$$c = 8<$$

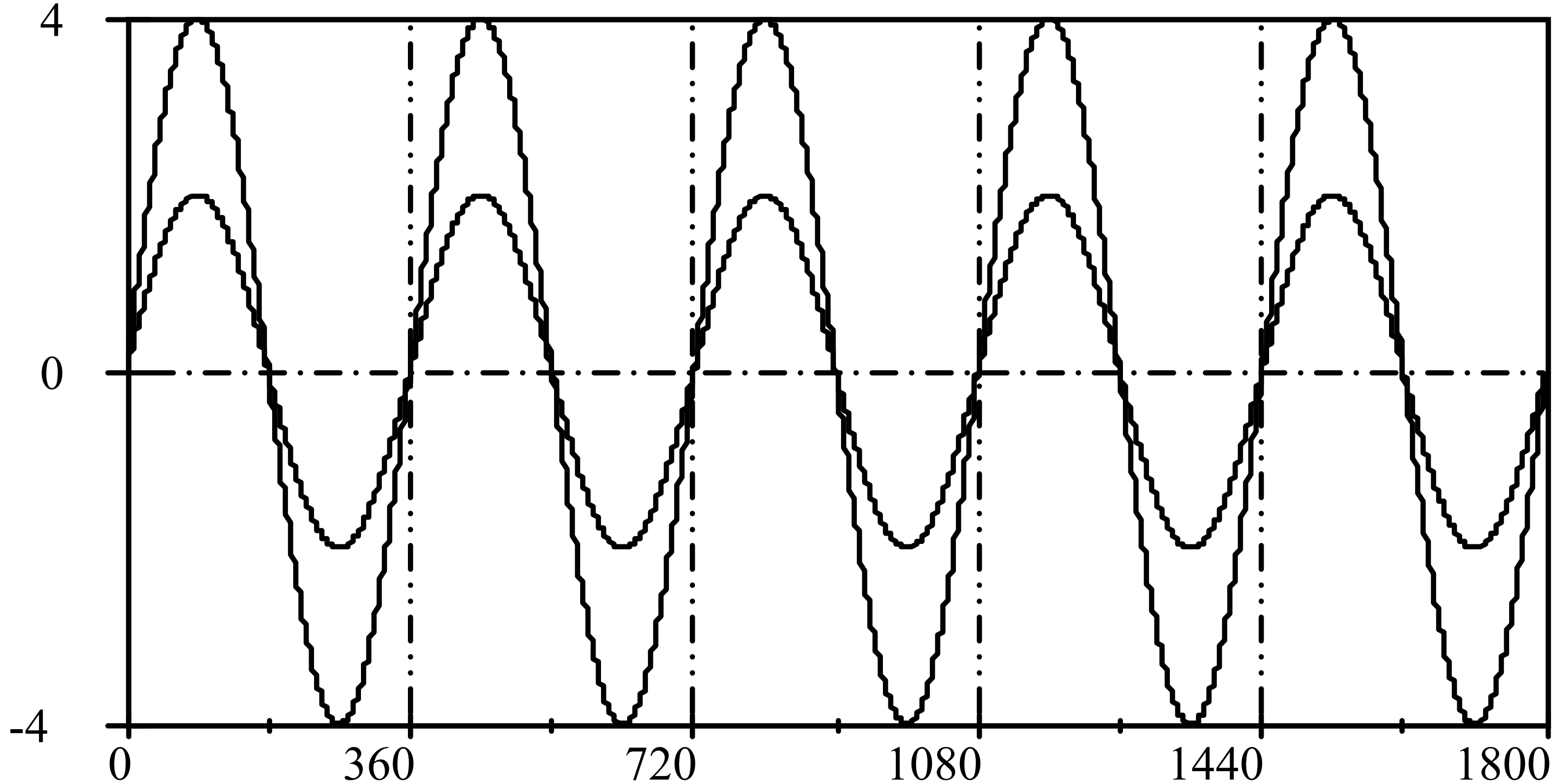
$$)E = h<$$

$$\lambda = \frac{h}{mv}$$

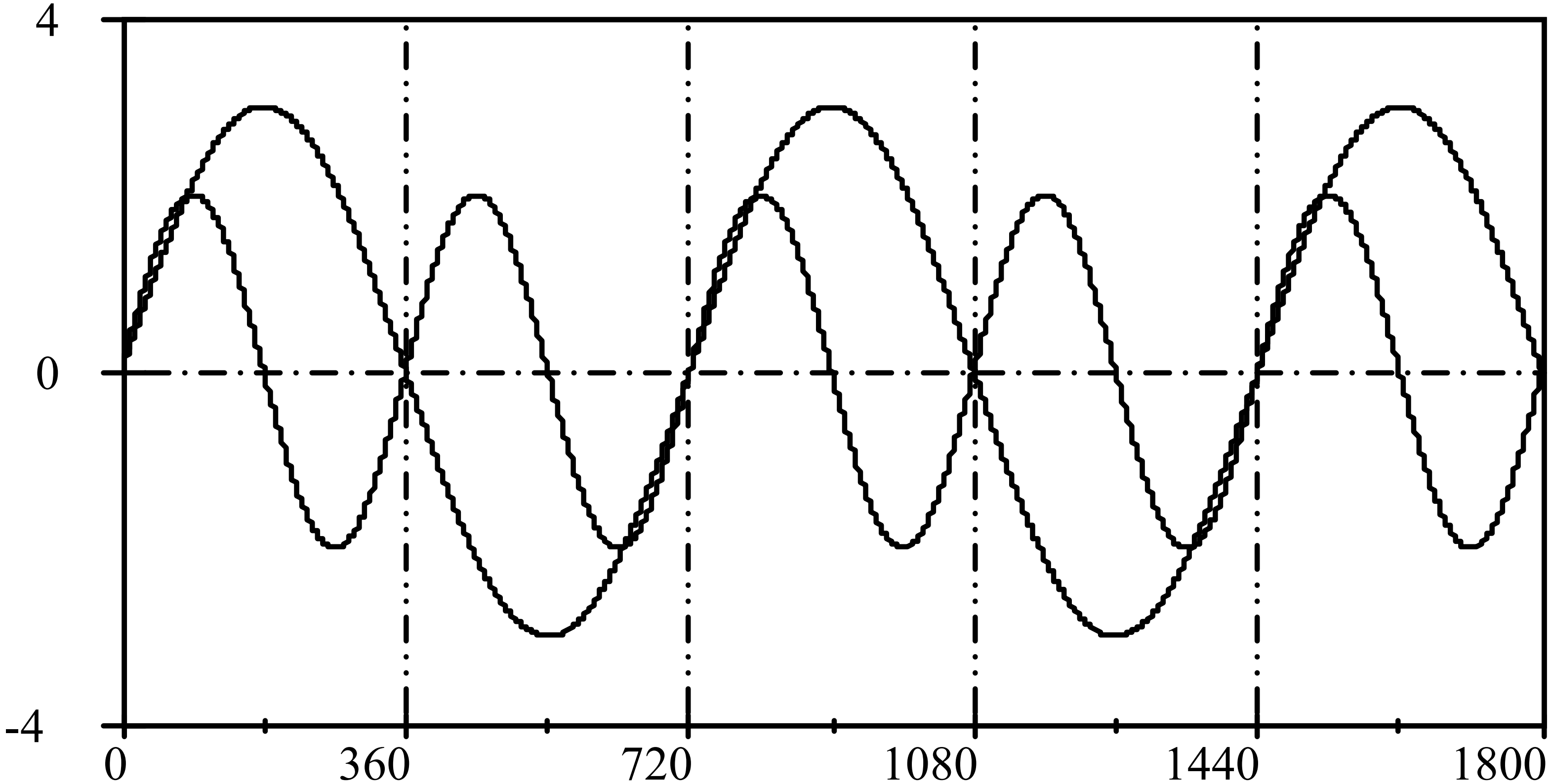
$$p = mv$$

$$E_a = \frac{-2.178 \text{ x } 10^{-18}}{n^2} \text{ joule}$$

Waves with different amplitudes



Waves with different wavelengths



Calculate the frequency of a visible line of light with a wavelength of 500 nm.

C = 8<

c = speed of light, 3.00 x 10^{8} m/s
< = frequency of the light wave, Hz or 1/sec^{-1}
8 = wavelength, meters

3.00 x 10^{8} m/s = 500 nm x <
3.00 x 10^{8} m/s = 500 x 10^{-9} m x <

< = 6.00 x 10^{14} hz

A short wave UV light has a wavelength of 350 nm. Calculate the energy of one 350 nm photon, of one mole of 350 nm photons.

C = 8<

3.00×10^{8} m/s = < x 350×10^{-9} m.
< = 8.57×10^{14} hz

)E = *h*<

)E = change in energy, in joules
h = Planck's constant, 6.626×10^{-34} J@s
< = frequency, is Hertz (hz or s^{-1})

)E = (6.626×10^{-34} J@s)(8.57×10^{14} s^{-1}/photon)
)E = 5.69×10^{-19} J/photon

5.69×10^{-19} J x 6.023×10^{23} $mole^{-1}$ = 3.42×10^{5}J
)E = 342 kJ/mol

Calculate the energy liberated when one mole of photons are produced at 656.3 nm. The red line in the hydrogen spectra.

C = 8<

3.00 x 10^{8} m/s = < x 656.3 x 10^{-9} m.
< = 4.57 x 10^{14} hz

) E = *h*<

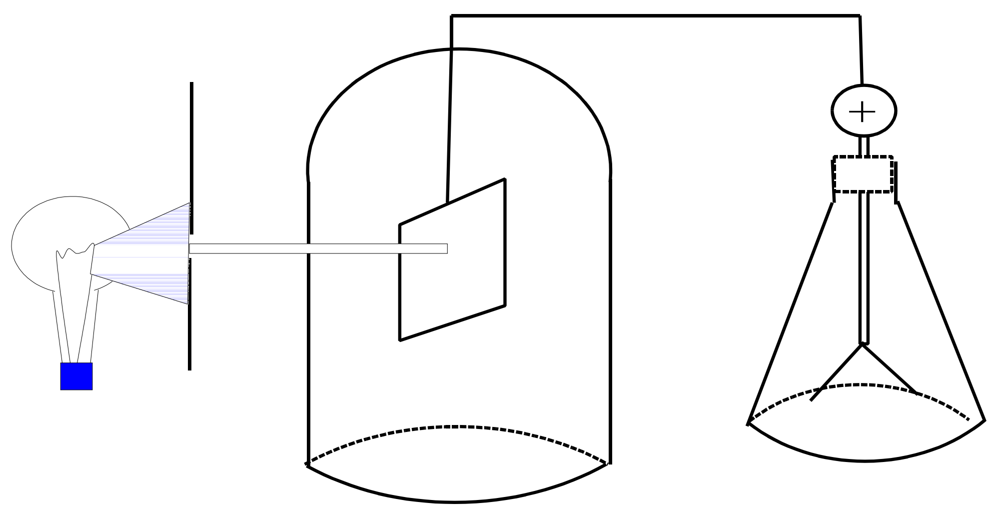
) E = (6.626 x 10^{-34} J@s)(4.57 x 10^{14} s^{-1}/photon)
) E = 3.03 x 10^{-19} J/photon

3.03 x 10^{-19} J x 6.023 x 10^{23} $mole^{-1}$ = 1.82 x 10^{5}J
) E = 182 kJ/mol

The Photoelectric Effect

The photoelectric effect was first observed by H. Hertz in 1887. Hertz noticed that when a beam of light strikes a metal surface, electrons were emitted from the metal surface leaving the metal positively charged.

The energy of these emitted photoelectrons is dependent on the energy of the incident light. The number of photoelectrons is related to the intensity if the incident light.



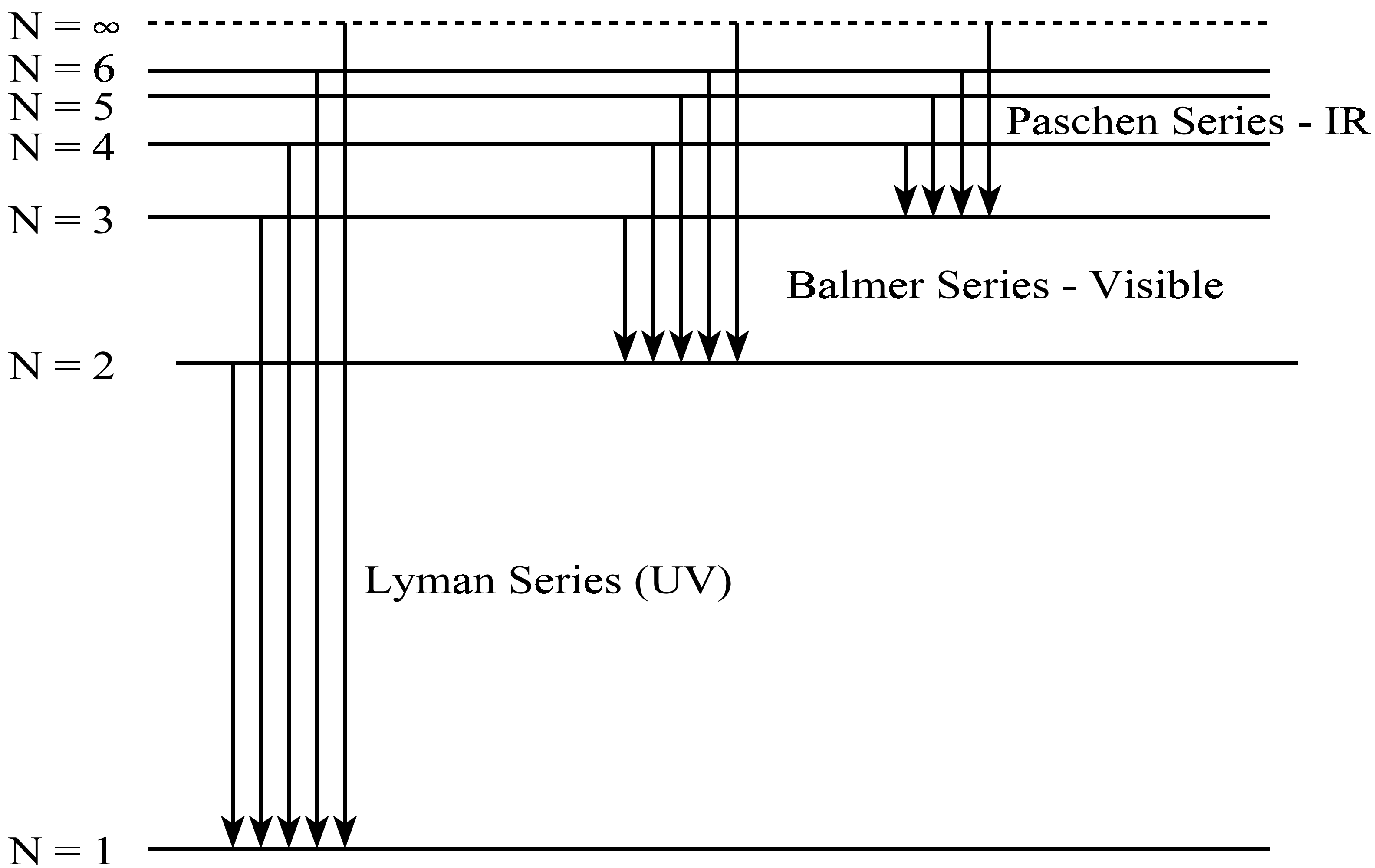
Albert Einstein won the Noble Prize for his explanation of the photoelectric effect.

$$\Delta E = h\upsilon - w$$

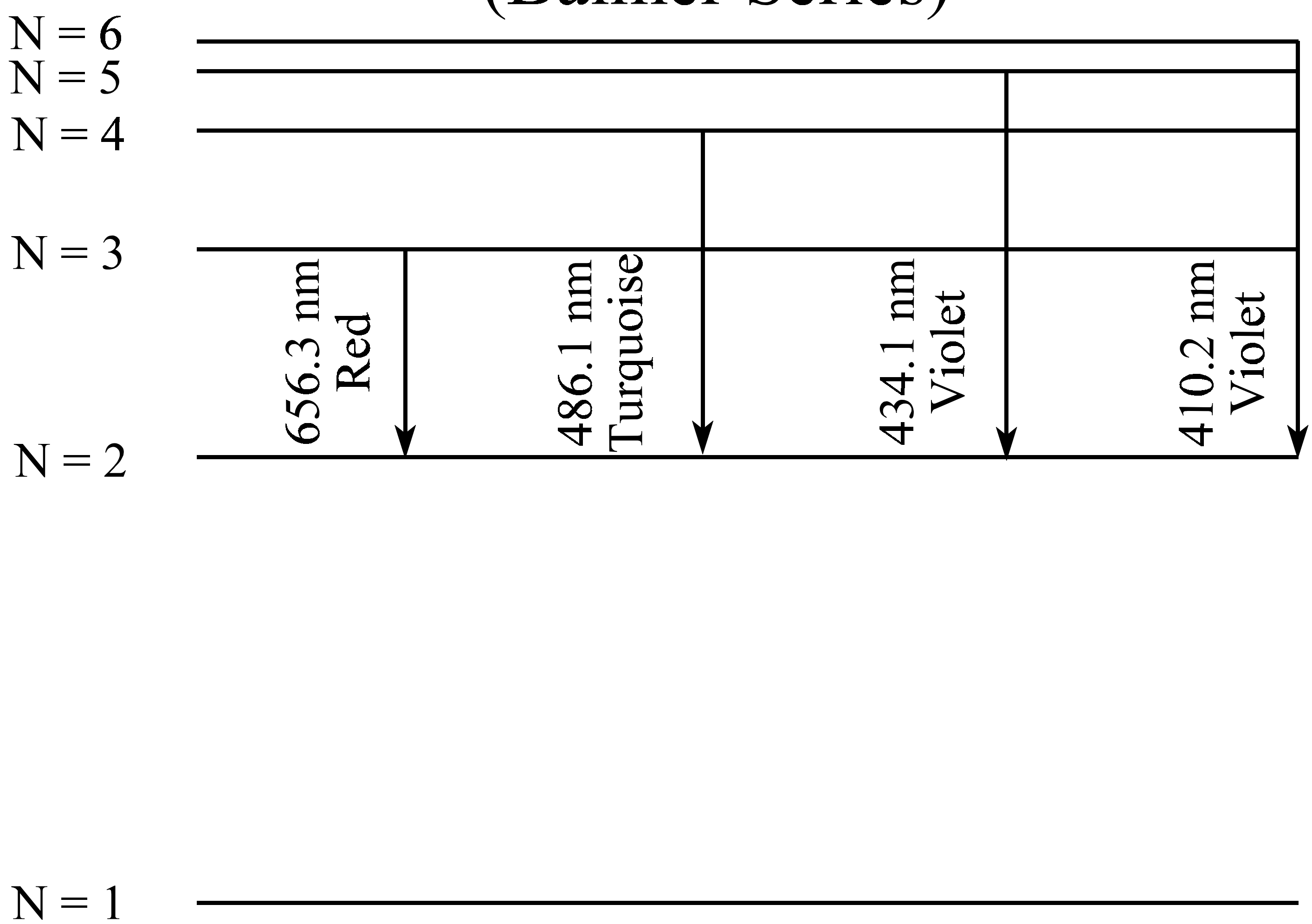
The energy of the emitted photoelectron equals the energy of the incident photon, *h*<, minus the work function, *w* (the ionization energy of the metal).

The photoelectric effect demonstrated the particle nature of light.

The Emission Spectra of Hydrogen



Visible Line Spectra of Hydrogen
(Balmer Series)



Calculate the energy of both the 2nd and 3rd energy levels in hydrogen in Joules per photon.

$$E_n = \frac{-2.178 \times 10^{-18}}{n^2} J$$

$$E_2 = \frac{-2.178 \text{ x } 10^{-18} J}{2^2} = -5.45 \text{ x } 10^{-19} J$$

$$E_3 = \frac{-2.178 \text{ x } 10^{-18} J}{3^2} = -2.42 \text{ x } 10^{-19} J$$

Calculate the wavelength of the photon produced when an electron drops from the third energy level to the second energy level in hydrogen.

$$)\,E = E_3 - E_2$$

$$)\,E = (-2.42 \text{ x } 10^{-19} \text{ J}) - (-5.45 \text{ x } 10^{-19} \text{ J}) = 3.03 \text{ x } 10^{-19}\text{J}$$

$$)\,E = h<$$

$$3.03 \text{ x } 10^{-19} \text{ J} = (6.626 \text{ x } 10^{-34}\text{J@s})<$$
$$< = 4.57 \text{ x } 10^{14} \text{ hz}$$

$$c = <8$$

$$3.00 \text{ x } 10^{8} \text{ m/s} = (4.57 \text{ x } 10^{14} \text{ s}^{-1})8$$
$$8 = 6.56 \text{ x } 10^{-7} \text{ m} = 656 \text{ nm}$$

<u>The Wave Properties of Matter</u>

Louis Victor de Broglie (1892-1987) reasoned that if light can have particle properties, then why can't particles have wave properties. In 1925, de Broglie proposed that a free electron of mass m, moving with a velocity of *v*, should have an associated wavelength according to the equation:

$$\lambda = \frac{h}{m\upsilon}$$

It was later discovered that de Broglie was correct when electrons were found to be diffracted by a thin piece of metal foil and became the basis for the electron microscope.

Why don't we observe the wave properties of matter?

A 114-gram baseball traveling at 90 mph (. 40 m/s) would have a wavelength of only 1.4×10^{-34} meters, too small to be detected.

What is the wavelength of an electron (mass = 9.1×10^{-31} kg) traveling at 40.0% the speed of light?

$$\lambda = \frac{h}{m\upsilon} = \frac{6.626 \times 10^{-34}\, J \cdot s}{(9.1 \times 10^{-31}\, kg)(0.4 \times 3.00 \times 10^{8}\, m/s)} = 6.1 \times 10^{-12}\, m$$

6.1×10^{-12} meters = 0.0061 nm = $^{1}/_{20}$ the diameter of the H atom

Heisenberg Uncertainty Principle

It is impossible to determine both the position and momentum of any particle simultaneously.

$$\Delta x \cdot \Delta(m\upsilon) > h$$

The uncertainty in the position of an electron,) *x*, times the uncertainty in the momentum,) *mv*, must be greater than Planck's constant, *h*.

In other words, since we can never know with certainty either the position or momentum of an electron, we can only discuss electrons in terms of probabilities. That is, we can only discuss the probability of finding an electron in some volume of space (energy levels, sublevels, and orbitals).

The Schrödinger Equation

$$\frac{\partial \Psi}{\partial x^2} + \frac{\partial \Psi}{\partial y^2} + \frac{\partial \Psi}{\partial z^2} + \frac{8\pi^2 m}{h^2}\left(E + \frac{e^2}{r}\right)\Psi = 0$$

In the Schrödinger equation, the square of the wave function, Q^2, (psi squared) gives the probability of finding an electron at point P in space around the nucleus.

To solve the Schrödinger equation for an electron in a three dimensional world, three integer numbers, n, R, and m_R must be introduced. These quantum numbers may have only certain combinations of numbers, as defined below.

<u>The Magnetic Properties of Matter</u>

Paramagnetic - Substances which are attracted to a magnetic field are said to be “**paramagnetic.**”

Paramagnetic materials have unpaired electrons.

Diamagnetic - Substances which are not magnetic and therefore cannot be attracted to a magnetic field are said to be **“diamagnetic”**

Diamagnetic substances have no unpaired electrons.

Diamagnetic substances are repelled slightly by a magnetic field.

The more unpaired electrons an atom has, the more strongly it is attracted to a magnetic field.

Which of the following substances are paramagnetic and which are diamagnetic?

V^{2+} B^{1+} Na^{1+} Ca^{0} O^{2-}

1s	2s	2p	2p	2p	3s	3p	3p	3p	3d	3d	3d	3d	3d	4s

Pauli Exclusion Principle

No two electrons may have the same four quantum numbers"

The Pauli exclusion principle limits occupancy of an orbital to two electrons with opposite spins.

Hund's Rule (of Maximum Multiplicity)

Whenever orbitals of equal energy are available, electrons are assigned to these orbitals singly before any pairing of electrons occurs.

Summary of the Quantum Numbers

Principal Quantum Number , n
Values = 1, 2, 3 . . . 4
Represents Orbital Size (distance from the nucleus) and Energy

Angular Momentum Quantum Number, R
Values = 0 . . . n-1
Represents Orbital Shape - circular, teardrop shaped, multiple teardrop shape, etc.
(subshell s, p, d, f)

Magnetic Quantum Number, m_R
Values = -R . . . 0 . . . +R
Represents Orbital Orientation - on which axis, the x, y, or z axis, the probability graph resides.
(number of orbitals in the subshell 1, 3, 5, 7)

Spin Quantum Number, m_s
Values = ±½
Represents the electron spin

Principal Quantum Number n	Angular Momentum Quantum Number R	Magnetic Quantum Number m_R
1	0 (1s)	0 (1 orientation)
2	0 (2s) 1 (2p)	0 (1 orientation) 1, 0, -1 (3 orientations)
3	0 (3s) 1 (3p) 2 (3d)	0 (1 orientation) 1, 0, -1 (3 orientation) 2, 1, 0, -1, -2 (5 orientations)
4	0 (4s) 1 (4p) 2 (4d) 3 (4f)	0 (1 orientation) 1, 0, -1 (3 orientations) 2, 1, 0, -1, -2 (5 orientations) 3, 2, 1, 0, -1, -2, -3 (7 orientations)

Give one of the two possible sets of quantum numbers for the valence electron in sodium, Na.

$_{11}Na\ 1s^2, 2s^2, 2p^6, 3s^1$

Valence electrons of sodium = $3s^1$

Therefore, n = 3

R = 0 (s type sublevel - spherical shell)

m_R = 0 (only 1 orbital or shape)

m_s = +½ or -½ (the arrow can be pointed either way)

Written: 3, 0, 0, +½ or 3, 0, 0, -½

Give all the possible sets of quantum numbers for the valence electrons in arsenic, As.

$_{33}As\ 1s^2, 2s^2, 2p^6, 3s^2, 3p^6, 4s^2, 3d^{10}, 4p^3$

Valence electrons $4s^2, 4p^3$

$_{33}As$	8 9	8	8	8
	4s	4p	4p	4p
m_R =	0	+1	0	-1

n = 4, R = 0 (s), m_R = 0, m_s = ±½

R = 1 (p), m_R = 1, 0, -1, m_s = ±½ (but must be the same)

4, 0, 0, +½, 4, 0, 0, -½ 4, 1, 1, +½, 4, 1, 0, +½, 4, 1, -1, +½ or

4, 0, 0, +½,   4, 0, 0, -½,   4, 1, 1, -½,   4, 1, 0, -½,   4, 1, -1, -½

Which of the following sets of quantum numbers are unacceptable?

(a) 1, 0, ½, -½

(b) 2, 2, 1, +½

(c) 3, 2, 1, 1

(d) 3, 0, 0, +½

(e) 4, 3, -2, +½

(a) Incorrect - no possible $m_\ell = \frac{1}{2}$

(b) Incorrect - ℓ from 0 to n-1, ℓ always $< n$

(c) Incorrect - $m_s = \pm\frac{1}{2}$ only

(d) Correct - 3s electrons

(e) Correct - 4f electron

Which sublevel, 1s, 2s, 2p, etc., has the set of quantum numbers

4, 2, -1, +½

What is the maximum number of orbitals which can be identified by each of the following sets of quantum numbers? If none, explain.

(a) $n = 3$, $R = 0$, $m_R = 0$

(c) $n = 5$, $R = 1$

(d) $n = 7$, $R = 5$

(d) $n = 4$, $R = 2$, $m_R = -2$