$$E = \frac{12400}{\lambda} = \frac{12400}{1400} = 8.857eV$$

The energy of the emitted photon of wavelength 1850Å

$$E = \frac{12400}{\lambda} = \frac{12400}{1850} = 6.702eV$$

Since the energy of the absorbed photon must equal to the total energy of the emitted photon

$$E_2 = E - E_1 = 8.857 - 6.702 = 2.155eV$$

$$\lambda_2 = \frac{12400}{E_2} = \frac{12400}{2.155} = 5754 \text{ Å}$$

<u>Ionization</u>

As most loosely bond-electron of an atom is given more and more energy, it moves into stationary states which are farther and farther away from the nucleus. The energy required to move the electron completely out of atom is called *ionization potential*.

Collisions of Electron with Atom

In order to excite or ionize an atom, energy must be supplied to it. This energy may be supplied to the atom in various ways, one of them is electron impact. Suppose that an electron is accelerated by the potential applied to a discharge tube. When this electron (has sufficient energy) collides with an atom, it may transfer enough of its energy to the atom to elevate it to one of the higher quantum state. If the energy of the electron at least equal to the ionization potential of the gas, it may deliver this energy to an electron of the atom and completely remove it from the parent atom. Three charged particle result from such ionizing collision; two electrons and a positive ion.

Example

(a) What is the minimum speed of an electron in order that a collision between it and a neon atom may result in ionization of this atom? (b) What is the minimum frequency that a photon can have to cause a photoionization of neon atom?

Note: The *ionization potential of neon is* 21.5eV

Solution

(a) The energy of bombared electron or photon must be at least equal to ionization potential

$$E_k = \frac{mv^2}{2}$$

$$\frac{mv^2}{2} = 21.5 \times 1.602 \times 10^{-19}\,J$$

$$v = \sqrt{\frac{2 \times 21.5 \times 1.602 \times 10^{-19}\,J}{9.109 \times 10^{-31}}} = 2.74 \times 10^{6}\,m/s.$$

(b)

$$\lambda = \frac{12400}{E} = \frac{12400}{21.5} = 576.74 \quad \text{Å}$$

$$f = \frac{c}{\lambda} = \frac{3 \times 10^{8}}{576.74 \times 10^{-10}} = 5.2 \times 10^{15}\,Hz$$

Collisions of Photon with Atoms

An Atom may absorb a photon of frequency f and thereby move from the level of energy E_1 to E_2 where

$$E_2 = E_1 + hf$$

An extremely important feature of excitation by photon capture is that the photon will not be absorbed unless its energy corresponds exactly to the energy difference between two stationary levels of the atom with which it collides.

If the frequency of the photon is sufficiently high, it may have enough energy to ionize the gas. The photon vanishes with the appearance of an electron and a positive ion. If the photon has more than ionizing energy of the excess will appear as the kinetic energy of the emitted electron.

The Dual Nature of Matter

In 1924 de Broglie postulated that all matters behave like a wave

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

Where p is the momentum of mass

Example

Calculate the de Broglie wavelength of

(a) a ball of mass 0.5 kg traveling at 0.016 m/s

(b) a dust particle of mass 1×10^{-4} kg traveling at 0.1 m/s

(c) an electron traveling at 5×10^{5} m/s

Solution

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

(a)

$$\lambda = \frac{6.626\times10^{-34}}{0.5\times0.016} = 8.28\times10^{-32}\,m$$ (b)

$$\lambda = \frac{6.626*10^{-34}}{10^{-4}*0.1} = 6.26*10^{-29}\,m$$

(c)

$$\lambda = \frac{6.626*10^{-34}}{9.109*10^{-31}*5*10^{5}} = 1.4*10^{-9}m$$

The kinetic energy is given by the formula

$$E_k = \frac{m.v^2}{2}$$

Also the kinetic energy is obtained from the accelerating voltage V

$$E_k = eV$$

$$\lambda = \frac{h}{(2meV)^{1/2}}$$

__Example__

Calculate mass contained in red light of wavelength 7×10^{-7}m which hits an electron at rest the energy contained in the light wave is completely lost to the electron. Calculate the velocity and the de Broglie wavelength of electron.

__Solution__

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

$$m = \frac{h}{\lambda v} = \frac{6.626\times10^{-34}}{7\times10^{-7}\times3\times10^{8}} = 3.14\times10^{-36}kg$$

The energy of light E=hf

$$E = \frac{hc}{\lambda} = \frac{6.626\times10^{-34}\times3\times10^{8}}{7\times10^{-7}} = 2.84\times10^{-19}J$$

$$v=\left(\frac{2E_k}{m}\right)^{1/2}=\left(\frac{2\times 2.84\times 10^{-19}}{9.109\times 10^{-31}}\right)^{1/2}=7.9\times 10^{10}\, m/s$$

$$\lambda=\frac{h}{mv}=\frac{6.626\times 10^{-34}}{9.109\times 10^{-31}\times 7.9\times 10^{10}}=9.25\times 10^{-25}\, m$$

$1Å=10^{-10}$ m
$\lambda= 9.25\times10^{-15}$ Å

Atoms with many electrons

The number of electrons in any one orbit is limited when these limit is reached a new orbit of greater radius is started. Each possible electron orbital is uniquely defined by a set of four quantum numbers.

Quantum Numbers: four numbers used to describe the electrons in an atom, are the *principal* (*n*), *angular* (*l*), *magnetic* (m_l), and *spin* (m_s) quantum numbers. These quantum numbers describe the size, shape, and orientation in space of the orbitals on an atom.

1. *Principal* (shell) quantum number - *n*

Describes the energy level within the atom.

- The quantum number *n* is integer.
- The principal quantum number (*n*) cannot be zero. *n* must be 1, 2, 3, etc. Also known as K, L ,M, etc
- Energy levels are 1 to 7
- The number of orbitals in a shell is n^2: $1^2 = 1$, $2^2 = 4$, $3^2 = 9$.
- Maximum number of electrons in n is $2n^2$

2. *Momentum* (subshell) quantum number - *l*

Describes the sublevel in n

- The quantum number *l* is integer.
- Each energy level has n sublevels. Sublevels of different energy levels may have overlapping energies.
- The angular quantum number (*l*) can be any integer between 0 and *n* - 1. If *n* = 3, for example, *l* can be 0, 1, or 2, the following lowercase letters are used to indicate different subshells.

s: $l = 0$
p: $l = 1$
d: $l = 2$
f: $l = 3$

- The number of orbitals in a subshell is therefore 2(*l*) + 1. There is one orbital in an *s* subshell (*l* = 0), three orbitals in a *p* subshell (*l* = 1), and five orbitals in a *d* subshell (*l* = 2).
- Maximum number of electrons in *l* is 2(2*l*+1)

Table 1.1 Principal and subshell quantum numbers

Principal Energy Level	Principal Quantum Number (*n*)	Maximum Number of Electrons in Each Sublevel *s* *p* *d* *f*	Total Number of Electrons ($2n^2$)
K	1	2	2
L	2	2 6	8
M	3	2 6 10	18
N	4	2 6 10 14	32

3. *Magnetic* quantum number – m_l

Describes the orbital within a sublevel

- The magnetic quantum number (*m*) can be any integer between -*l* and +*l*. If *l* = 2, *m* can be either -2, -1, 0, +1, or +2.

s: *l* = 0 *has 1 orbital*
p: *l* = 1 *has 3 orbitals*
d: *l* = 2 *has 5 orbitals*
f: *l* = 3 *has 7 orbitals*

4. Spin quantum number – m_s

This fourth quantum number describes the spin of the electron.

- Electrons in the same orbital must have opposite spins.
- Possible spins are clockwise or counterclockwise, spin quantum number m_s is arbitrarily assigned the numbers $+^1/_2$ and $-^1/_2$.

Pauli Exclusion Principle:

No two electrons in an atom have the same set of four quantum numbers

Electron configuration: **The electron configuration for chlorine is (Z=17)**

$$1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^5$$

- The **large numbers** represent the **energy level**.
- The **letters** represent the **sublevel**.
- The **superscripts** indicate the **number of electrons** in the sublevel

Example