TOPIC 3.1

ATOMIC THEORY: DEMOCRITUS AND DALTON

What is everything made of, and how do we know?

The idea of the atom has been around for a loooong time. Let's journey through time and see how the **model** of the atom has evolved through the work of many scientists.

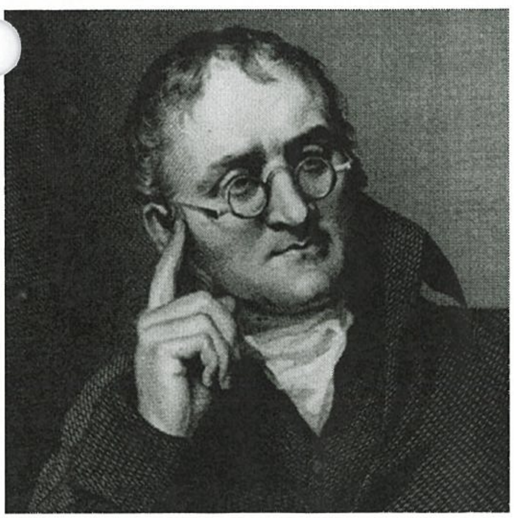
DEMOCRITUS

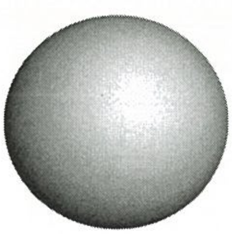
In 400 B.C., whilst enjoying a long walk on the beach in Greece...

the atom is the simplest, indivisible form of matter

"Now you might think that once Democritus came up with the general idea of atoms, it'd be pretty easy for someone else to take that little, indivisible ball and run with it. But you'd be wrong. The next major developments in atomic theory didn't come along for nearly 2300 YEARS. " – Hank Green

JOHN DALTON





FATHER OF ATOMIC THEORY:

Dalton was able to make a series of scientific claims based on experimental evidence about the structure and composition of matter:

1. All substances are made of __atoms__, which are the smallest particles of matter (__indivisible__)

2. All atoms of the same __element__ are alike and have the same mass. Atoms of different elements are different and have different masses.

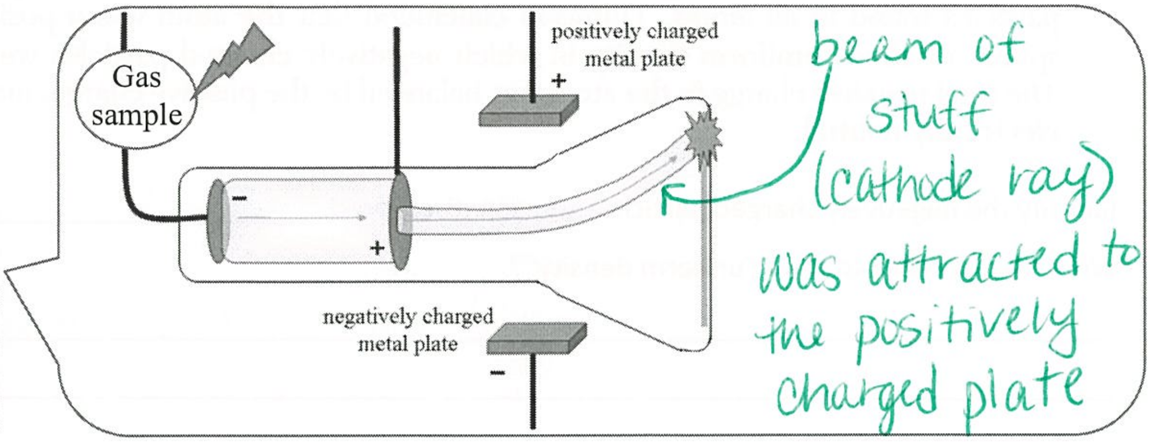
3. Atoms join together to form __compounds__, and a given compound always consists of the same kinds of atoms in the same proportions.

TOPIC 3.2

J.J. THOMSON AND THE ELECTRON

Can atoms be broken down into smaller parts?

Let's pick up where we left off...the idea of an atom as indivisible, or unable to be broken down. This turns out not to be true, and one scientist to prove this was **J.J. Thomson**, in 1897.

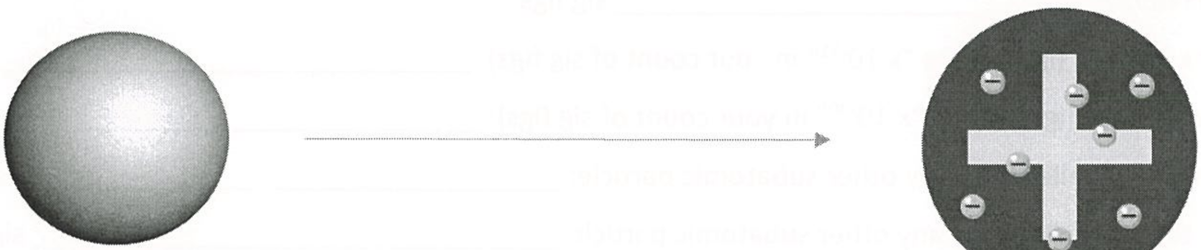


To explain his findings, Thomson overturned one of Dalton's claims. We CAN, in fact, break an atom down into smaller parts—**subatomic particles** called **electrons**.

Subatomic particle: smaller than the atom

Electron: negatively charged subatomic particles

Those **electrons** must be **embedded in some positively charged sphere** (to keep things neutral), kind of like chocolate chips embedded within a plain cookie.



Dalton Thomson

TOPIC 3.3

RUTHERFORD & THE NUCLEUS

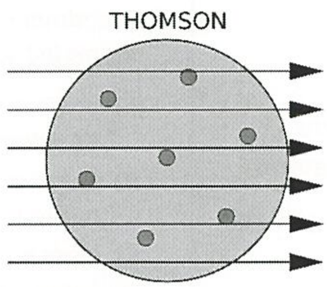
How can we learn more about the structure of something we can't even see?

Thomson's model revealed that there are particles smaller than the atom: **subatomic particles.** Other scientists were now eager to see if there was more hidden inside the atom. Trouble is, when the atom is so small, it's hard to see much. **Ernest Rutherford** came up with a way to **indirectly** "see" the inside of an atom:

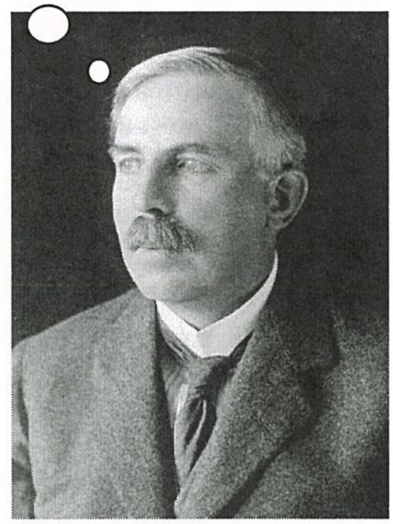
Rutherford's Gold Foil Experiment

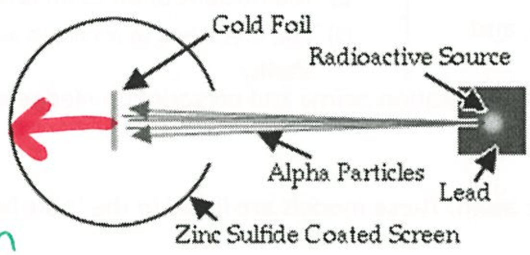
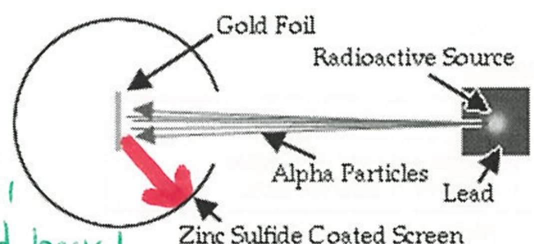
If Thomson's model of the atom is correct, then we would predict:

all the particles will pass straight through

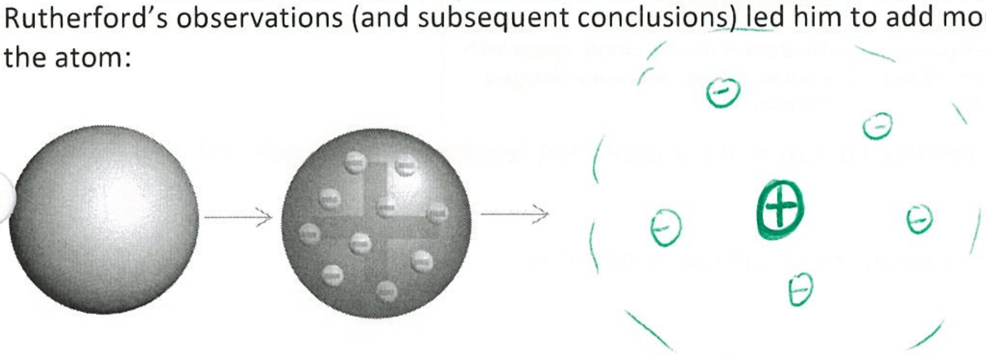


ERNEST RUTHERFORD



Experimental Result (Observation)		Explanation of Result (Conclusion)
most particles went straight through		the atom is mostly empty space!
Some particles were deflected, or bounced back!		there must be a small, dense, positively charged center → NUCLEUS

Rutherford's observations (and subsequent conclusions) led him to add more detail to the developing model of the atom:



TOPIC 3.4

ISOTOPES & MASS NUMBER

How can two atoms of the same element be different?

After Rutherford's research regarding the interior of the atom, two more subatomic particles were identified and named. We now know that each atom is made up of **three different subatomic particles**. Each of these subatomic particles contributes something unique to the atom.

atomic mass units

Subatomic Particle	Charge	Location	Mass (amu)	Responsible for...
Electron	−1	outside nucleus	0	reactivity of atom
Proton	+1	nucleus	1	identity of atom
Neutron	0	nucleus	1	stability of nucleus

Nuclear charge: charge of the nucleus (always positive, equal to the # of protons)

Net charge: overall charge of the atom (atoms are neutral because # protons = # electrons)

How do we know how many **neutrons** are in an atom? Chemists use different **symbols** to represent the *mass number* and *atomic number* of a given atom. Here's a couple different ways to represent an atom of carbon:

mass # (protons + neutrons) → 12

$^{12}_{6}C$

atomic # (protons) → 6

carbon – 12

↑ element name ↑ mass #

Isotopes: atoms of the same element, but have different masses (due to diff. # of neutrons!)

Fill in the missing element symbols and circle the two **isotopes**.

$^{32}_{16}$ S $\qquad$ $^{34}_{17}$ Cl $\qquad$ $^{34}_{16}$ S

TOPIC 3.5

AVERAGE ATOMIC MASS

How can we discuss the mass of a large sample of atoms of an element?

We've determined that the mass of an atom comes from its protons and neutrons

35.453 –1 +1 +5 +7
Cl
17
2-8-7

What's the deal with the decimal? We can't have partial protons or neutrons!

Average atomic mass: weighted average of all naturally occuring isotopes of an element

isotopes → same element, diff. masses (b/c diff. # of neutrons)

Naturally Occurring Isotopes of Chlorine

Isotopes	Atomic Mass of the Isotope (u)	Natural Abundance (%)
^{35}Cl	34.97	75.76
^{37}Cl	36.97	24.24

$(34.97)\times(0.7576) + (36.97)\times(0.2424)$

$26.493272 + 8.961528$

35.45 amu

Calculate the average atomic mass of sulfur if 95.00% of all sulfur atoms have a mass of 31.972 amu, 0.76% has a mass of 32.971 amu and 4.22% have a mass of 33.967amu.

$(31.972)\times(0.9500) + (32.971\times 0.0076) + (33.967\times 0.0422)$

$30.3734 + 0.2505796 + 1.4334074$

32.06 amu

Identify the **most abundant** *isotope of fluorine. How do you know? Do all fluorine atoms have this mass?*

$^{19}_{9}F$ fluorine-19

→ Average atomic mass (18.9984) is closest to 19.

→ No! Just a very large percentage do

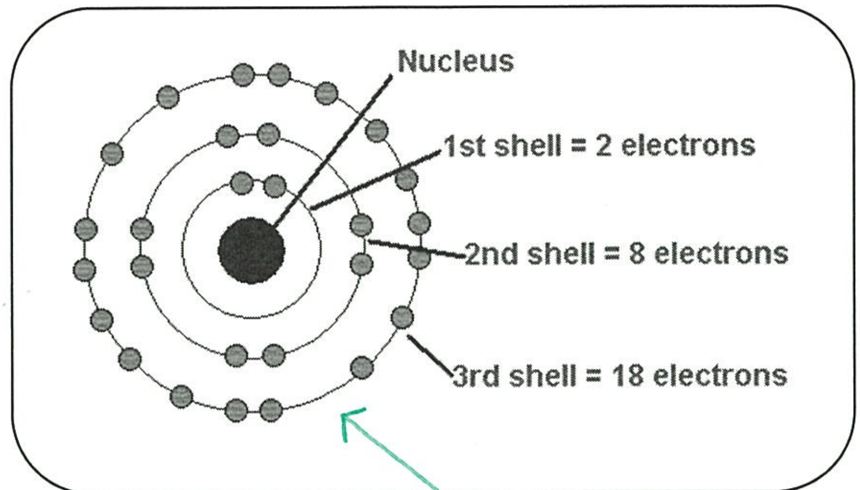
TOPIC 3.6

NIELS BOHR and ENERGY LEVELS

How are electrons arranged within an atom?

Niels Bohr was the scientist that gives us the most useful model of the atom to work with: the Bohr model, also known as the planetary model. Through experimentation which we will learn about in the next set of mini lessons, Bohr determined how electrons were arranged within the atom.

NIELS BOHR

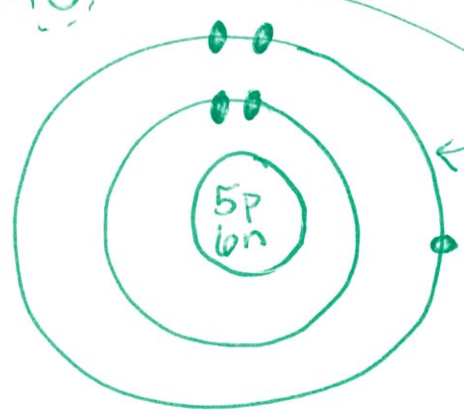


Bohr's model shows that electrons are arranged (or configured) in energy levels. Each **energy level** (also called **energy shells**) has a distinct amount of energy, which increases as you get *further* from the nucleus. So, electrons in the shell closest to the nucleus have the *least* amount of energy, and electrons in the shell furthest from the nucleus have the *most* energy.

Example: electron configuration and Bohr diagram for the most abundant isotope of boron

2-3

mass # = 11 (round atomic mass)



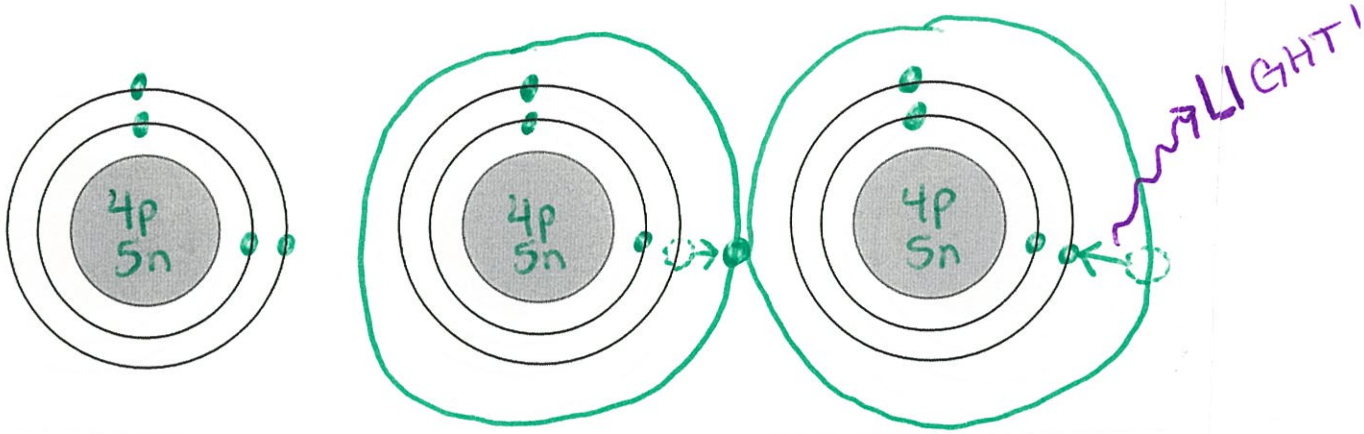
# of electrons in the last shell are called valence electrons

TOPIC 3.7

GROUND vs EXCITED STATE

How do electrons move within an atom when they absorb energy?

Bohr's conclusions about the structure of the atom came from the analysis of hydrogen atoms (and a lot of ensuing math). He "hit" hydrogen atoms with a surge of energy and observed the resulting changes. Here's the overview of what he concluded he'd seen:



Step 1: Start with an atom in its "regular" state

Beryllium-9

e⁻ config: 2-2

ground state

Step 2: Add energy to "excite" an electron to a higher energy state

e⁻ config: 2-1-1

excited state

Step 3: Watch as electron releases energy in the form of light as it returns to its more stable "ground state"

e⁻ config: 2-2

ground state

Ground State Stable, lowest energy electron configuration (on the Periodic Table)

Excited State unstable, high energy electron configuration (NOT on Periodic Table)

TOPIC 3.8

Bright Line Spectra

How can we use the behavior of electrons to identify elements?

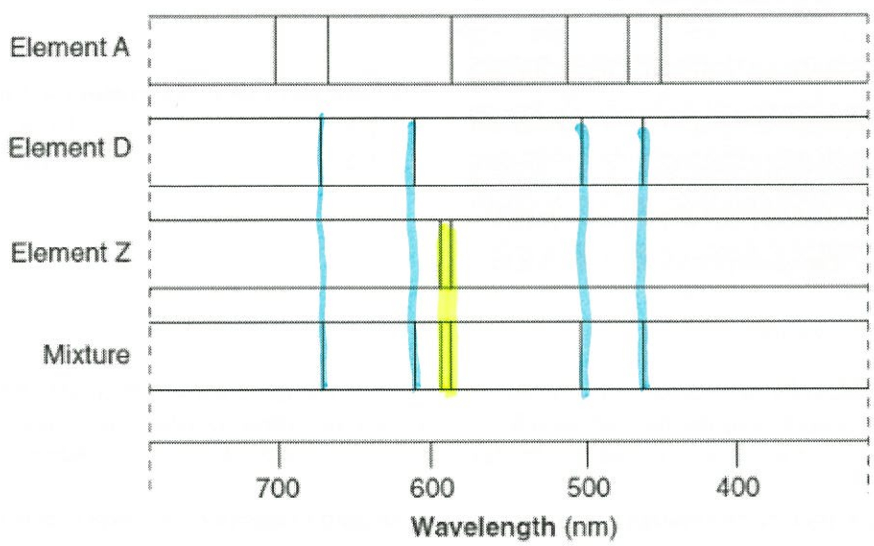
The structure of atoms of each element is consistent across the *entire universe.* Helium here on Earth is exactly the same as helium on the Sun, which is exactly the same as helium from the most distant start we can detect. The elements are the common building blocks of the Universe.

Each element also leaves a characteristic "fingerprint" that we can detect as an atom moves from the excited state back down to the ground state.

Bright-line spectrum: "fingerprint" that shows the wavelengths of light (= individual colors) that an element emits

This "fingerprint" is more precise than the color seen on a **flame test** (just heating and observing), since a bright-line spectrum breaks down the visible light into its component wavelengths. Let's take a look at what that means/looks like:

Selected Bright-Line Spectra



In terms of electrons and energy states, explain how the lines in the bright-line spectra were generated.

Electrons from higher energy states emit light when they drop to lower energy states.

Which elements make up the mixture? How can you tell?

D & Z → all spectral lines match up!

TOPIC 3.9

Schrodinger and the Orbital

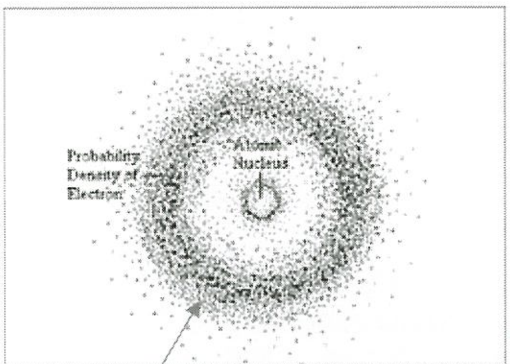
How did modern physics overturn the model of the atom?

In our story uncovering the history of atomic theory, we left off with Bohr's nice and neat model of the atom. However, in the early twentieth century, work in quantum theory by Erwin Schrodinger and others totally changed our understanding of electron behavior.

"Anyone who is not shocked by quantum theory has not understood it."

SCHRODINGER

WAVE-MECHANICAL/CLOUD MODEL



we can't know the exact location of an electron, but can predict the odds of its location

Orbital: the region that is the most probable location of an electron in an atom

1. Four statements about the development of the atomic model are shown below.

   A: Electrons have wavelike properties. (4) Schrodinger
   B: Atoms have small, negatively charged particles. (2) Thomson
   C: The center of an atom is a small, dense nucleus. (3) Rutherford
   D: Atoms are hard, indivisible spheres. (1) Dalton

   Which order of statements represents the historical development of the atomic model?

   A) $C \rightarrow D \rightarrow A \rightarrow B$ B) $C \rightarrow D \rightarrow B \rightarrow A$
   C) $D \rightarrow B \rightarrow A \rightarrow C$ D) $D \rightarrow B \rightarrow C \rightarrow A$