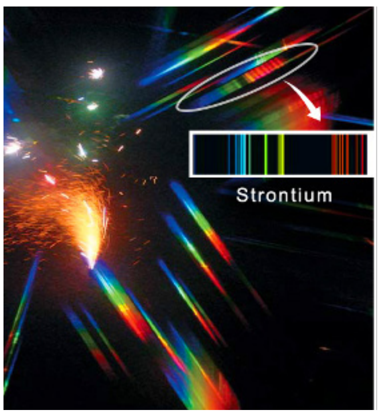
Chapter 4: Detailed Summary

Subatomic Particles



A *physical model* represents an object on a different scale. A toy airplane is a physical model of a real airplane. A *conceptual model* is used to describe the behavior of a system. Conceptual models of the weather that help us to predict whether it will rain or shine. Atoms are smaller than the wavelength of visible light. This makes it impossible for use to build an accurate physical model. An atom, however, is a system of interacting subatomic particles. As such, atoms are best described using conceptual models. These models are very useful in that they allow us to predict the behavior of atoms.

By the late 1800s it was realized that atoms themselves consist of even tinier particles. The first subatomic particle to be discovered was the *electron*. Its discoverer was the English physicist J.J. Thomson (1856–1940), who studied beams of electrons within cathode ray tubes. The mass of an electron being smaller than the smallest known atom was confirmed by the American physicist Robert Millikan (1868–1953) through his famous oil-drop experiment.

Other subatomic particles where then discovered within a matter of decades. Putting the pieces together in the first model of the inner structure of the atom was the New Zealand physicist Ernest Rutherford (1871–1937). Rutherford proposed that the atom is mostly empty space with most of its mass concentrated in a very tiny center he called the *atomic nucleus*.

Today we know the atomic nucleus to consist of *protons* and *neutrons*. Protons have a positive electric charge, while neutrons are neutral. Each of these subatomic particles is about 2000 times more massive than the electron, which has a negative charge and whizzes about the nucleus at high speeds.

Important terms that describe the arrangement of subatomic particles within an atom include *atomic number*, *isotope*, *mass number*, and *atomic mass*. Atomic number is the number of protons in a nucleus. Two atoms are "isotopes" of each other if they have the same atomic number but a different number of neutrons. The mass number of an atom is a count of the number of protons and neutrons, which are collectively called *nucleons*. The atomic mass is the average mass of all the isotopes of an element.

Modern conceptual models of the atom arose from the interaction between atoms and light. The chapter, therefore, spends some time discussing the nature of light, which also goes by the name *electromagnetic radiation* (EMR). Visible light is only a narrow region within the broad electromagnetic radiation spectrum. All EMR comes to us in the form of electromagnetic waves. We are able to decipher the different frequencies of light waves coming from a light source using the *spectroscope*.

When a spectroscope is pointed toward a glowing element, we see only particular frequencies of light. In fact, the frequencies of light emitted by an element are unique to that element and can be considered as the element's fingerprint. Early investigators noted some remarkable mathematical patterns within atomic spectra. These patterns, however, could not be explained until Niels Bohr proposed that the energy levels within an atom are *quantized*, which means that an electron within an atom can possess only particular quantities of energy. This, in turn, means that the atom can absorb or emit only particular frequencies of light.

Further work in quantum theory led to the discovery that electrons are not just particles of mass. Because of their kinetic energies, electrons also exhibit wavelike properties. It is wrong, therefore, to think of an electron merely as a particle orbiting the atomic nucleus much like a planet orbiting the Sun. Instead, the electron is a wave of matter and energy surrounding the nucleus. Only particular waves form because only particular waves are self-reinforcing. Each self-reinforcing wave corresponds to a particular energy level, as described by Bohr.

A simplified, yet powerful, conceptual model of the atom is the *shell model* in which electrons behave as though they are arranged in a series of shells concentric to the atomic nucleus. Each shell can only hold a limited number of electrons. For example, the first shell has a capacity for only two electrons. The second shell, however, has a capacity for eight electrons. Relative to the periodic table, each shell represents a period (horizontal row). Note that the number of elements within the periods corresponds to the electron capacities of the shells: two for the first shell, eight for the second shell, and so on.

Using this shell model, we can understand why it is that atoms get smaller in moving from left to right across the periodic table. In moving in this direction, the nuclear charge gets stronger. This has the effect of pulling electrons in closer to the nucleus. In moving down a group (vertical column), atoms of elements get larger because they have more filled shells. By a similar rationale, we learn that atoms to the upper right of the periodic table hold onto their electrons most tightly. The electrons of atoms towards the lower left of the periodic table, however, are more readily lost.

From this chapter you should realize that chemistry is more about applying conceptual models than it is about memorizing facts. With facts we can only regurgitate. With a good conceptual model, we can understand.

Summary of Terms

Atomic Mass The total mass of an atom. The atomic mass of each element presented in the periodic table is the *average* atomic mass of the various isotopes of that element occurring in nature.

Atomic Nucleus The dense, positively charged center of every atom.

Atomic Number The number of protons in the atomic nucleus of each atom of a given element.

Atomic Spectrum The pattern of frequencies of electromagnetic radiation emitted by the atoms of an element, considered to be an element's "fingerprint."

Conceptual Model A representation of a system that helps us predict how the system behaves.

Effective Nuclear Charge The nuclear charge experienced by outer-shell electrons, diminished by the shielding effect of inner-shell electrons and also by the distance from the nucleus.

Electromagnetic Spectrum The complete range of waves, from radio waves to gamma rays.

Electron An extremely small, negatively charged subatomic particle found outside the atomic nucleus.

Electron Configuration The arrangement of an atom's electrons within orbitals.

Energy-level diagram A schematic drawing used to arrange atomic orbitals in order of increasing energy levels.

Inner-Shell Shielding The tendency of inner-shell electrons to partially shield outer-shell electrons from the attractive pull exerted by the positively charged nucleus.

Ionization Energy The amount of energy needed to pull an electron away from an atom.

Isotope Any member of a set of atoms of the same element whose nuclei contain the same number of protons but different numbers of neutrons.

Mass Number The number of nucleons (protons plus neutrons) in the atomic nucleus. Used primarily to identify isotopes.

Neutron An electrically neutral subatomic particle found in atomic nuclei.

Nucleon Any subatomic particle found in an atomic nucleus. Another name for either proton or neutron.

Physical Model A representation of an object on some convenient scale.

Proton A positively charged subatomic particle in atomic nuclei.

Quantum A small, discrete packet of energy.

Quantum Number An integer that specifies the quantized energy level within an atom.

Shell A graphic representation of a collection of orbitals of comparable energy in a multielectron atom. A shell can also be viewed as a region of space about the atomic nucleus within which electrons may reside.

Spectroscope A device that uses a prism or diffraction grating to separate light into its color components and measure their frequencies.