

Chapter 4

Development of a New Model

- Electrons behave like particles in some experiments, and like waves in others.
- The electron's 'wave/particle duality' has no real analogy in the everyday world.
- The *quantum theory* that describes the behavior of electrons is a cornerstone in modern chemistry.
- Quantum theory can be used to explain:
 - why atoms are stable,
 - why things have the color they do,
 - why the periodic table has the structure it does,
 - why chemical bonds form,
 - why different elements combine in different ratios with each other.

Properties of Light

- Light and electrons both behave quantum mechanically.

– **Waves**

- Waves are an oscillation that moves outward from a disturbance (ripples moving away from a pebble dropped into a pond)

- Properties of waves

property	definition	symbol	SI unit
velocity	distance traveled per second	c	m/s
amplitude	peak height above midline	A	varies with type of wave
wavelength	peak-to-peak distance	Lambda λ	m
frequency	number of peaks passing by per second	nu ν	s ⁻¹ (called Hertz)

- relationship between frequency and wavelength

- distance per cycle \times cycles per second = distance per second = c

$$C = \lambda \nu$$

- examples

- *The speed of sound in air is 330 m/s. Humans can hear sounds with wavelengths between 17 m and 17 mm. What is the highest sound frequency that is audible?*

- interference
 - **constructive interference:** amplitudes add
 - peaks, troughs of interfering waves occur in the same positions (waves are **in phase**)
 - **destructive interference:** amplitudes cancel
 - peaks of one wave are in same position as troughs of the other (waves are **out of phase**)

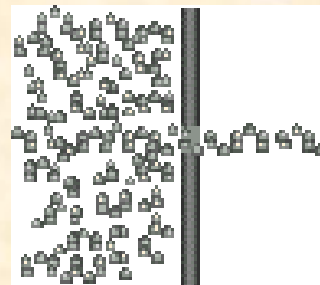
Diffraction

- The ability of a wave to bend around the edges of obstacles or holes. The effect is most noticeable when the obstacle or hole is comparable to the size of the wavelength



- Waves can bend around small obstacles and fan out from pinholes.

- particles effuse from pinholes.



- a wave can't bend around obstacles much larger than its wavelength
- what does this imply about the wavelength of sound waves? radio waves? visible light?
- **waves are delocalized** (spread out in space)

wave behavior	particle behavior
waves interfere	particles collide
waves diffract	particles effuse
waves are delocalized	particles are localized

Is light a stream of particles or a wave?

- Thomas Young, 1801
 - pass light through two tiny adjacent slits
 - if light were particles:
 - target would be brightest where light passing through the slits overlapped
 - target would darken steadily moving away from the overlap region
 - this was not observed!
 - a pattern of light and dark stripes was observed instead
 - Young explained the stripes as a combination of diffraction and interference
 - these **interference fringes** are a sure sign of wave behavior
 - White areas are peak-peak or trough-trough overlaps (constructive interference)
 - black areas are peak-trough overlaps (destructive interference).

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Force Fields

- **force field:** a region where forces act on an object; strength of forces vary with position
- gravitational fields
 - larger mass at center of field = stronger forces
 - larger distance from center of field = weaker forces
- electric fields
 - opposite charges attract each other, but like charges repel each other
 - larger charge at center of field = stronger forces
 - larger distance from center of field = weaker forces
- magnetic fields
 - can be produced by moving charges (electromagnets)
 - a moving magnetic field can produce an electric field (electric generator)

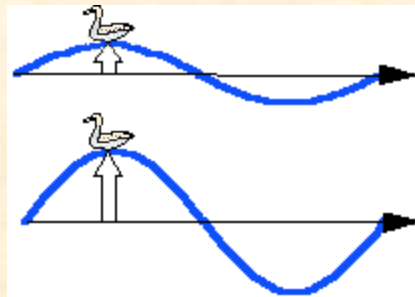
Electromagnetic radiation

- James Clerk Maxwell ca. 1855
 - changes in electric and magnetic fields are **always** coupled: electromagnetism
- making e/m waves with a vibrating charge
 - both electric and magnetic fields oscillate
 - oscillations are at right angles
 - electric oscillation produces magnetic oscillation, which produces another electric oscillation, ...and on and on
 - vibrating charge creates a ripple in the electromagnetic field
- **The speed of electromagnetic radiation was computed to be around 3×10^8 m/s**
- The same speed had been determined experimentally for light!
- hypothesis: light is a form of electromagnetic radiation (Maxwell, 1862)

Electromagnetic spectrum

Energy of electromagnetic radiation

- radiation carries energy through space
 - work is done on charges in the e/m field
 - transmitter loses energy; receiver gains energy



- higher amplitude means higher energy per peak
- amplitude squared determines the intensity or brightness of light
- therefore, brighter light should carry more energy per peak than dimmer light

Max Planck

- Did an experiment to measure the energy carried by an electromagnetic wave
 - **photoelectric effect**: shining light on alkali metals knocks electrons out of metal
 - strategy: measure kinetic energy of ejected electrons; then measure light energy per ejected electron.
 - surprise:
 - brightness has NO EFFECT on the kinetic energy per ejected electron
 - brighter light ejects MORE electrons.
 - surprise #2:
 - red light can't eject any electrons, but blue light can!
 - below a threshold frequency , there are no ejected electrons!
 - frequency is a property of the metal being used

Planck's mathematical equation

- Quantum (energy)= minimum quantity of energy that can be lost or gained by an atom
- Energy = planck's constant x frequency
- $E = h \nu$

$$h = 6.626 \times 10^{-34} \text{ js}$$

Albert Einstein

- Albert Einstein's interpretation of the photoelectric effect (Nobel Prize, 1921)
 - maybe light is like a stream of massless particles (call them **photons**)
 - brighter light has more photons, but bluer light has higher energy photons
- Changed Planck's work to:
- $E_{\text{photon}} = h \nu$

- examples
 - *What is the energy of a photon of red light with wavelength 700 nm?*
 - *What is the wavelength of a photon which has an energy of $1 \times 10^{-18} \text{ J}$?*
 - *Shining light of 400 nm on a metal causes electrons with a kinetic energy of $5 \times 10^{-19} \text{ J}$ to be ejected. What is the minimum energy required to eject an electron from the metal?*
- summary: light moves like a wave, but transfers energy like a stream of particles; the particles (photons) have energy equal to $h\nu$.

The collapsing atom paradox

- what's the electron doing in an atom?
- electrons within the atom can't be stationary:
 - positively charged nucleus will attract the negatively charged electron
 - electron will accelerate towards the nucleus
- if electrons within the atom move,
 - moving charges emit electromagnetic radiation
 - emission will cause electrons to lose energy and spiral into the nucleus
 - the atom will collapse!
- why don't atoms collapse?
 - classical physics has no answer!
 - key: electrons have wave/particle duality

Electrons as Waves

- the de Broglie hypothesis ([Nobel Prize, 1929](#))
 - connect wave and particle nature of matter using a relationship that applies to photons: $\lambda = h/p$ where p is the momentum of the particle ($p = \text{mass times velocity}$).
- experimental evidence of electron wave/particle duality
 - electron diffraction
 - C. J. Davisson and G. P. Thomson observed interference fringes when electron beams hit crystal surfaces and thin metal films (Nobel Prize, 1937)

Quantum Model of Atom

- Louis deBroglie
 - Electrons considered waves confined to the space around an atomic nucleus
 - Electrons can exist only at specific frequencies
 - Electrons have wave like properties
 - Created equation that demonstrates---anything that has mass and volume has wavelength

Werner Heisenberg

- E- detected by their interaction with photons
- Photons have the same energy as e-, any attempt to locate a specific e- with a photon will knock the e- off course
- Created Heisenberg uncertainty principle—“it is impossible to determine simultaneously both the position and velocity of an e- or any other particle.”

Erwin Schrodinger

- Used quantization of e- energies in equation, e- dual wave particle
- Quantum theory was derived by using Heisenberg Uncertainty Principle and Schrodinger wave equation
- E- are in orbitals (3-D)

Quantum Numbers

- Defn.: numbers that specify the properties of atomic orbitals and the properties of e⁻ in orbitals
- There are 4
 - Principal Quantum number
 - Symbol n
 - Main energy level occupied by an e⁻ (currently no more than 7)

– Angular Momentum number

- Symbol l
- Indicates the shape
 - S (sphere)
 - P(peanut)
 - D(dumbell)
 - F (flower)

– Magnetic quantum number

- Symbol m
- Orientation of an orbital
 - S : 1 orbital
 - P : 3 orbital
 - D: 5 orbital
 - F: 7 orbital

– Spin quantum number

- Symbol $+1/2$ or $-1/2$
- Indicates the two fundamental spins of the e^-

Electron Configuration Notation, Orbital Notations, Electron-dot notations

- Arrangement of e- using 4 quantum numbers
 - Atoms of each element have distinctive electron configurations
 - Atoms electrons tend to assume arrangements that have the lowest possible energies
- Rules governing:
 - Aufbau principle
 - States an electron occupies the lowest energy orbital that can receive it

– Hund's Rule

- States orbitals of equal energy are each occupied by one electron before any one orbital is occupied by a second electron. All electrons in a singly occupied orbitals must have same spin.

– Pauli Exclusion Principle

- States no two electrons in the same atom can have the same set of four quantum numbers
- Two values of spin quantum number permit two electrons of opposite spins to occupy the same orbital

s-block

1 New Designation

IA Original Designation

s-block

18

VIIIA

1	1 H 1.0094	2 He 4.00260	Non-Metals															
			13 III A	14 IV A	15 V A	16 VI A	17 VII A											
	p-block																	
2	3 Li 6.941	4 Be 9.0122	d-block Transition Metals										5 B 10.81	6 C 12.011	7 N 14.007	8 O 15.999	9 F 18.998	10 Ne 20.179
3	11 Na 22.990	12 Mg 24.305	3 IIIB	4 IVB	5 VB	6 VIB	7 VIIB	8 VIIIB	9 VIIIB	10 VIIIB	11 IB	12 IIB	13 Al 26.982	14 Si 28.086	15 P 30.974	16 S 32.06	17 Cl 35.453	18 Ar 39.948
4	19 K 39.098	20 Ca 40.08	21 Sc 44.956	22 Ti 47.88	23 V 50.942	24 Cr 51.996	25 Mn 54.938	26 Fe 55.847	27 Co 58.933	28 Ni 58.69	29 Cu 63.546	30 Zn 65.39	31 Ga 69.72	32 Ge 72.59	33 As 74.922	34 Se 78.96	35 Br 79.904	36 Kr 83.80
5	37 Rb 85.468	38 Sr 87.62	39 Y 88.906	40 Zr 91.224	41 Nb 92.906	42 Mo 95.94	43 Tc (98)	44 Ru 101.07	45 Rh 102.91	46 Pd 106.42	47 Ag 107.87	48 Cd 112.41	49 In 114.82	50 Sn 118.71	51 Sb 121.75	52 Te 127.60	53 I 126.91	54 Xe 131.29
6	55 Cs 132.91	56 Ba 137.33	57 to 71	72 Hf 178.49	73 Ta 180.95	74 W 183.85	75 Re 186.21	76 Os 190.2	77 Ir 192.22	78 Pt 195.08	79 Au 196.97	80 Hg 200.59	81 Tl 204.38	82 Pb 207.2	83 Bi 208.98	84 Po (209)	85 At (210)	86 Rn (222)
7	87 Fr (223)	88 Ra 226.03	89 to 103	104 Unq (261)	105 Unp (262)	106 Unh (263)	107 Uns (262)	108 Uno (265)	109 Uue (266)	110 Uun (267)	(Mass Numbers in Parentheses are from the most stable of common isotopes.)						Phases Solid Liquid Gas	
	Metals																	

Rare Earth

Elements

Lanthanide Series

Actinide Series

	d-block														f-block						
57	58	59	60	61	62	63	64	65	66	67	68	69	70	71							
La 138.91	Ce 140.12	Pr 140.91	Nd 144.24	Pm (145)	Sm 150.36	Eu 151.96	Gd 157.25	Tb 158.93	Dy 162.50	Ho 164.93	Er 167.26	Tm 168.93	Yb 173.04	Lu 174.97							
89	90	91	92	93	94	95	96	97	98	99	100	101	102	103							
Ac 227.03	Th 232.04	Pa 231.04	U 238.03	Np 237.05	Pu (244)	Am (243)	Cm (247)	Bk (247)	Cf (251)	Es (252)	Fm (257)	Md (258)	No (259)	Lr (260)							

- Each block contains a number of columns equal to the number of electrons that can occupy that subshell
 - The s-block has 2 columns, because a maximum of 2 electrons can occupy the single orbital in an s-subshell.
 - The p-block has 6 columns, because a maximum of 6 electrons can occupy the three orbitals in a p-subshell.
 - The d-block has 10 columns, because a maximum of 10 electrons can occupy the five orbitals in a d-subshell.
 - The f-block has 14 columns, because a maximum of 14 electrons can occupy the seven orbitals in a f-subshell.

Do orbital notation, electron-configuration notation, Nobel-Gas Notation and electron dot notation for each of the following:

Rows 1, 3, 5 Do these

- Chlorine
- Terbium
- Iron
- Silver
- Iodine

Rows 2 and 4 Do these

- Krypton
- Phosphorus
- Cerium
- Aluminum
- Lead