Structure of the Atom

The first questions about the examination of what matter is come to us from the ancient Greeks. It was they who first attempted to explain the nature of matter. The ancient Greeks searched for knowledge. The Greek word for knowledge is 'soph'and the Greek word for 'admirer' or 'seeker' is 'Philo.' Those that searched out understanding were called Philosophers. Those that specialize in the understanding of nature and those things surrounding them were therefore referred to as "natural philosophers." In fact, hopefully, many of you may eventually achieve the title of Doctor of Philosophy (PhD) in one of the natural science branches.

MODELS OF MATTER

A model is a tentative description of a system or theory that accounts for all of its known properties. Models are invented, for the most part, to interpret the invisible and relate experimental results. In ancient Greece, philosophers argued on two concepts of matter.

A. Democritus Model

- Matter is made up of particles that are indivisible ("atomos" Greek for indivisible).
- These particles are small, hard, incompressible, and indestructible.
- Two atoms cannot occupy the same space at the same time.
- Atoms of different materials differ from each other in shape, mass, and size.
- Atoms are made up of only two things: atoms and emptiness, or a void.
- Believers in the indivisibility of matter were called ?atomists.?

B. Aristotle's Model

- Matter can be divided into an infinite number of parts.
- Matter is continuous.
- Matter is made up of four elements: fire, water, earth, and wind.
- Differences in substances lay in the proportion of these four elements making up every material.
- Aristotle tutored Alexander the Great, thus his explanation stood for 1000 years.

C. John Dalton's Model (1806)

By the beginning of 19^{th} century, precision of analytical chemistry had improved so greatly that chemists were able to show that simple compounds contained fixed and unvarying amounts of their constituent elements**.**

In certain cases, more than one compound could be formed between the same elements**.** Joseph Gay-Lussac, at around the same time, in France, showed that the volume ratios of reacting gases were small whole numbers**.**

Dalton, in 1806 provided a major step in explaining this with his chemical atomic theory, or the Particle Theory**.**

The seven postulates of Dalton′s Particle Theory are:

1. *All matter is composed of extremely small particles called atoms.*
2. *Atoms can neither be subdivided nor changed into one another.*
3. *Atoms cannot be created or destroyed.*
4. *Atoms of one element are the same shape, size, mass, etc.*
5. *Atoms of one element differ in properties from other elements? atoms.*
6. *Chemical change is the union or separation of atoms.*
7. *Atoms combine in small whole-number ratios (1:1, 1:2, 1:3, etc.)*

therefore:

an atom is the smallest particle of an element that has all the chemical properties of that element

a molecule is the smallest particle of a compound that shows all the chemical properties of that compound

Problems with Dalton′s Atomic Theory:

- did not account for law of multiple proportions
- made no distinctions between atoms and molecules
- thus, he could not distinguish between formulas for water HO and H_2O
- could not explain why density of water vapor was less than that of oxygen

Solution to these problems found in 1811 by Amedeo Avogadro, who suggested that oxygen combined with hydrogen by splitting the "double atom" of oxygen to combine with two hydrogen atoms (O_2 and H_2)

However, it took almost 50 years for his ideas to prevail; the early-mid 1800's were filled with confusion amongst chemists and their calculations
In 1860, Stanislao Cannizarro reintroduced Avogadro′s hypotheses

Homework:

Read: p.80-84
Explain: the term atomic mass unit
Do: p.84, Practice Exercises 4-1, 4-2; p.88, Part One Review # 1-6

D. William Crookes′ Cathode Ray Tube

As the study of static charging of objects progressed, it became more clear that Dalton′s model of the atom was incomplete.

It was shown that the charging of an object was the result of collection of charged particles, of which two types were discovered**.**

Whether objects attracted or repelled each other depended on the nature of these particles**.**

Discoveries were made that there were different gases, and thus it became important to investigate all the properties of these newly discovered materials**.**

Volta invented the electrical battery in 1800, and experiments were carried out, using electrical conductivity as a characteristic property for materials

William Crookes performed research on the conduction of gases, and came to invent the Crookes' Tube, known today as the "Cathode Ray Tube"

Homework:
Draw and label: a Cathode Ray Tube:

Useful definitions:

- **Cathode ray tube:** vacuum, glass tubes with an electrode at each end
- **Cathode:** the negative side of the cathode ray tube
- **Anode:** the positive side of the cathode ray tube
- **Cathode ray:** the ray made of negatively charged particles (electrons)

Crookes′ original intentions for the tube were to test out the conductivity of various gases, and it never occurred to him that his invention would be used as the source of explanation that would improve Dalton′s model.

Dalton′s model needed mending, in that it could not explain the phenomena of static electricity (why is there attraction and repulsion, why are there only two types, not three or four, etc.).
All these questions bothered scientists at the time, and every effort was made to try and explain these natural phenomena.

E. Joseph J. Thompson′s Model

He saw Crookes′ tube and realized that Dalton was wrong.
He then built the pinwheel cathode ray tube to help prove his cause.
As air or gas was removed from the tube, there would be a glow; rays appeared to begin at the cathode and traveled toward the anode (+ to -).
The "things" flying in the tube were particles of matter, capable of spinning the pinwheel. This meant that they had a mass, but they were much smaller than atoms.

The rays were then put against charged ebony, which repelled them. The ebony being classified as negative material at the time, Thompson thought that, for repulsion to occur, the rays must have been charged in the same way.

The battery had two connections, one of which was referred to as the cathode, the other as the anode. Knowing the rays came out of the cathode side of the tube were repelled by ebony, Thompson deduced that they were also negatively charged. They, therefore determined that the cathode was negative and the anode was positive.

Hence, Thompson realized that Dalton had erred. The indivisible particle of matter, the atom, now had at least two parts. Negatively charged material, as well as positively charged material.

Thompson was thus credited with the discovery of the **electron**, the negative particle of matter. Electrons are identical no matter what material is present in the tube. He determined that these particles did in fact have a mass, but was only able to determine the charge:mass ratio (e/m).

Thomson′s model was very similar to Dalton's, but said that the atom is a positive body that contains embedded electrons. He used the analogy of raisins in raisin bread, since the bread could represent a bundle of positive charge, the raisins the tiny electrons. Thus, it came to be known as the **"Raisin Bread Model."**

He claimed matter is naturally neutral, but sometimes becomes charged negatively by gaining excess electrons, or positively by losing electrons. This difference in charges makes for attraction and/or repulsion between objects.

Homework:

Read: p.84-87
Describe: characteristics of electrons and protons; Goldstein?s experiments
Do: p.88 #7-12, p.120 #4-5, p.121 #3-5

F.   Rutherford′s Model

His model came after several major advancements in science:

Roentgen, in 1895, experimented with CRT′s and found that, under high voltage, X-Rays were emitted. The high frequency and high energy allow for them to pass through soft tissue, making them extremely useful in medicine.

Becquerel, in 1896, left an unexposed film under a chunk of uranium ore, and found it had been exposed, even though it was not exposed to any light. He discovered that uranium spontaneously emitted powerful invisible rays, and that substances that do this are called **radioactive**.

Pierre and Marie Curie succeeded in isolating two new radioactive elements; polonium and radium. Marie Curie went on to hold a professorship at Sorbonne, and became the first person to win the Nobel Prize twice.

Rutherford′s preliminary experiment consisted of placing grains of radium salt in the bottom of a hole in a block of lead. The radiation escaped from the hole and was allowed to strike a screen painted with zinc sulfide (individual atoms are far too small to be seen, but when the screen is struck by an atom or an atomic particle it produces a tiny scintillation of light). The radiation hitting the screen caused the zinc sulfide to glow in one spot.
By bringing a magnet close to the rays, three spots of light could be seen. Rutherford found that, therefore, three different types of rays are emitted by radioactive substances:

Beta rays are the lightest rays, and were attracted to the positive plate. They are made of negative particles, and can pass through several mm of aluminum.

Alpha rays are positively charged, since they were attracted to the negative plate. They are helium nuclei, with a mass of 4 amu, and can be stopped with a piece of paper.

Gamma rays are not deflected by a magnet, therefore they are not charged. They have no mass, but instead, are like waves of light. They have high frequencies and high energies, and can pass through 30 cm of steel or 5 cm of lead.

Homework:
***Read:* p.85-86 (Rutherford′s Model of the Atom)**

Rutherford performed a gold foil experiment at McGill University, Montreal. He bombarded thin gold foil with alpha particles: he noticed that while most penetrated the foil undeflected, about 1 in 2000 suffered serious deflections as they penetrated, and a similar number bounced back instead of penetrating.

Thomson′s model did not explain these deflections. This type of behaviour would be expected only if the positive charge of an atom were highly concentrated in a small region. Rutherford called this region the **nucleus**. The approach of an alpha particle to a nucleus of high positive charge and mass would lead to repulsive forces strong enough to reverse the direction of the particles.
Rutherford formed his **Planetary Model** based on these findings:

- *Most of the mass and all of the positive charge of an atom are centered in a very small region called the nucleus. The atom is mostly empty space, since there is nothing between the nucleus and the electrons).*
- *The magnitude of the charge on the nucleus is different for different atoms and is approximately one half of the numerical value of the atomic mass of the element.*
- *There must be a number of electrons outside the nucleus of an atom that is equal to the number of units of nuclear charge.*

Homework:
Read: p. 90-91
Explain: Rutherford's experiments that led to the concept of the nuclear atom
Describe: Rutherford′s nuclear model and two major problems that it has
Do: p. 104 #1-5, p. 120 #6-8, p. 121 #6-7
Calculate: how many times smaller a nucleus is than an atom
(an atom′s diameter ~ 10^{-10} m; a nucleus′ diameter ~ 10^{-15} m)

Weaknesses of Rutherford′s Model:
Classical physicists observed that a moving electric charge such as an electron ,which changes its direction in space, must release, or radiate, energy. If it gives off energy, it will slow down and therefore not be able to resist the attraction of the positive nucleus and hence should rapidly spiral into the nucleus and eventually a small mushroom cloud will be seen. The collapse of the atom should be observed, however, atoms do not collapse. Therefore, there must be a flaw in Rutherford′s model.

Homework:
Read: p.96-103, 103-117
Define: **continuous spectrum, line spectrum**
Do: p.104 #1-3, p.122 #16-21

SPECTROSCOPY

The spectroscope is an instrument used to see the spectrum of light emitted by different elements.
Each element has its own characteristic light spectrum.

Visible light spectrum
The white light is formed by different colours: **ROYGBIV**
Each coloured ray has its characteristic wavelength and frequency. There are other rays that can be detectable but not visible.

All the rays form the **ELECTROMAGNETIC SPECTRUM**

Radio, Waves, TV Waves	Micro-waves, Radar, Sonar	infra-Red	Visible light	Ultra-Violet	X-Rays	γ -rays

Long waves ---> Short Waves
Low Frequenc--> High Frequency
Low Energy--> High Energy

G. Bohr′s Model

A Danish scientist, he started looking at the scientific evidence in a different way, and came up with a series of postulates (a postulate is something assumed to be true) ...

- In every hydrogen atom, there are only certain paths in which an electron is allowed to move (these are called energy levels)
- Each energy level corresponds to an orbit, a circular path in which the electrons can move around the nucleus
- Electrons can travel in allowed energy levels without loss of energy
- Electrons may jump from one energy level to another

When electrons are stationary, they are stable and do not radiate energy. An electron in the lowest possible energy level is said to be in the **ground state**. When heated or given energy, it occupies a higher energy level (**excited state**).

Atoms have only certain allowed orbits (electrons orbit the nucleus on different energy levels, and the energy of an electron is quantized, or limited to a certain set of values). The energy differences are therefore specific values.

An atom radiates energy only when the electron jumps from one allowed state to a lower one. An excited electron drops to a lower energy level and emits radiation of a certain frequency. This frequency depends on the size of the jump as well as the final level that the electron reaches.

The frequencies of the colored lines in the visible spectrum of hydrogen matched perfectly with the theory. Bohr?s atomic model gained credibility when he was able to predict lines outside the visible region (ultraviolet). The fact that atoms could give off distinct spectra has many applications:

- Neon signs, sodium lamps, mercury lamps, flares
- Identifying elements; spotting of chemicals in space from their spectra

Drawbacks:

This theory only worked for hydrogen, and for no other element. When applied to multi-electron atoms, several series of lines were seen that should not have been there. These extra lines were given names:

Sharp
Principle
Diffuse
Fundamental

Bohr was also wrong in thinking that an electron′s position and motion can be specified at a given time, as well as thinking that orbits have fixed radii.

Homework:

Summarize: four atomic models: Dalton, Thomson, Rutherford, and Bohr

H. The Present Model of the Atom

Physics progressed up to the point where electrons were thought to be waves as well as particles. Werner Heisenberg came up with one of the most quoted principles of science: the **Heisenberg Uncertainty Principle**. To explain this principle, consider the following situation:

Waves in the ocean, next to a beach. In order for a wave to be disturbed, it must hit something of roughly the same size or of bigger size than the wavelength. What will happen when the wave runs into a pebble? A boulder instead?
Hence, to see exactly where an electron is, what would be a characteristic of the light that we would bounce off of it?

If light with a very small wavelength (high frequency, high energy) struck the electron, what would happen? ?We would lose information about the electron′s velocity. If we hit it with a long wavelength (low energy and frequency), we would not gather information about its position (see boulder-pebble analogy above).

Thus, we will never be able to know an electron′s velocity and position at the same time. The best we can do is calculate the **probability** of where an electron might be.

Wave mechanics considers electrons as waves, and we can therefore apply mathematical methods to come up with pictures of electron distribution. Now, we do not talk about electrons occupying orbits, like in the planetary model, but **orbitals**, which describe how their changes are spread out in small regions in space.

Definition: An **orbital** is a 3D space where an electron is likely to be found.

The principal quantum number (n) identifies energy possessed by electrons:

- *$n = 1 \longrightarrow$ infinity*
- *$2n^2$ number of electrons in an energy level*
- *n types of orbitals*
- *n^2 actual orbitals*

There are an infinite number of energy levels, but as you move further away from the nucleus the levels are closer and soon become indistinguishable.

There are four types of orbitals whose shapes have been worked out, using wave mechanics. They are referred to by letters (s, p, d, and f). These are the initial letters of the words ?sharp?, ?principal?, ?diffuse?, and ?fundamental?, originating from work carried out on the hydrogen spectrum which led to our present-day view of energy levels and sublevels.

n = 1:	1 type of orbital	1 actual orbital	‵s′ orbital
n = 2:	2 types of orbitals	n^2 orbitals (4)	‵s′ (bigger), 3 ‵p′
n = 3:	3 types of orbitals	n^3 orbitals (9)	‵s′, 3 ‵p′, 5 ‵d′
n = 4:	4 types of orbitals	n^4 orbitals (16)	s, 3p, 5d, 7f

Questions:

1. Explain why 'p' orbitals are labeled p_x, p_y, and p_z.
2. How many electrons can be held in a 's' orbital? a set of 3 'p' orbitals?
3. Use one word to describe the shape of a 's' orbital. a 'p' orbital.
4. Fill in the following table:

Energy Level (n)	Number and Type of Orbitals (n, n^2)	Max. Number of Electrons in each Set of Orbitals	Max. Number of Electrons in Energy Level ($2n^2$)
n = 1	1s		
n = 2	1s		
	3p		
n = 3	1s		
	3p		
	5d		
n = 4	1s		
	3p		
	5d		
	7f		

With more orbitals available, many more electronic transitions are possible, thus explaining all the extra lines in the spectra observed earlier. The letters correspond to the first letter in each series of lines observed.

Electrons fill energy levels according to the following set of rules:
The Aufbau Rule
Hund's Rule
Pauling's Exclusion Principle

In an atom the orbitals are filled in order of increasing energy starting from 1s. An aid to remember this order is to write down the orbitals in columns.

Homework:

Do: p.112 #4-7, #4-9; p.114 #4-10, #4-11; p.122 #27, 28a-e, 31, 32
Find out about: the three rules mentioned above and fill in short explanations