# Chapter 5: Electrons in Atoms Note Taking Guide Key

## Electrons in Atoms

### Light and Quantized Energy

Rutherford's nuclear model of the atom does not explain chemical reactions because it does not explain the nature of electrons or where they occur in the atom

Electronic structure of atoms is revealed by _____

1. The interaction of electrons with light
2. Analysis of spectra of the elements

### The Wave Nature of Light

Characteristics of waves

Wavelength ( $\lambda$ ): the distance between equivalent points on a wave

Frequency ( $\nu$ ): the number of waves passing a given point in one second

Amplitude: the height of a wave from origin to crest

Speed of Light ( c ): $2.997\ 924\ 58 \times 10^{8}$ m $s^{-1}$ in a vacuum

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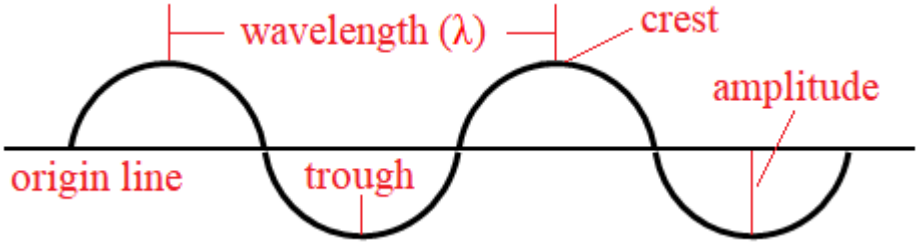
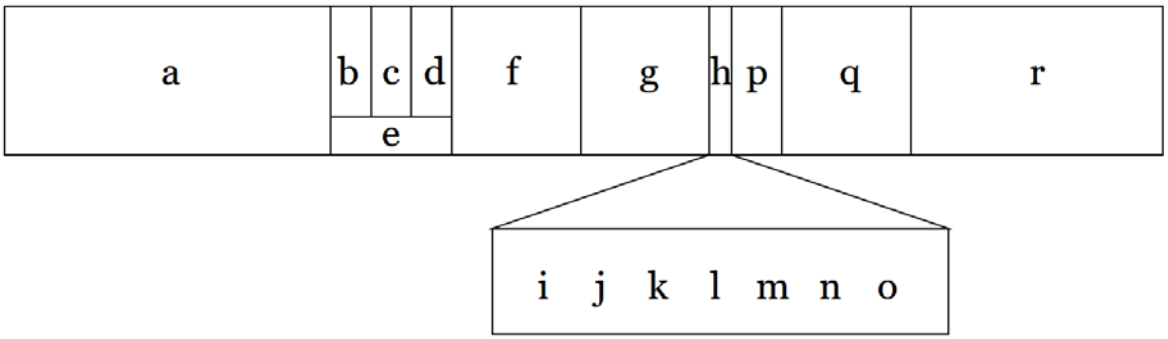
Show the following characteristics of waves on the diagram below: crest, trough, origin, wavelength, and amplitude




The Electromagnetic Spectrum




| | | | | | |
|---|---|---|---|---|---|
| a | long radio waves | b | AM radio waves | c | unused frequencies |
| d | FM radio waves | e | radio waves | f | microwaves |
| g | infrared light | h | visible light | i | red light |
| j | orange light | k | yellow light | l | green light |
| m | blue light | n | indigo light | o | violet light |
| p | ultraviolet light | q | x-ray light | r | gamma radiation |

Some properties that only waves exhibit

1. Refraction: bending of waves when they enter a new medium

   Example: a straw in water appears to be bent

2. Diffraction: bending of waves around objects

   Example: a stick in water is hit by a ripple

3. Interference: waves add up or cancel out

   Example: the wa-wa-wa sound when strings are tuned just out of tune

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The Particle Nature of Light

The Quantum Concept

In 1900 the German physicist Max Plank

studied blackbody radiation

Classical mechanics predicted an emitter could produce all wavelengths

but this is not possible because even one single infinitely short wavelength would

contain a nearly infinite amount of energy

Plank's quantum theory predicted an emitter produced small

packets of finite energy and as the total energy increased

more packets (or quanta) of higher energy photons (short wavelength or high

frequency) could be produced

In 1905 the German physicist Albert Einstein

proposed that quanta nicely explained the photoelectric effect

Photoelectric effect: incident light energy can eject an electron from a metal surface

Wave theory alone predicts that multiple waves should add energies by constructive

interference

which explains how lasers are produced.

But the energies ***do not combine*** indicating that light exists as discrete packets of

energy (called a quanta or a photons) instead of like one high energy wave

Therefore, light exists as both a wave and a particle at the same time.

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Atomic Emission Spectra

Unlike blackbody emitters which produce continuous spectra (the entire rainbow) ,

atoms produce an atomic emission spectrum (emission or absorption)

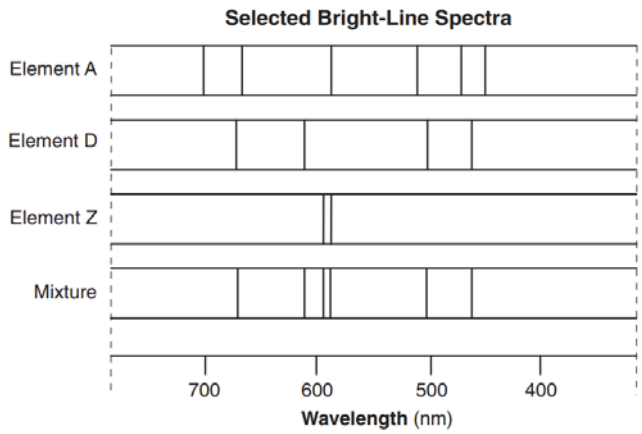
Atomic emission spectra are characteristic and like fingerprints

can be used to identify the elements

Example: Which element is not found in the mixture?

Element A is missing

Explain: None of the spectral lines found in element A show in the mixture




Bright-line spectra form when electrons in an electron shell which is farther

from the nucleus fall closer to the nucleus and give off (emit) energy in the form of

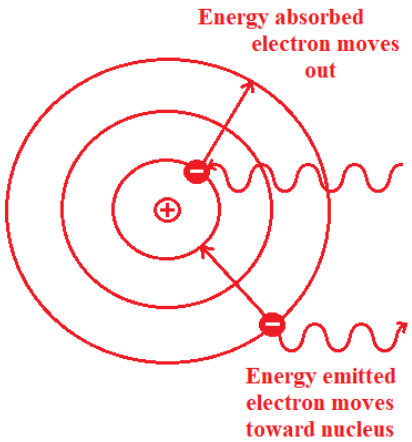
photons of the exact energy of the difference between the two electron shells

Electrons in a lower energy electron shell can absorb energy

and will move to a higher electron shell farther from the nucleus

and produce a dark line spectrum or absorption spectrum

Draw a sketch of a Bohr model and show electron excitation and electron relaxation. Be sure to indicate if energy is absorbed or emitted. Also add photons of light entering or leaving the atom to emphasize energy loss or gain.

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| History of the development of the quantum theory of the atom | | | |
|---|---|---|---|
| Name | Date | Theory | Experiment and/or Equation |
| Henrey GJ Moseley | 1913 | Atomic number (Z) – the number of protons | x-ray absorption edges show a linear increase in energy |
| Niels Bohr | 1913 | Planetary model with electron shells  | Explained bright line spectra of the elements |
| Louis de Broglie | 1924 | Electron wave theory describes an orbiting electron as a standing wave | Wrote the electron wave theory: $\lambda = (h/m\nu)$ |
| Werner Heisenberg | 1925 | Uncertainty principle implies no fixed electron paths | Cannot know electron velocity and position simultaneously |
| Erwin Schrödinger | 1926 | Probability wave function (shows the shapes of electron sublevels, s, p, d, f) | $\frac{ih}{2\pi}\frac{\partial\Psi}{\partial t} = -\frac{h^2}{8\pi^2 m}\left(\frac{\partial^2\Psi}{\partial x^2} + \frac{\partial^2\Psi}{\partial y^2} + \frac{\partial^2\Psi}{\partial z^2}\right) + V(x,y,z,t)\Psi$ |
| James Chadwick | 1932 | Discovered the neutron $^{9}_{4}\mathrm{Be} + ^{4}_{2}\mathrm{He} \rightarrow ^{13}_{6}\mathrm{C} \rightarrow ^{12}_{6}\mathrm{C} + ^{1}_{0}\mathrm{n}$ | Bombarded Be with alpha particles |

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Quantum Numbers

Rules for Quantum Numbers

1. Aufbau Principle – build from lowest energy to highest energy

   An atom with all the $e^-$ in lowest energy is in the ground state

   An atom with any $e^-$ in any higher electron shell is in the excited state

2. Hund's Rule – maximize multiplicity (each orbital gets 1 $e^-$ before any get two)

3. Pauli Exclusion Principle – no two $e^-$ can have the same four quantum numbers

Valence Electrons: electrons in the outermost electron shell (or energy level)

Electron Notations – There are many electron notations, the Regents uses only two

Electron configuration (energy level diagram)

$_{8}O$: 2 – 6

$_{34}Se$: 2 – 8 – 18 – 6

Lewis Electron Dot (LED) notation


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