

Atomic Structure

Prentice Hall *Physical Science* – Chapter 4

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Section 4.1 – Studying Atoms

History of Atomic Theory

(or how do we study atoms?)

- Atoms are difficult to study because they are so small .
- Experiments have evolved over time to help increase our understanding of the atom.
- Recall that a scientific model is a representation of an object or event that makes it easier to understand things that are difficult to observe directly.

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Ancient Greek Model of Atom

(or: Democritus v. Aristotle, and Aristotle won)

- ~ **2500 years ago**
- Democritus' model:
 - Matter composed of *extremely* small particles that cannot be divided
 - He called the particles *atoms*, from the Greek atomos (“uncut” or “indivisible”).
- Aristotle's model:
 - No limit to how much matter could be divided
 - All matter was built from earth, air, fire and water
- Aristotle's model prevailed until the 1800's.



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Dalton's Atomic Theory

- ~ **250 years ago**
- Dalton's evidence for atoms
 - Measured masses of elements that combine when compounds form
 - The ratio of masses of elements in the compound was always the same regardless of sample size
- Dalton's Theory:
 - All elements are composed of atoms
 - All atoms of the same element have the same mass, and atoms of different elements have different masses
 - Compounds contain atoms of more than one element
 - In a particular compound, atoms of different elements always combine the same way

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Thompson's Model of the Atom (or the "Plum Pudding" Model)

- ~ **110 years ago**
- J. J. Thompson used an electric current to learn about atoms
- Discovered that particles in the beam
 - Were negatively charged
 - Came from inside atoms
 - Were identical regardless of the metal source
 - Had about 1/2000 the mass of a hydrogen atom
- First evidence of subatomic particles
- Thompson's Model: Negative charges evenly scattered throughout an atom filled with a positive mass

Rutherford's Atomic Theory

- ~ **90 years ago**
- Rutherford discovered positively charged subatomic particles (alpha particles, or α particles)
- Gold Foil Experiment
 - Rutherford asked a student, Ernest Marsden, to find out what happens to α particles when they pass through gold foil
 - Hypothesis, based on Thomson's model, was that most α particles would pass straight through
 - More α particles were deflected than expected – about 1 out of 20,000 was deflected more than 90°
 - Concluded that positive charge is not evenly distributed
- Rutherford's model: *all of an atom's positive charge is concentrated in the **nucleus***

Subatomic Particles

- Protons
 - Positively charged (assigned charge of +1)
 - Found in nucleus
- Electrons
 - Negatively charged (assigned charge of -1)
 - Found outside of nucleus
- Neutrons
 - Existence proven by James Chadwick in 1932
 - Neutral particle (0 charge)
 - Found in nucleus

Comparing Subatomic Particles

- Protons, electrons and neutrons are distinguished by
 - Mass
 - Charge
 - Location in an atom

Properties of Subatomic Particles				
Particle	Symbol	Relative Charge	Relative Mass (Proton = 1)	Actual Mass (g)
Electron	e^-	1 ⁻	1/1836	9.11E-28
Proton	p^+	1 ⁺	1	1.67E-24
Neutron	n	0	1	1.68E-24

Atomic Number and Mass Number

- Atomic number
 - Atoms of any given element always have same number of protons
 - Atoms of different elements always have different number of protons
 - *Atomic number of an element* = # of protons in one atom of that element
 - Atoms are neutral, so each proton in an atom (with its positive charge) is balanced by one electron (with a negative charge); thus, atomic number also equals number of electrons
- Mass number
 - *Mass number of an atom* = # of protons + # of neutrons in nucleus of that atom
 - *Number of neutrons* = *Mass number* – *Atomic number*

Isotopes

- All atoms of an element have the same number of protons
- NOT ALL atoms of an element have the same number of neutrons
- **Isotopes** of an element: atoms of the same element that have different numbers of neutrons, and thus different mass numbers
- Isotopes are referred to by their name and mass number when needed (example: hydrogen-1 and hydrogen-2)
- Example: heavy water is made up of two hydrogen-2 atoms bound to oxygen

Bohr's Model of the Atom

- Focused on electrons
 - Electrons move with constant speed in fixed orbits around the nucleus
 - Each electron has a specific amount of energy
 - **Energy levels**: the possible energies that electrons can have
 - No two elements have the same set of energy levels
- An electron in an atom can move from one energy level to another when an atom gains or loses energy
- Evidence for energy levels: Light given off in fireworks

Electron Cloud Model

- ~ **80 years ago**
- Bohr's model assumed electrons moved in predictable orbits, but later evidence showed they move more randomly
- Electron cloud: visual model of the most likely locations for electrons in an atom
 - Cloud is denser where electrons are more likely to be
 - Represents all the **orbitals** in an atom (**orbital**: a region of space around the nucleus where an electron is likely to be found)

Atomic Orbitals

- Electrons at certain energy levels can only occupy certain orbitals
- Orbitals can only contain 2 electrons
- The lowest energy level (where the lowest energy electrons are) only has 1 orbital, whereas higher energy orbitals have more than one (see table below)

Energy Level	Number of Orbitals	Maximum Number of Electrons
1	1	2
2	4	8
3	9	18
4	16	32

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Electron Configurations

- **Electron configuration:** the arrangement of electrons in the orbitals of an atom
- Most stable configuration: electrons occupy lowest-energy orbitals (called the **ground state**)
- An atom in an **excited state** has absorbed enough energy for one electron to move to a higher-energy orbital
- Example: Neon atoms (and atoms of other noble gases) emit energy in the form of light when their electrons return from an excited state to the ground state ("neon lights")

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