Module 6 summary: Electronic structure & periodic properties

6.1: Electromagnetic energy

Light and other forms of **electromagnetic radiation** move through a vacuum with a constant speed, c, of 2.998 E8 m/s. This radiation shows wavelike behavior, which can be characterized by a **frequency, ν**, and a **wavelength, λ**, such that **c = λν**. Light is an example of a **travelling wave**. Other important wave phenomena include **standing waves**, periodic oscillations, and vibrations. Standing waves exhibit **quantization**, since their wavelengths are limited to **discrete integer multiples** of some characteristic lengths. Electromagnetic radiation that passes through two closely spaced narrow slits having dimensions roughly similar to the wavelength will show an interference pattern that is a result of **constructive and destructive interference** of the waves. Electromagnetic radiation also demonstrates properties of particles called **photons**. The energy of a photon is related to the frequency (or alternatively, the wavelength) of the radiation as **E = hν (or E = hc/λ)**, where h is Planck's constant. That light demonstrates both wavelike and particle-like behavior is known as **wave-particle duality**. All forms of electromagnetic radiation share these properties, although various forms including X-rays, visible light, microwaves, and radio waves interact differently with matter and have very different practical applications. Electromagnetic radiation can be generated by **exciting matter to higher energies**, such as by heating it. The emitted light can be either continuous (incandescent sources like the sun) or discrete (from specific types of excited atoms). Continuous spectra often have distributions that can be approximated as **blackbody radiation** at some appropriate temperature. The **line spectrum** of hydrogen can be obtained by passing the light from an electrified tube of hydrogen gas through a prism. This line spectrum was simple enough that an empirical formula called the **Rydberg formula** could be derived from the spectrum. <u>Three historically important paradoxes from the late 19th and early 20th centuries that could not be explained within the existing framework of classical mechanics and classical electromagnetism were the blackbody problem, the photoelectric effect, and the discrete spectra of atoms. The resolution of these paradoxes ultimately led to quantum theories that superseded the classical theories.</u>

6.2: The Bohr model

Bohr incorporated Planck's and Einstein's quantization ideas into a model of the hydrogen atom that resolved the paradox of atom stability and discrete spectra. The **Bohr model** of the hydrogen atom explains the connection between the quantization of photons and the quantized emission from atoms. Bohr described the hydrogen atom in terms of an electron moving in a circular orbit about a nucleus. He postulated that the electron was restricted to certain orbits characterized by discrete energies. Transitions between these allowed orbits result in the **absorption or emission of photons**. When an electron moves from a higher-energy orbit to a more stable one, energy is emitted in the form of a photon. To move an electron from a stable orbit to a more excited one, a photon of energy must be absorbed. Using the Bohr model, we can calculate the energy of an electron and the radius of its orbit in any one-electron system.

6.3: Development of quantum theory

Macroscopic objects act as particles. Microscopic objects (such as electrons) have properties of both a particle and a wave. Their exact trajectories cannot be determined. The

quantum mechanical model of atoms describes the three-dimensional position of the electron in a probabilistic manner according to a mathematical function called a wavefunction, often denoted as ψ. **Atomic wavefunctions are also called orbitals.** The squared magnitude of the wavefunction describes the distribution of the probability of finding the electron in a particular region in space. Therefore, atomic orbitals describe the areas in an atom where electrons are most likely to be found. An atomic orbital is characterized by **three quantum numbers**. The **principal quantum number, n,** can be any positive integer. The general region for value of energy of the orbital and the average distance of an electron from the nucleus are related to n. Orbitals having the same value of n are said to be in the same shell. The **angular momentum quantum number, l,** can have any integer value from 0 to n – 1. This quantum number describes the shape or type of the orbital. Orbitals with the same principal quantum number and the same l value belong to the same subshell. The **magnetic quantum number, ml**, with 2l + 1 values ranging from –l to +l, describes the orientation of the orbital in space. In addition, each electron has a **spin quantum number, ms**, that can be equal to ±12. No two electrons in the same atom can have the same set of values for all the four quantum numbers.

6.4: Electronic structure of atoms (electron configuration)

The relative energy of the subshells determine the order in which atomic orbitals are filled (1s, 2s, 2p, 3s, 3p, 4s, 3d, 4p, and so on). Electron configurations and orbital diagrams can be determined by applying the **Pauli exclusion principle** (no two electrons can have the same set of four quantum numbers) and **Hund's rule** (whenever possible, electrons retain unpaired spins in degenerate orbitals). Electrons in the outermost orbitals, called **valence electrons**, are responsible for most of the chemical behavior of elements. In the periodic table, elements with analogous **valence electron configurations** usually occur within the same group. There are some exceptions to the predicted filling order, particularly when half filled or completely filled orbitals can be formed. The periodic table can be divided into three categories based on the orbital in which the last electron to be added is placed: **main group elements** (s and p orbitals), **transition elements** (d orbitals), and **inner transition elements** (f orbitals).

6.5: Periodic variations in element properties

Electron configurations allow us to understand many periodic trends. **Covalent radius** increases as we move down a group because the n level (orbital size) increases. Covalent radius mostly decreases as we move left to right across a period because the effective nuclear charge experienced by the electrons increases, and the electrons are pulled in tighter to the nucleus. **Anionic radii** are larger than the parent atom, while **cationic radii** are smaller, because the number of valence electrons has changed while the nuclear charge has remained constant. **Ionization energy** (the energy associated with forming a cation) decreases down a group and mostly increases across a period because it is easier to remove an electron from a larger, higher energy orbital. **Electron affinity** (the energy associated with forming an anion) is more favorable (exothermic) when electrons are placed into lower energy orbitals, closer to the nucleus. Therefore, electron affinity becomes increasingly negative as we move left to right across the periodic table and decreases as we move down a group. For both IE and electron affinity data, there are exceptions to the trends when dealing with completely filled or half-filled subshells