Carbon-14
- Allotropes: graphite, diamond, and fullerenes

4.4 Unstable Nuclei & Radioactive Decay

Changes that can take place in the nucleus of an atom

Reactions in which atoms of an element change into atoms of another element.

Nuclear Reactions involve a change in the atom's nucleus

Radioactivity

When substances spontaneously emit radiation due to the nuclei of their atoms becoming unstable

Radiation is the rays and particles emitted by radioactive materials.

Radioactive elements undergo significant changes that can alter their identities Changing Proton Number

Radioactive Decay

Unstable atoms lose energy to gain stability.

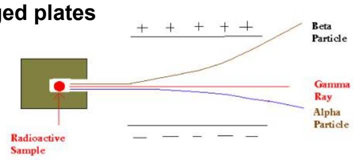
They lose energy by emitting radiation in a process that does not require energy. (Really long process)

Unstable radioactive atoms undergo radioactive decay until they become stable atoms.

Types of Radiation

3 different types of radiation are emitted during radioactive decay

This was determined by directing radiation between two electrically charged plates



Nuclear Equation

Shows the mass number and atomic number of the particles involved in radioactive decay.

Atomic and mass number must be conserved!

$$\boxed{{}^{222}_{88}\mathrm{Ra} \longrightarrow {}^{4}_{2}\mathrm{He} + {}^{218}_{86}\mathrm{Rn}}$$

Nucleon number or mass number

${}^{12}_{6}\mathrm{C}$ ${}^{238}_{92}\mathrm{U}$ Element symbol (C or U)

Proton number or atomic number

${}^{226}_{88}\mathrm{Ra} \rightarrow {}^{222}_{86}\mathrm{Rn} + {}^{4}_{2}\mathrm{He}$ Alpha emission (+)

${}^{89}_{38}\mathrm{Sr} \rightarrow {}^{89}_{39}\mathrm{Y} + {}^{0}_{-1}\mathrm{e}$ Beta emission (−)

${}^{60}_{27}\mathrm{Co} \rightarrow {}^{60}_{27}\mathrm{Co} + \boxed{{}^{0}_{0}\gamma}$ Gamma emission

Alpha Radiation

Radiation deflected toward negative plate

Made up of Alpha Particles, which contain 2 protons and 2 neutrons, therefore a 2+ charge.

It was attracted to the negative plate and repelled by the positive plate due to its positive charge.

He is the alpha particle.

Alpha

$${}^{238}_{92}\mathrm{U} \longrightarrow {}^{234}_{90}\mathrm{Th} + \boxed{{}^{4}_{2}\mathrm{He}}$$

Beta Radiation

Radiation deflected toward positive plate and repelled by the negative plate due to it's negative charge.

Made up of Beta Particles, which is an electron with a 1- charge.

A neutron is converted into a proton, hence, then increase in atomic number.

$${}^{14}_{6}\mathrm{C} \longrightarrow {}^{14}_{7}\mathrm{N} + {}^{0}_{-1}\beta$$

P = 6, N = 8   P = 7, N = 7

Gamma Radiation

Radiation not deflected toward either plate

High energy radiation with no mass and no electrical charge

Accounts for most energy lost during radioactive decay

Since no mass, it cannot result in the formation of new atoms.

${}^{0}_{0}\gamma$

$${}^{238}_{92}\mathrm{U} \longrightarrow {}^{234}_{90}\mathrm{Th} + {}^{4}_{2}\alpha + \gamma$$

uranium   thorium   alpha particle   gamma ray

Nuclear Stability

The primary factor in determining an atom's stability is the ratio of neutrons to protons.

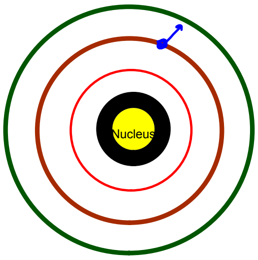
Atoms with too many or too few neutrons are unstable

Very few radioactive atoms are found in nature because most have already decayed into stable atoms.

Chapter 4 Test - 100 Points

- **15 Matching** *(3 sections of 5)*
  - Scientists, Radiation, Parts of Atom
- **35 Multiple Choice**
  - Everything in the notes is fair game
  - Focus on concepts, not memorizing
  - Calculating atomic mass.
  - Fill in the chart.
  - Which is correct for P=13, N=14, E=10?

Energy Levels of an Atom



Electrons can absorb or release energy