Models of the Atom

Objectives

- Describe Dalton’s model of the atom.
- Describe Thompson’s “plumb pudding” model of the atom.
- Describe Rutherford’s nuclear model of the atom.
- State the relative charge and approximate mass of the electron, proton, and neutron.

Objectives

- Describe the difference between isotopes of the same element.
- Given an atom's atomic notation; state the number of protons, neutrons, and electrons in the atom.
- Given the number of protons or electrons, and the number of neutrons, write the atomic notation for the atom.

Objectives

- Calculate the atomic mass of an element given the mass and abundance of the naturally occurring isotopes.
- State the relationship between the frequency and the wavelength of light.
- State the relationship between the frequency and the energy of light.
- Describe the Bohr model of the atom.

Objectives

- Explain the relationship between energy levels in an atom and the lines in its emission spectrum.
- Determine the maximum number of electrons that can occupy any given energy level.
- List the sublevels according to increasing energy.

Objectives

- Write the predicted electron configurations for selected elements.
- Describe the quantum mechanical model of the atom.
- Describe the relative sizes and shapes of the *s* and *p* orbitals.

Early Models of the Atom

- Democritus’ Model - Matter is composed of indivisible particles. Democritus called these particles atoms from the Greek word atomos meaning indivisible.
- Aristotilian Model - Matter is continuous. It can be divided an infinite number of times. (Accepted for 21 centuries because of Aristotle’s stature as a thinker.)

Dalton’s Model (1803)

- The model:
  - An element is composed of tiny indivisible, indestructible particles called atoms.
  - All atoms of an element are identical and have the same properties.
  - Atoms of different elements combine to form compounds.
  - Compounds contain atoms in small whole number ratios.
  - Atoms may combine in more than one ratio to form different compounds.

Dalton’s Model

- Basis for Dalton’s Model:
  - Robert Boyle - Gases are composed of tiny particles, solids and liquids are probably composed of tiny particles as well.
  - Antoine Lavoisier - Mass is conserved in chemical reactions.
  - Joseph Proust - Verified the Law of Definite Composition - Compounds always contain the same elements in the same proportion by mass.

Thompson’s Model (1903)

- Thompson’s Model:
  - Atoms are composed of smaller particles.
    - Electrons with a negative charge and small mass.
    - Protons with a positive charge and large mass.
  - Atoms consist of homogeneous spheres of positive charge with electrons stuck in them like raisins in a plumb pudding.

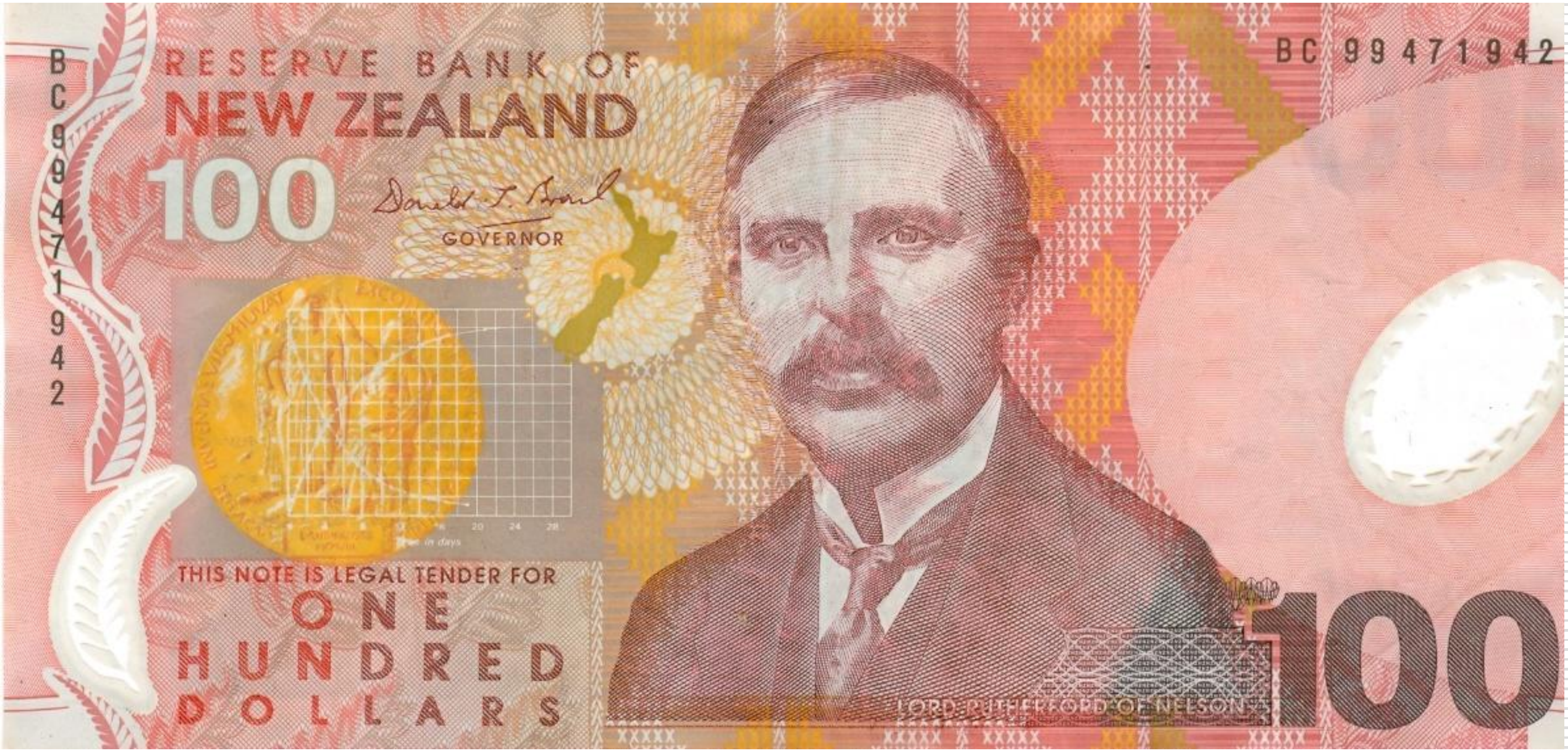
Thompson’s Model (1903)

- Basis for Thompson’s Model:
  - Cathode Rays - William Crooks observed that cathode rays could be deflected by a magnetic field.
  - Thompson discovered that cathode rays can be deflected by electric fields.   Consequently, cathode rays must be streams of particles, not radiation.
  - Thompson measured the charge to mass ratio for the particle and is consequently credited with the discovery of the cathode ray particle, the electron.
  - Atoms are electrically neutral. If they contain negative charge, they must also contain positive charge.

Rutherford’s Model (1911)

- Rutherford’s Model:
  - Positive charge is concentrated in a small dense nucleus. Particles with neutral charge also reside in the nucleus.
  - Negatively charged electrons are distributed about the nucleus.
- Basis for Rutherford’s Model:
  - The Rutherford alpha particle scattering experiment.
  - James Chadwick’s discovery of the neutron in 1932.

Ernest Rutherford



Atomic Notation

- Atomic Number - The number of protons in the nucleus of the atom. It is also equal to the number of electrons that the neutral atom contains (Z).
- Mass Number - The number of protons and neutrons in the nucleus (A).
- Isotopes - Atoms with the same number of protons, but different numbers of neutrons.

Atomic Notation

- Symbol for the element.
- Superscript indicating the mass number.
- Subscript indicating the atomic number.
- Examples:
  - $^{23}_{11}Na$
  - $^{19}_{9}F$
  - $^{109}_{47}Ag$

Atomic Mass

- Atomic Mass Unit - 1/12 the mass of the Carbon-12 atom.
- Atomic Mass - The weighted average of all of the naturally occurring isotopes of the element.
- Reasons for using atomic mass:
  - Atomic mass provides a convenient way of measuring a known number of atoms.
  - Most elements occur as a mixture of isotopes, so just using the mass number does not work for this purpose.

Weighted Averages

- Weighted Average Examples:
  - What is the average mass of a shotput?
    - 12 pound balls: 12 lbs 25% 3 lbs
    - 16 pound balls: 16 lbs 75% 12 lbs
    - Average weight: 15 lbs
  - What is the atomic mass of Carbon?
    - C-12: 12.000 AMU’s 98.89% 11.87 AMU’s
    - C-13: 13.003 AMU’s 1.11% .144 AMU’s
    - Atomic Mass of Carbon: 12.01 AMU’s

Atomic Mass and the Periodic Table

- Symbol in the center of the square.
- Atomic Number above symbol.
- Atomic Mass below the symbol.
- Mass Number - For radioactive elements, the mass number of the most stable or best known isotope is listed in place of the atomic mass. This number is enclosed in parentheses.

Characteristics of Light

- Wavelength - The distance between two identical points on a wave.
- Frequency - The number of times a wave repeats itself per second.
- Speed - The rated at which a wave travels through a medium.

Relationships Pertaining to Light

- The speed of light is constant: 3 X 10^8 m/s
- Frequency and wavelength are inversely proportional - as one increases, the other decreases.
- The wavelength of visible light is very small: 400 - 700 nm.
- The energy of light is directly proportional to its frequency.

The Electromagnetic Spectrum

- Radio Waves
- Microwaves
- Infrared
- Visible Light
- Ultraviolet
- X Rays
- Gamma Rays

The Particle Nature of Light

- Light is consists of a stream of particles called photons.
- Each photon has a discrete amount or quantum of energy.
- The amount of energy a photon has determines its frequency or color.
- When light is emitted, exactly the right amount of energy must be used to create the photon.
- When light is absorbed, energy must be released that is exactly equal to the amount of energy used to create the photon.

The Problem with Rutherford’s Model

- Objects that travel in curved paths accelerate.
- Charged objects that accelerate radiate energy.
- When an electron looses energy it moves closer to the nucleus.
- Atoms should be unstable because electrons should spiral down and collapse into the nucleus.

Bohr’s Model

- Bohr’s Model
  - Electrons can only be found at certain distances from the nucleus - they occupy certain energy levels.
  - Electrons can only emit or absorb energy when they jump from one energy level to another.
  - Electrons occupy the lowest possible energy levels.

Bohr’s Model

- Basis for Bohr’s Model
  - Hydrogen atoms emit light of certain colors only.
    - Red - Transition from energy level 3 to energy level 2
    - Bluegreen - Transition from energy level 4 to energy level 2
    - Violet - Transition from energy level 5 to energy level 2
  - Spectra of other elements
    - Helium
    - Neon

Energy Levels and Sublevels

- Main energy levels - 1, 2, 3, 4, 5, . . .
- Energy Sublevels - sharp, principle, diffuse, fine.
- Relationships among energy levels and sublevels
  - Number of sublevels:
    - Level 1 - 1 sublevel, s
    - Level 2 - 2 sublevels, s and p
    - Level 3 - 3 sublevels, s,p, and d
    - Level 4 - 4 sublevels, s. p. d. and f

Energy Levels and Sublevels

- Relationships among energy levels and sublevels
  - Maximum number of electrons:
    - s - 2 electrons
    - p - 6 electrons
    - d - 10 electrons
    - f - 14 electrons

Electron Configurations

- Hydrogen: $1s^1$
- Lithium: $1s^2 2s^1$
- Oxygen: $1s^2 2s^2 2p^4$
- Iron: $1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^6$

The Quantum Mechanical Model

- The Heisenberg Uncertainty Principle - It is impossible to measure both the precise position and the precise speed of an object.
  - Energy levels are not sharply defined, they are regions in space where the probability of finding an electron is high.
  - Because electrons can’t be thought of as traveling in fixed paths or orbits, these regions are called orbitals

The Quantum Mechanical Model

- The Pauli Exclusion Principle - Each orbital in the atom can be occupied by no more than 2 electrons, one with spin up and the other with spin down.