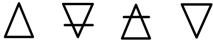


History of the Development of the Atomic Model

1. The first civilization that looked into the Atom is the Greeks
 - a. ~400 B.C.E.
 - b. Greeks tried to understand matter and broke matter into four groups:
 - i. Fire : Earth : Air : Water



- c. Democritus - Greek Philosopher
 - a. came up with the idea of "atomos"
 - i. Atomos = 'indivisible'
 - ii. 'Atom' is derived
 - b. No experiments to support idea
 - c. No protons, electrons, or neutrons but a solid that was indestructible

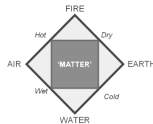


Democritus
(400 B.C.)

Quote from Democritus: "To understand the very large, we must understand the very small."

2. Aristotle - around 340 B.C.E.

- a. Another Greek philosopher
- b. Did not believe in the theory of atoms
- c. Thought that the idea of not being able to break down atoms any further would be putting a restriction on the Gods
- d. The concept that gods had unlimited power was quite popular and kept the idea of the atom from progressing for ~2,000 years.
- e. Introduced the relationship between the four elements and came up with the fifth element (ether)
 - i. thought that ether made up the heavens and could convert cheaper metals into gold and cure disease and old age.



3. After the Greeks, Chemistry was ruled by alchemy.

- a. during this time period, several elements were found
- b. equipment, procedure, and experimental techniques improved



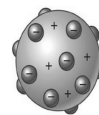
4. Robert Boyle - 17th century - wrote the book "Skeptical Chemist"

- a. In his book he stated:
 - i. A substance was an element unless it could be broken down to two or more substances
 - ii. Air therefore could not be an element because it could be broken down into many pure substances.
 - iii. This supported Democritus' thoughts on the atom and began the end of the thoughts of the mystical makeup of matter
 - iv. Stated that the term element needs a more precise meaning.
 - Earth can't be an element because gold, mercury, iron, and many others come from the Earth



5. John Dalton's Atomic Theory - 1808

- a. Elements are made of indivisible particles called atoms
- b. Atoms of the same element are exactly alike, in particular, they have the same mass

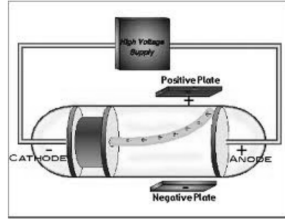


- c. For example: this gold bar can only be divided to the point that you end up with pieces (spheres) of gold that can no longer be divided.

Note: John Dalton was the first scientist to have evidence to support his atomic model.

6. Sir John Joseph Thomson - 1897

a. found that electrons can be stripped off of metal atoms



- b. The glass tube was stripped of most of its air
- c. An electron beam was shot from one side of the tube to the other
- d. It was found that the positive and negative plate could influence the direction of the beam
- e. He found that the particles that were bent were a couple thousand times lighter than the hydrogen atom (1 proton)

f. Conclusions from the experiment:

- i. Particles had been stripped off (means that atoms are divisible)
- ii. This was the discovery of the electron
- iii. Electrons had a very small mass



g. William Thomson (a.k.a. Lord Kelvin)

- i. Since atom was known to be electrically neutral, he proposed the plum pudding model.

Equal quantities of + and - charges
Charges evenly distributed in atom
+ is 2000 times bigger than -



Thomson's Model
"Plum pudding"

Robert Millikan's Oil Drop Experiment (1909)

a. experiment to determine the electric charge of an electron

b. The experiment entailed observing tiny charged droplets of oil between two horizontal metal electrodes.



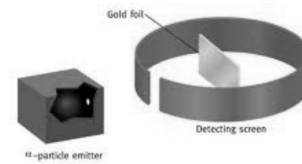
c. Oil droplet's terminal velocity was measured, which means that gravity and drag are equal, therefore the mass could be determined.

d. Knowing the mass and gravitational pull, a known electric field was applied and the oil droplets were suspended to where they were at mechanical equilibrium meaning their electrical force equaled their gravitational force.

e. From this they found the charge of oil droplets and they were always some multiple of 1.5924×10^{-19} C, about 0.6% difference from the currently accepted value of $1.602176487 \times 10^{-19}$ C.

f. They proposed that this was the charge of a single electron.

Atomic Target Practice Lab

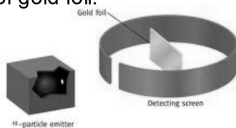


What Rutherford expected Actual experimental results Rutherford's model

Radius of nucleus is 1/100,000 of radius of atom.

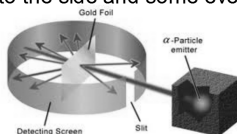
7. Ernest Rutherford (1909) - Gold Leaf Experiment

a. A radioactive source emitted alpha particles (a positively charged particle) that shot out in one direction towards a thin sheet of gold foil.



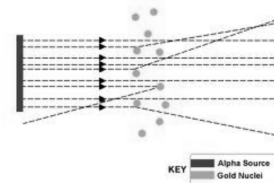
b. It was found that most of them passed right through the foil like it wasn't there.

c. Surprisingly, every once in awhile an alpha particle would hit the screen off to the side and some even bounced backwards.



d. conclusions from his experiment

- i. The positive proton was not spread out over the whole atom
- ii. Almost the entire mass of the atom was in a tiny area in the center (nucleus = 1/100,000 radius of atom)

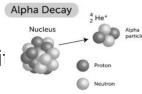


- iii. This is why the fast moving alpha particle would usually miss the tiny nucleus
- iv. If it did get real close, it would be pushed away by the positive repulsion of the protons in the nucleus.
- v. the negatively charged particles (electrons) orbit the nucleus

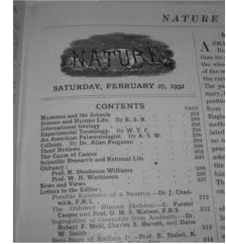
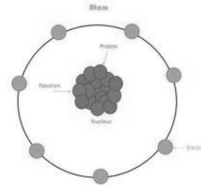
8. James Chadwick (1932)



- a. Chadwick joined Rutherford in research after being a POW during WWI
- b. Chadwick's research focused on radioactivity
- c. As they studied atomic disintegration (radiation), they kept seeing that the atomic number was less than the atomic mass.
 - i. Ex: A helium atom has an atomic mass of 4, but an atomic number (positive charge) of 2
- d. One idea was there was 4 protons in the nucleus along with 2 electrons to yield a charge of 2+



- e. After studying other researchers (Frederic and Curie), he looked for a particle with no charge and a mass near that of a proton
- f. He was successful and determined the neutron existed and published his findings in a paper entitled "Possible Existence of Neutron."
- g. The point of the neutron is for nucleus stability



Recent Atomic Models

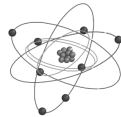
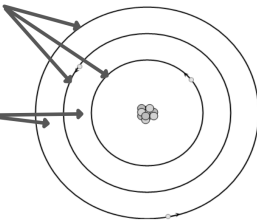
1. Niels Bohr (1913) - Bohr Model



- a. also known as the Planetary Model - electrons follow a set path much like the planets in the solar system
- b. Electrons can only possess certain amounts of energy therefore can only be certain distances from the nucleus.

Electrons found here

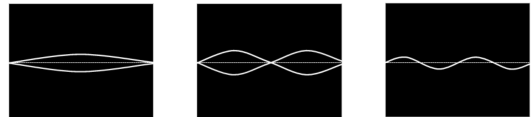
Electrons never found here



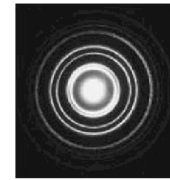
- c. explains how light is emitted from some atoms such as hydrogen

2. Louis de Broglie - 1924

- a. found that electrons exhibit wave properties
- b. wavelengths are quantized (meaning that there are only whole numbers involved: no in between)



Visible Light



Electrons

3. Werner Heisenberg - 1926



- a. Heisenberg Uncertainty Principle
 - i. States that we never know for certain the velocity or position of an electron at the same time

4. Erwin Schrodinger - 1926

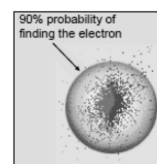
- a. wave equation of quantum waves
- b. only certain solutions, no in between #s
- c. quantized energy levels
- d. defines probability of finding an electron



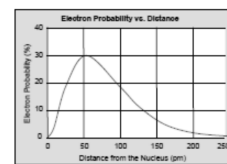
$$-\frac{\hbar^2}{2m} \frac{d^2 \psi}{dx^2} + V(x)\psi = E\psi$$

5. Quantum Mechanical Model (Model that is accepted today)

- a. also known as the electron cloud model
- b. Orbital ("electron cloud") - region in space where there is a 90% probability of finding an electron

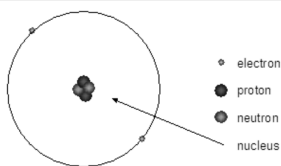


Orbital



6. Basics of the Atom

The Subatomic Particles



Particle	Charge	Location in the Atom	Mass
proton	1+	in nucleus	~ 1 u.a.m.u
neutron	0	in nucleus	~ 1 u.a.m.u
electron	1-	orbits nucleus	~ 0 u.a.m.u

$$1 \text{ u.a.m.u.} = 1.66 \times 10^{-27} \text{ kg}$$

7. Properties of the basics of the atom:



a. u.a.m.u. - unified atomic mass unit

- used to measure the mass of atoms
- Equal to 1/12 of the mass of an unbound neutral atom of carbon-12 in its ground state.
- has a value of $1.66 \times 10^{-27} \text{ kg}$

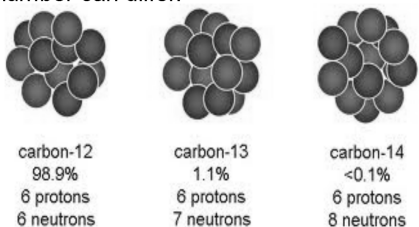
b. Atomic Number

- the number of protons for that element
- the whole number of Periodic Table
- determines the identity of atoms
- change this number, we change the identity of the atom

Atomic number	→ 6		
		C	← Elemental symbol
Element name	→ Carbon		
		12.01	← Atomic weight → Atomic mass

c. Mass Number (Atomic Mass)

- The number of protons + number of neutrons
- Protons and Neutron are the only two particles consider to have any mass
- This number can differ.



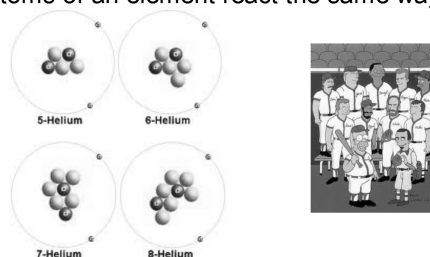
$$M.N. = \# \text{ of } p^+ + \# \text{ of } n^0$$

$$6+6=12 \quad 6+7=13 \quad 6+8=14$$

Isotopes

- Different varieties of the same type of element
 - have different number of neutrons
 - different atomic masses
 - some are radioactive, others are not

- All atoms of an element react the same way chemically.



Average Atomic Mass

- This is the weighted average mass of all atoms of an element that occur in nature and are not radioactive
 - measured in u.a.m.u.
- To find average atomic mass: (knowing % abundance)
 - You take the mass of an isotope and multiply times the relative abundance (how often it appears in nature)
 - Do the same for all isotopes
 - Add up the sum the numbers found in step a and b

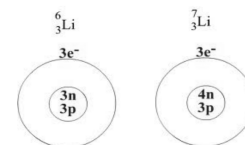
$$AAM = \text{mass A (R.A. of A)} + \text{mass B (R.A. of B)} + \dots$$

*Round using sig figs

Example:

Lithium has two isotopes. Li-6 occurs 7.5% of the time while Li-7 occurs 92.5% of the time. What is the average atomic mass of Lithium?

Answer:



2. To find percent abundance: (knowing weighted average)
 - a. You must set one of the % abundances equal to x and the other must be equal to (1-x)
 - b. Plug in numbers
 - c. Solve for x
 - d. The isotope you set equal to x is equal to your answer from c
 - i. multiply by 100 to find percentage
 - e. Solve for the other one by subtracting x from 1
 - i. multiply by 100 to find percentage

AAM = mass A (%A) + mass B (%B)

AAM = mass A (x) + mass B (1-x)

*Round using sig figs

Example:

Iridium is composed essentially of two isotopes: Ir-191 and Ir-193. The average atomic mass of an iridium atom is 192.2 uamu. Determine the natural abundance of each of these isotopes in a naturally occurring sample.

Answer:

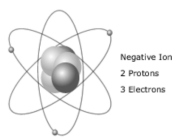
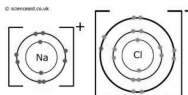
Ions

1. a charged atom
2. involves differences in protons and electrons

3. Two types of ions

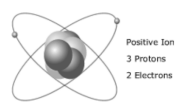
a. anion

- i. a (-) ion
- ii. more electrons than protons
- iii. formed when atoms gain electrons
- iv. usually found on the right side of the periodic table



b. cation

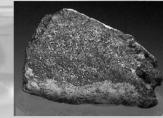
- i. a (+) ion
- ii. more protons than electrons
- iii. formed when atoms lose electrons
- iv. found on the left side of the periodic table



Question: Why can't we just add or take away protons?

Examples:

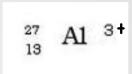
Description	Net Charge	Atomic #	Mass #	Ion Symbol
15 protons 16 neutrons 18 electrons				
protons neutrons electrons			128	Te ²⁻



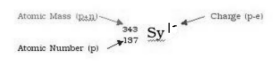
Complete Atomic Designation

1. gives precise information about an atomic particle

mass # charge (if any)



element
symbol





atomic #

Examples:

Protons	Neutrons	Electrons	Complete Atomic Designation
11	12	10	
			59 ³⁺ Co 27

6. The Mol

- a. Atoms are so small, it is impossible to count them by the dozens, thousands, or even millions. 
- b. To quantify atoms, we use the concept of the mol.
- c. A mol of any element is defined as the number of atoms of that element that equal to the number of atoms in exactly 12.0 grams of carbon-12.
- d. This solved a major issue of going between number of atoms, the u.a.m.u., and mass. 

7. Avogadro's Number

- a. The number of atoms or compounds in one mol of substance
- b. Value is: 6.02×10^{23} items/mol



Named in honor of Amedeo Avogadro

- Actually found by other scientists
- Millikan was first to accurately measure it
- X-ray technology today supports value

c. How big is 6.02×10^{23} ?



If you cover the entire Earth with 6.02×10^{23} standard marbles, they would be 2 miles deep.

6.02×10^{23} \$1 bills stacked face-to-face would stretch from the Sun to Pluto and back 7.5 million times. It takes light 9,500 years to travel that far.

602,000,000,000,000,000,000,000



- For any element on the Periodic Table, one mol of that element (i.e., 6.02×10^{23} atoms of that element) has a mass in grams equal to the atomic mass on the Table for that element.

- Example: 1 mol C-12 = 6.02×10^{23} atoms of C = 12 g C

7. To solve problems where we go between mass, atoms, and mols, we must use conversion factors.

- Conversion Factors we will be using are:

$$\left(\frac{1 \text{ mol of a substance}}{6.02 \times 10^{23} \text{ atoms/compounds}} \right) \quad \left(\frac{1 \text{ mol of a substance}}{\text{atomic mass of substance (g)}} \right)$$

Conversion factors can be flipped, need to figure out what we need to get rid of

You will change the unit of atomic mass from u.a.m.u. to grams

Examples:

1. How many mols is 3.79×10^{25} atoms of zinc?

Answer:

2. How many atoms is 0.68 mols of zinc?

Answer:

Steps:

1. Put known info
2. Place c.f. with unit you need to get rid of in appropriate place
3. multiply the top
4. divide the bottom

3. How many grams is 5.69 mols of uranium?

Answer:

4. How many grams is 2.65×10^{23} atoms of neon?

Answer:

5. How many atoms is 421 g of promethium?

Answer:

How many atoms are in 125g of Calcium?

How many gram are in 3.5×10^{26} atoms of Nickel?

How many grams are in 3.5×10^{26} atoms of Radium?

How many atoms are in 0.02 mols of Tin?