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Chapter 5

Section 5.1 – Revising the Atomic Model

Models of the atom throughout history:

1. Dalton's "Billiard Ball" Model
    a. The atom is a tiny, indestructible particle with no internal structure.
2. Thomson's "Plum Pudding" Model
    a. The atom is a sphere of positive electrical charge with electrons evenly distributed throughout it
    b. Knew that the overall charge of an atom is neutral (electrons = protons)
3. Rutherford's Atomic Model
    a. Most of an atom's mass is concentrated in the small, positively charged nucleus. The electrons are in motion in the space around the nucleus.
    b. Limitations:
        i. It explained only a few simple properties of atoms
        ii. It couldn't explain the chemical properties of elements
4. Bohr's Atomic Model
    a. Electrons move in circular orbits at fixed distances from the nucleus.
    b. He changed Rutherford's model to incorporate newer discoveries about how the energy of an atom changes when the atom absorbs or emits light.
    c. He proposed that an electron is found only in specific circular paths, or orbits, around the nucleus
        i. Each possible electron orbit in Bohr's model has a fixed energy
            1. The fixed energies an electron can have are called energy levels.
            2. A quantum of energy is the amount of energy required to move an electron from one energy level to another energy level
                a. Similar to rungs on a ladder
                    i. A person cannot stand between the rungs. Electrons in an atom cannot exist between energy levels
                    ii. Energy levels are unequally spaced. Higher energy levels are closer together
        ii. The Rutherford model couldn't explain why elements that have been heated to higher and higher temperatures give off different colors of light. Bohr model explains

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how the energy levels of electrons in an atom change when the atom emits light.

5. Quantum Mechanical Model

a. Previous models described the motion of electrons in the same way as large objects move. The quantum mechanical model uses mathematical equations to describe the behavior of electrons.

i. Schrödinger used theoretical calculations to describe the behavior of the electron in a Hydrogen atom

ii. The modern description of the electrons in atoms came from the mathematical solutions to the Schrödinger equation.

b. Similar to Bohr model:

i. Quantum mechanical model restricts the energy of electrons to certain values

c. Different from Bohr model:

i. Doesn't specify an exact path the electron takes around the nucleus

d. The quantum mechanical model determines the allowed energies an electron can have and how likely it is to find the electron in various locations around the nucleus of an atom.

e. In the quantum mechanical model, the probability of finding an electron within a certain area surrounding the nucleus can be represented as a cloudlike region.

f. The Schrödinger equation gives the energy levels that an electron can have.

i. Atomic orbital – each energy level is represented as a region of space where there is a high probability of finding an electron

1. The energy levels of electrons are labeled by principal quantum numbers (n). Quantum numbers are assigned values (n=1, 2, 3, 4)
2. For each principal energy level greater than 1, there are several orbitals with different shapes and at different energy levels.
3. Forms energy sublevels.

ii. Each energy sublevel corresponds to one or more orbitals of different shapes. The orbitals describe where an electron is likely to be found.

iii. Different atomic orbitals are represented by letters

1. s orbitals are spherical

2. p orbitals are dumbbell shaped
3. d orbitals are balloon -shaped
   a. 4 of the 5 d orbitals have the same shape but different orientations in space

iv. The number and types of atomic orbitals depend on the principal energy level. It also determines the maximum number of electrons
1. Sublevels: equal to the quantum number, n
2. Number of orbitals: equal to n^2
3. Number of electrons: equal to 2n2

6. Section 1 Summary
   a. Energy levels in atoms. Electrons in atoms are found in fixed energy levels
      i. Bohr proposed that electrons move in specific orbits around the nucleus.
      ii. In these orbits, each electron has a fixed energy called an energy level.
      iii. A quantum of energy is the amount of energy needed to move an electron from one energy level to another.
   b. Quantum Mechanical Model. The quantum mechanical model determines how likely it is to find an electron in various locations around the atom.
      i. The quantum mechanical model is based on mathematics, not on experimental evidence.
      ii. This model does not specify an exact path an electron takes around the nucleus, but gives the probability of finding an electron within a certain volume of space around the nucleus.
      iii. This volume of space is described as an electron cloud, which has no boundary.

      The electron cloud is denser where the probability of finding the electron is high.
   c. Atomic Orbitals
      i. An atomic orbital describes where an electron is likely to be found.
      ii. Numbered outward from the nucleus, each energy level is assigned a principal quantum number, *n*, which is also the number of sublevels.

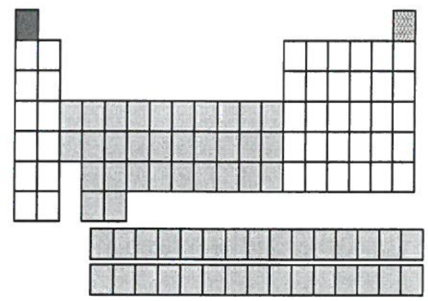
iii. Each energy sublevel differs in shape and orientation and contains orbitals, each of which can contain up to 2 electrons.

iv. Each energy level contains a maximum of $2n^2$ electrons.

Section 5.2 Electron Arrangement in Atoms

7. Electron Configuration

a. There are 4 "blocks" on the Periodic Table



i. s block Groups 1 & 2

ii. d block Groups 3 – 12

iii. p block Groups 13 – 18

iv. f block

Bottoms 2 rows of the Periodic Table

b. Each block can hold a specific amount of electrons

i. s block = 2 electrons

ii. p block = 6 electrons

iii. d block = 10 electrons

iv. f block = 14 electrons

c. For each element, we can write out the configuration of its electrons.

d. We can use the "Electron Configuration" to account for all of an atom's electrons.

i. A combination of letters and numbers are used.

1. The number will help us figure out which period it is in and the letter tells us which block it is in.

ii. Example – the electron configuration of Hydrogen is

$1s^1$ → one electron

1st period

s block

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	H 1 $1s^1$																	He 2 $1s^2$ *
$1s^2$	Li 3 $2s^1$	Be 4 $2s^2$ *										$1s^2 2s^2$	B 5 $2p^1$	C 6 $2p^2$	N 7 $2p^3$	O 8 $2p^4$	F 9 $2p^5$	Ne 10 $2p^6$ *
$1s^2 2s^2 2p^6$	Na 11 $3s^1$	Mg 12 $3s^2$ *										$1s^2 2s^2 2p^6 3s^2$	Al 13 $3p^1$	Si 14 $3p^2$	P 15 $3p^3$	S 16 $3p^4$	Cl 17 $3p^5$	Ar 18 $3p^6$ *
$1s^2 2s^2 2p^6 3s^2 3p^6$	K 19 $4s^1$	Ca 20 $4s^2$ *	Sc 21	Ti 22	V 23	Cr 24	Mn 25	Fe 26	Co 27	Ni 28	Cu 29	Zn 30 $3d^{10}$ *	Ga 31 $4p^1$	Ge 32 $4p^2$	As 33 $4p^3$	Se 34 $4p^4$	Br 35 $4p^5$	Kr 36 $4p^6$ *
	Rb 37 $5s^1$	Sr 38 $5s^2$ *	Y 39	Zr 40	Nb 41	Mo 42	Tc 43	Ru 44	Rh 45	Pd 46	Ag 47	Cd 48 $4d^{10}$ *	In 49 $5p^1$	Sn 50 $5p^2$	Sb 51 $5p^3$	Te 52 $5p^4$	I 53 $5p^5$	Xe 54 $5p^6$ *
	Cs 55 $6s^1$	Ba 56 $6s^2$ *	La 57	Hf 72	Ta 73	W 74	Re 75	Os 76	Ir 77	Pt 78	Au 79	Hg 80 $5d^{10}$ *	Tl 81 $6p^1$	Pb 82 $6p^2$	Bi 83 $6p^3$	Po 84 $6p^4$	At 85 $6p^5$	Rn 86 $6p^6$ *
	Fr 87 $7s^1$	Ra 88 $7s^2$ *	Ac 89															

Ce 58	Pr 59	Nd 60	Pm 61	Sm 62	Eu 63	Gd 64	Tb 65	Dy 66	Ho 67	Er 68	Tm 69	Yb 70	Lu 71 $4f^{14}$ *
Th 90	Pa 91	U 92	Np 93	Pu 94	Am 95	Cm 96	Bk 97	Cf 98	Es 99	Fm 100	Md 101	No 102	Lr 103 $5f^{14}$ *

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e. We can find the electron configuration for any element and write out a long chain to show all the "Stopping points" along the way.
   i. Superscripts (look like exponents) add up to the atomic number
f. 1st period
   i. H $1s^1$
   ii. He $1s^2$ — Stopping point for the 1s layer
g. 2nd period
   i. Li $1s^2 2s^1$
   ii. Be $1s^2 2s^2$ — Stopping point for the 2s layer
   iii. B $1s^2 2s^2 2p^1$ — Just entered the p block
   iv. C $1s^2 2s^2 2p^2$
   v. N $1s^2 2s^2 2p^3$
   vi. O $1s^2 2s^2 2p^4$
   vii. F $1s^2 2s^2 2p^5$
   viii. Ne $1s^2 2s^2 2p^6$ — Stopping point for the 2p layer (p=6 electrons)
h. 3rd period
   i. Mg $1s^2 2s^2 2p^6 3s^2$ — Stopping point for the 3s layer (s =2)
   ii. Ar $1s^2 2s^2 2p^6 3s^2 3p^6$ — Stopping point for the 3p layer (s=6)
i. 4th period
   i. Ca $1s^2 2s^2 2p^6 3s^2 3p^6 4s^2$
      1. Stopping point for 4s layer (s=2)
   ii. Zn $1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^{10}$
      1. Stopping point for 3d layer (d=10)
   iii. Kr $1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^{10} 4p^6$
      1. Stopping point for 4p layer (p=6)
j. 5th period
   i. Sr $1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^{10} 4p^6 5s^2$
      1. Stopping point for 5s layer (s = 2)
   ii. Cd $1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^{10} 4p^6 5s^2 4d^{10}$
      1. Stopping point for 4d layer (d=10)
   iii. Xe $1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^{10} 4p^6 5s^2 4d^{10} 5p^6$
      1. Stopping point for 5p layer (p=6)

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k. 6^{th} period

i. Ba $1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^{10} 4p^6 5s^2 4d^{10} 5p^6 6s^2$

1. Stopping point for 6s layer (s=2)

ii. Lu $1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^{10} 4p^6 5s^2 4d^{10} 5p^6 6s^2 4f^{14}$

1. Stopping point for 4f layer (f=14)

iii. Hg $1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^{10} 4p^6 5s^2 4d^{10} 5p^6 6s^2 4f^{14} 5d^{10}$

1. Stopping point for 5d layer (d=10)

iv. Rn $1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^{10} 4p^6 5s^2 4d^{10} 5p^6 6s^2 4f^{14} 5d^{10} 6p^6$

1. Stopping point for 6p layer (p=6)

8. Valence electrons

a. Valence electrons are those in the highest orbital levels. They are the electrons that are able to be gained, lost, or shared to form chemical bonds.

i. Example Chlorine (Cl)

1. $1s^2 2s^2 2p^6 \boxed{3s^2 3p^5}$ 2+5= Valence electrons = 7

ii. Example: Germanium (Ge)

1. $1s^2 2s^2 2p^6 3s^2 3p^6 \boxed{4s^2} 3d^{10} \boxed{4p^2}$ 2+2= Valence electrons = 4

b. Another way to find the number of valence electrons – look at the group number

i. s block only

1. Group 1 – all elements only have 1 valence electron
2. Group 2 – all elements only have 2 valence electrons

ii. p block only

1. Take the group number and subtract 10

a. Group 13 – all elements have 3 valence electrons
b. Group 14 – all elements have 4 valence electrons
c. Group 15 – all elements have 5 valence electrons
d. Group 16 – all elements have 6 valence electrons
e. Group 17 – all elements have 7 valence electrons
f. Group 18 – all elements have 8 valence electrons

i. 1 exception – Helium ($1s^2$)

1. It is in Group 18, but it only has a total of 2 electrons

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iii. d block (f block)

1. Look at the electron configuration
2. Add up the electrons in the highest levels
   a. Copper $1s^2 2s^2 2p^6 3s^2 3p^6 \boxed{4s^1} 3d^{10}$
      i. Has 1 valence electron
   b. Titanium $1s^2 2s^2 2p^6 3s^2 3p^6 \boxed{4s^2} 3d^2$
      i. Has 2 valence electrons

9. Noble gas Notation
   a. "Shortcut" Method
      i. Find the element whose electron configuration you need to write
      ii. Find the Noble Gas from the period above it (element in Group _____)
      iii. Place the symbol of the noble gas in brackets [ ]
      iv. Write out only the "stopping points" for the information that follows
         1. Silicon
            a. Electron Config $1s^2 2s^2 2p^6 3s^2 3p^2$
            b. Noble gas before it Neon
            c. Noble gas notation $[Ne]\ 3s^2 3p^2$
         2. Polonium (has 84 electrons)
            a. Noble gas before it Xenon
            b. Noble gas notation $[Xe]\ 6s^2 5d^{10} 4f^{14} 6p^4$
   b. The noble gas notation is listed for <u>every</u> element on our "Colorful" Periodic Table (at the bottom of each box)

Group 7

25	1.5
Mn	
Manganese 54.93805	
$[Ar]\ 4s^2 3d^5$	

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10. There are 3 rules that tell you how to find the electron configurations of atoms:
    a. Aufbau principle
        i. Electrons occupy the orbitals of lowest energy first.
    b. Pauli exclusion principle
        i. At atomic orbital may describe, at most, 2 electrons
        ii. To occupy the Same orbital, two electrons must have opposite spins (electron spins must be paired).
            1. Spin is a quantum mechanical property of electrons and may be thought of as clockwise or counterclockwise
            2. Arrows are used to indicate electrons and the direction of spin (↑or ↓)
            3. An orbital containing paired electrons is written as $\boxed{\uparrow\downarrow}$
    c. Hund's rule
        i. Electrons occupy orbitals of the Same energy in a way that makes the number of electrons with the same spin direction as large as possible.

11. Orbital Notation
    a. Each line can hold 2 electrons (Pauli exclusion principle)
    b. Three electrons (in the p orbital) would occupy three orbitals of equal energy (Hund's Rule) $\boxed{\uparrow}\boxed{\uparrow}\boxed{\uparrow}$

S orbitals	s^1	↑	s^2	↑↓
P orbitals	p^1	↑ __ __	p^4	↑↓ ↑ ↑
	p^2	↑ ↑ __	p^5	↑↓ ↑↓ ↑
	p^3	↑ ↑ ↑	p^6	↑↓ ↑↓ ↑↓
D orbitals				
	d^1	↑ __ __ __ __	d^6	↑↓ ↑ ↑ ↑ ↑
	d^2	↑ ↑ __ __ __	d^7	↑↓ ↑↓ ↑ ↑ ↑
	d^3	↑ ↑ ↑ __ __	d^8	↑↓ ↑↓ ↑↓ ↑ ↑
	d^4	↑ ↑ ↑ ↑ __	d^9	↑↓ ↑↓ ↑↓ ↑↓ ↑
	d^5	↑ ↑ ↑ ↑ ↑	d^{10}	↑↓ ↑↓ ↑↓ ↑↓ ↑↓

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F orbitals

f^1	↑	___	___	___	___	___	___	f^8	↑↓	↑	↑	↑	↑	↑	↑
f^2	↑	↑	___	___	___	___	___	f^9	↑↓	↑↓	↑	↑	↑	↑	↑
f^3	↑	↑	↑	___	___	___	___	f^{10}	↑↓	↑↓	↑↓	↑	↑	↑	↑
f^4	↑	↑	↑	↑	___	___	___	f^{11}	↑↓	↑↓	↑↓	↑↓	↑	↑	↑
f^5	↑	↑	↑	↑	↑	___	___	f^{12}	↑↓	↑↓	↑↓	↑↓	↑↓	↑	↑
f^6	↑	↑	↑	↑	↑	↑	___	f^{13}	↑↓	↑↓	↑↓	↑↓	↑↓	↑↓	↑
f^7	↑	↑	↑	↑	↑	↑	↑	f^{14}	↑↓	↑↓	↑↓	↑↓	↑↓	↑↓	↑↓

12. Paired and unpaired Electrons
    a. Easiest to see in orbital notation
        i. Paired electrons ↑↓
        ii. Unpaired electrons ↑
        iii. Example: Se [Ar] $4s^2 3d^{10} 4p^4$

↑↓ ↑↓ ↑↓ ↑↓ ↑↓ ↑↓ ↑↓ ↑ ↑

4s 3d 4p

1. So Se has 2 unpaired electrons

13. Remember that the sum of the superscripts equals the number of electrons in the atom. The arrows also equal the number of electrons in the atom.

Section 5.3 – Atomic Emission Spectra and the Quantum Mechanical Model

14. The Nature of Light – Wave Model
    a. Scientists thought that light consisted of waves
        i. Amplitude of a wave is the wave's height from zero to the crest
        ii. Wavelength is the distance between crests. Represented by the Greek letter lambda (λ)
        iii. Frequency is the number of wave cycles to pass a given point per unit of time. Represented by Greek letter nu (ν)
        iv. The SI unit of cycles per second is called the Hertz (Hz)

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b. The speed of light is the product of frequency and wavelength. The speed of light is a Constant, 2.998×10^8 m/s.

   i. $c = \lambda v$

c. Frequency (v) and wavelength (λ) of light are inversely proportional to each other.

   i. As the wavelength increases, the frequency decreases

   ii. As the frequency increases, the wavelength decreases

      1. Long wavelength = low frequency

      2. Short wavelength = high frequency

d. According to the wave model, light consists of electromagnetic waves

   i. Electromagnetic radiation includes radio waves, microwaves, infrared waves, visible light, ultraviolet waves, X-rays, and gamma rays

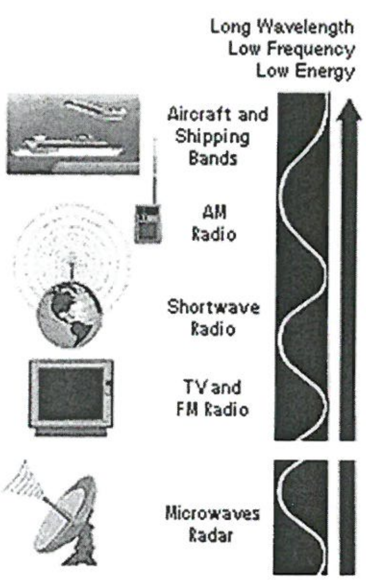
   ii. All electromagnetic waves travel in a Vacuum at a speed of 2.998×10^8 m/s

      1. Hotter, more energetic objects and events create higher energy radiation than cool objects.

e. Types of radiation, from longest wavelength to shortest:

1.

2.



1. Radio:
   - same kind of energy that radio stations emit into the air
   - Radio waves are also emitted by stars and gases in space.

2. Microwaves:
   - Used to cook popcorn in just a few minutes!
   - In space, microwaves are used by astronomers to learn about the structure of Milky Way galaxy and other galaxies

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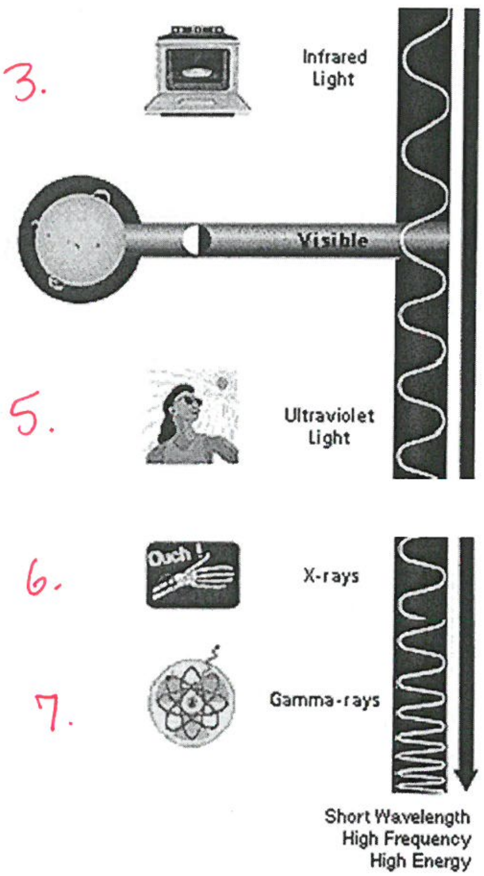

3.
4.
5.
6.
7.

3. Infrared:
- makes our skin feel warm
- In space, IR light maps the dust between stars.

4. Visible:
- the part that our eyes see
- Emitted by everything from fireflies to light bulbs to stars also by fast-moving particles hitting other particles.
- Order is ROYGBIV
  - Red
  - Orange
  - Yellow
  - Green
  - Blue
  - Indigo
  - Violet

5. Ultraviolet:
- a sunburn is the result of the Sun's UV rays
- Stars and other "hot" objects in space emit UV radiation.

6. X-rays:
- used to look at bones & teeth
- hot gases in the Universe also emit X-rays

7. Gamma rays:
- radioactive materials (either from natural sources or made by nuclear power plants) emit gamma-rays

15. Atomic Emission Spectra

a. The energy absorbed by an electron for it to move from its current energy level to a higher energy level is identical to the energy of the light emitted by the electron as it drops back to its original energy level.

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b. Each atom has a specific atomic emission spectrum that is characteristic for only that element. No 2 elements have the same emission spectrum.

16. Calculating frequency and wavelength

a. To calculate frequency (v):

i. 1st – convert the wavelength in nm to meters

ii. 2nd – plug into the formula $v = \frac{c}{\lambda}$

iii. 3rd – check significant digits and label as Hertz (Hz)

b. Example: Calculate the frequency of a light with a wavelength of 425 nm

1st $425\ nm \cdot \frac{1\ m}{1 \times 10^{-9}\ nm} = 4.25 \times 10^{-7}\ m$

2nd $v = \frac{c}{\lambda} = \frac{2.998 \times 10^{8}\ m/s}{4.25 \times 10^{-7}\ m} = 7.054117647 \times 10^{14}\ Hz$

3rd 3 digits
Label

$7.05 \times 10^{14}\ Hz$

c. To calculate wavelength (λ)

i. 1st – no conversion is needed, as long as the frequency is given in Hertz (Hz)

ii. 2nd – plug into the formula $\lambda = c/v$

iii. 3rd – check significant digits and label as meters (m)

d. Example: Calculate the wavelength of a light with a frequency of 3.05×10^{5} Hz

$\lambda = \frac{c}{v} = \frac{2.998 \times 10^{8}\ m/s}{3.05 \times 10^{5}\ Hz} = 982.9508197$

3 digits
Label

983 m

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17. The Nature of Light – Particle Model

a. The photoelectric effect – electrons are ejected when light shines on a metal.

b. Planck showed mathematically that the amount of radiant energy (E) of a single quantum absorbed or emitted by a body is proportional to the frequency of radiation (v).

   i. E = hv

      1. Planck's constant (h) has a value of 6.626×10^{-34} Js

c. Einstein used Planck's quantum theory to explain the photoelectric effect.

   i. The photoelectric effect couldn't be explained by classical physics (which described light as a form of energy). One assumption was that under weak light of any wavelength, an electron in a metal should eventually collect enough energy to be ejected. However, red light will not cause potassium to eject electron, no matter how intense the light. Yet a very weak yellow light shining on potassium begins the effect.

d. To explain the photoelectric effect, Einstein proposed that light could be described as quanta of energy that behave as if they were particles

   i. Light quanta are called photons

      1. Einstein's theory that light behaves as a stream of particles explains the photoelectric effect and many other observations.

   ii. Light behaves as a wave in some situations. Einstein concluded that light must have both wavelike and particle-like properties

e. The light emitted by an electron moving from a higher to a lower energy level has a frequency directly proportional to the energy change of the electron.

   i. Neon lights – each different gas has its own characteristic emission spectrum, creating different colors of light when excited electrons return to the ground state.