2.1 Matter and Structure
- Understanding the structure of the atom will help to understand the properties of the elements.
- Keep in mind that these, as all theories, are subject to constant refinement. The picture of the atom isn’t final.

2.2 Composition of the Atom
- **Atom** – the basic structural unit of an element.
  - The smallest unit of an element that retains the chemical properties of that element.
- **Radioactive Decay** – certain kinds of atoms can “split” into smaller particles and release large amounts of energy.
- Atoms consist of three **primary** particles.
  - Electrons
  - Protons
  - Neutrons
- **Nucleus** – small, dense, positively charged region in the center of the atom.
  - Contains:
    - Protons – positively charged particles
    - Neutrons – uncharged particles
- Surrounding the nucleus is a diffuse region of negative charge populated by:
  - Electrons – negatively charged particles
- **Atomic Number** – the number of protons in an atom
- **Mass Number** – sum of the number of protons and neutrons
- **Isotopes** – atoms of the same element having different masses.
  - Contain same number of protons
  - Contain different numbers of neutrons
  - Isotopes of the same element have identical chemical properties
  - Some isotopes are radioactive
- **Atomic Mass** – the weighted average of the masses of the isotopes that make up an element.
- The **weighted average** is an average corrected by the relative amounts of each isotope present in nature.
  - Chlorine consists of chlorine-35 and chlorine-37 in a 3:1 ratio.
  - Calculate the atomic mass of naturally occurring chlorine if 75.77% of chlorine atoms are chlorine-35 and 24.23% of chlorine atoms are chlorine-37.

  **Step 1:** Convert the percentage to a decimal fraction
  **Step 2:** Multiply the decimal fraction by the mass of that isotope to obtain the isotope contribution to the atomic mass.
  **Step 3:** Sum to get the weighted average atomic mass of chlorine.
• 2.2 Composition of the Atom
  o Ion – electrically charged particles that result from the gain or loss of one or more electrons by the parent atom.
    ▪ Results from the loss of electrons
  o Cation – positively charged.
  o Anion – negatively charged.
    ▪ Results from the gain of electrons

• 2.3 Development of the Atomic Theory
  o Dalton’s Atomic Theory – the first experimentally based theory of atomic structure of the atom.
    ▪ John Dalton
    ▪ Early 1800’s
  o Much of Dalton’s Theory is still regarded as correct today. (See starred items
  o Postulates of Dalton’s Atomic Theory
    ▪ All matter consists of tiny particles called atoms.*
    ▪ An atom cannot be created, divided, destroyed, or converted to any other type of atom.
    ▪ Atoms of a particular element have identical properties.
    ▪ Atoms of different elements have different properties.*
    ▪ Atoms of different elements combine in simple whole-number ratios to produce compounds (stable aggregates of atoms).*
    ▪ Chemical change involves joining, separating, or rearranging atoms.*

• Subatomic Particles: Electrons, Protons, and Neutrons
  o Electrons were the first subatomic particles to be discovered using the cathode ray tube.
  o Protons were the next particle to be discovered.
    ▪ Protons have the same size charge but opposite in sign.
    ▪ Proton is 1837 times as heavy as an electron.
  o Neutrons
    ▪ Postulated to exist in 1920’s but not demonstrated to exist until 1932.
    ▪ Almost the same mass as the proton.

• The Nucleus
  o The initial ideas of the atom did not have a “nucleus”
  o “Plum Pudding Model”
  o Earnest Rutherford’s “Gold Foil Experiment” led to the understanding of the nucleus.
  o Most of the atom is empty space.
  o Most of the mass is located in a small, dense region.

• 2.4 The Relationship between Light and Atomic Structure
  o Spectroscopy – absorption or emission of light by atoms.
    ▪ Used to understand the electronic structure.
  o To understand the electronic structure, we must first understand light and electromagnetic radiation.
  o Emission spectrum – light emitted when a substance is excited by an energy source.
The emission-spectrum of hydrogen leads to the modern understanding of the electronic structure of the atom.

2.5 The Bohr Atom
- Initial understanding of the atom by Niels Bohr
- Electrons exist in fixed energy levels surrounding the nucleus. The quantization of energy.
- Promotion of an electron occurs as it absorbs energy. The excited state.
- Energy is released as the electron travels back to lower levels. Relaxation.
- Orbit – what Bohr called the fixed energy levels.
- Ground state – the lowest possible energy state an electron can occupy.
- The orbits are also identified using “quantum numbers”: 1, 2, 3, …
- When the electron relaxes the energy released is observed as a single wavelength of light.

2.6 Modern Atomic Theory
- Bohr’s model of the atom when applied to atoms with more than one electron failed to explain their line spectra.
- One major change from Bohr’s model is that electrons do not move in orbits.
- Atomic orbitals – regions in space with a high probability of finding an electron.
- Electrons move rapidly within the orbital giving a high electron density in that region.