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Chapter 11 Modern Atomic Theory

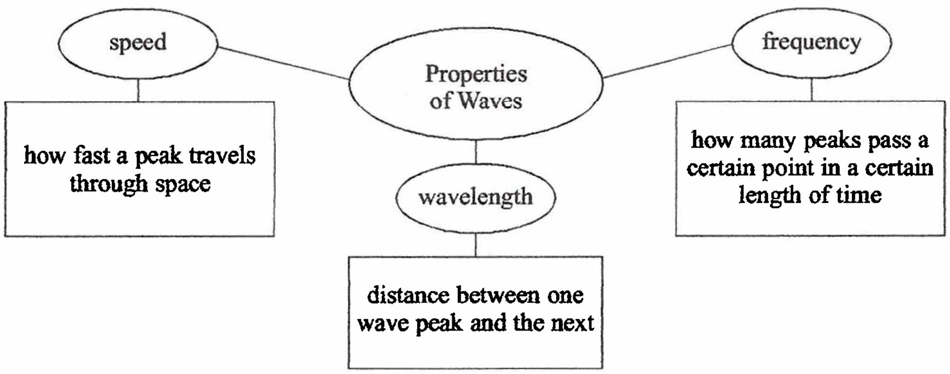
Summary

11.1 Rutherford's Atom

In Rutherford's concept of the nuclear atom, negatively charged electrons revolve around the nucleus as planets revolve around the sun.

11.2 Energy and Light

Energy can be transmitted by *electromagnetic radiation,* which travels through space in the form of waves. The various types of electromagnetic radiation differ in their wavelengths and amounts of energy. Each color of visible light has a different wavelength. Light also has properties that are more like those of particles rather than waves. A beam of light can therefore be thought of as a stream of tiny packets of energy called *photons.*



11.3 Emission of Energy by Atoms

When an atom receives energy from some source, it becomes excited, and it can release this energy by emitting (giving off) light, which is carried away by a photon. The energy of the photon is exactly the same as the energy change of the emitting atom. Short-wavelength light has high-energy photons, and long-wavelength light has low-energy photons.

11.4 The Energy Levels of Hydrogen

When hydrogen atoms are excited, only certain colors of visible light are emitted, that is, only certain types of photons are produced. This tells us that only certain energy changes are occurring. The hydrogen atom must have certain discrete (separate and unique) energy levels. Excited hydrogen atoms always emit photons with the same discrete colors (wavelengths). They never emit photons with energies in between. We can conclude that all hydrogen atoms have the same set of discrete energy levels. The energy levels of atoms are *quantized,* which means that only certain values are allowed.

11.5 The Bohr Model of the Atom

The Bohr model of the hydrogen atom showed that an electron moves in circular orbits that match the allowed orbits. The Bohr model worked well for hydrogen, but not for other atoms.

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11.6 The Wave Mechanical Model of the Atom

According to the *wave mechanical model* of the atom, an electron has the characteristics of both waves and particles. Electron states are described by orbitals. *Orbitals* are defined as probability maps, which are collections of points that indicate how likely it is to find the electron at a given position in space. The wave mechanical model can describe only the probabilities, or likelihoods, of finding the electron in any given point in space around the nucleus. The probability of the electron's presence is highest closest to the positive nucleus.

11.7 The Hydrogen Orbitals

The probability map for a hydrogen electron is called an orbital. The orbital size can be thought of as a surface containing 90 percent of the total electron probability. The electron can be found inside the sphere 90 percent of the time.

The discrete energy levels are called *principal energy levels,* and they are labeled with integers (whole numbers) symbolized by n. Level 1 is $n = 1$, level 2 is $n = 2$, and so on. The energy of the level increases as the value of n increases. Each of these levels is divided into *sublevels,* or types of orbitals. The number of sublevels present in a principal energy level equals n.

Orbitals have different shapes. They can be spherical (labeled s) or two-lobed (labeled p). When labeling an orbital, the n value is followed by a letter that indicates the type (shape) of the orbital. The designation $3p$ stands for an orbital in level 3 that has two lobes.

11.8 The Wave Mechanical Model: Further Development

All electrons appear to spin as a top spins on its axis. An electron can spin in only one of two directions. An orbital can be empty or it can contain one or two electrons, but never more than two. According to the *Pauli exclusion principle,* an atomic orbital can hold a maximum of two electrons, and those electrons must have opposite spins.

11.9 Electron Arrangements in the First Eighteen Atoms on the Periodic Table

For the first eighteen elements, the sublevels fill in the following order: $1s$, $2s$, $2p$, $3s$, $3p$. An electron is always most attracted to the $1s$ orbital because in this orbital the negatively charged electron is closer to the positively charged nucleus than in any other orbital. As n increases, the orbital becomes larger. The electron, on average, occupies space farther from the nucleus.

The arrangement of electrons in an atom is called its *electron configuration.* The hydrogen atom, with its lone electron in the $1s$ orbital, has an electron configuration of $1s^1$. This same configuration can be represented by an *orbital diagram.*

Valence electrons are those in the outermost, or highest, principal energy level of an atom. Valence electrons are those involved when atoms form bonds. The inner electrons are called *core electrons.* They are not involved in bonding.

11.10 Electron Configurations and the Periodic Table

The principles of orbital filling for elements from nineteen (potassium) on are different from those of the first eighteen. In these elements a higher principal energy level starts to fill before the previous one is complete. In a principal energy level that has d orbitals, the s orbital from the next level fills before the d orbitals in the current level. The $(n + 1)s$ orbitals always fill before the nd orbitals. Figure 11.32 shows the order in which atomic orbitals fill.

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Elements with similar properties repeat in a particular pattern on the periodic table because the same types of orbitals repeat as one principal energy level progresses to the next. As a result particular patterns of valence electrons repeat periodically. Elements with similar valence-electron patterns show very similar chemical behavior.

11.11 Atomic Properties and the Periodic Table

The basic classification of the chemical elements is into metals and nonmetals. *Metals* tend to lose electrons to form positive ions and are found at the left and center of the periodic table. The most chemically active metals have electrons in orbitals farthest away from the nucleus. *Nonmetals* tend to gain electrons to form negative ions. The most chemically active nonmetals can most easily pull electrons from metals.

The size of an atom is related to the number of protons in its nucleus. Atoms with their outermost electrons in the same principal energy level get smaller from left to right across a period (horizontal row on the periodic table). This is because their number of protons increases, causing an increase in positive charge, which pulls electrons closer to the nucleus.

Ionization energy, the energy required to remove an electron from a gaseous atom, decreases going down a group and increases going from left to right across a period.

Focus Questions

1. What are the three basic properties of waves?
2. Which kind of light has higher energy photons—short-wavelength light or long-wavelength light?
3. What does it mean to say that all atoms are quantized?
4. What is a probability map?
5. Where is the probability of an electron's presence highest?
6. What is the term for the probability map for an electron?
7. What is the relationship between a principal energy level and a sublevel?
8. Describe the Pauli exclusion principle.
9. What is the arrangement of electrons in an atom called?
10. For the first eighteen elements, in what order do the individual sublevels fill?
11. Which electrons occupy the outermost, or highest, principal energy level of an atom?
12. Describe the position of the electrons in the most chemically active metals.