

محاضرات التفريغ الكهربائي، المرحلة: ثالث سنة الفيزياء - د. وليد عيسى

«Electric Discharge»

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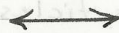
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$$p = \frac{1}{3} n m \bar{c}^2$$

where n is the number density or concentration of particles.
 m : the mass of each particle.
 \bar{c} : root mean square velocity irrespective of direction.

If this expression is compared with perfect gas law: $pV = RT$
 where V is the volume of one mole of gas.

1- Fundamental Processes

In order to understand the behaviour of any gas discharge it is necessary to know something of the detailed processes involving individual particles of the gas.

1-1 The kinetic Theory of a Simple Gas

The kinetic theory of gases relates the overall or macroscopic properties of a gas to the average behaviour of its particles.

By defining the collisions as perfectly elastic we mean that the energy exchanged in any collision is purely kinetic.

By considering the average force exerted on a wall by momentum transfer, it is readily shown that the pressure (P) of the gas can be expressed as:

$$P = \frac{nmC^2}{3} \quad (1)$$

Where n : is the number density or concentration of particles.

m : the mass of each particles.

C : root mean square velocity irrespective of direction.

If this expression is compared with perfect gas Law: $PV = RT$ (2)

Where V is the volume of one mole of gas.

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T ; is the absolute temperature (K)

R ; is the universal gas constant.

It follows that;

$$nmC^2V = 3RT$$

From which the mean energy of a particles is:

$$\frac{1}{2}mC^2 = \frac{3}{2} \frac{RT}{nV} \quad (3)$$

The $\left(\frac{R}{nV}\right)$ is written as (k) and known as Boltzmanns constant. It has the value 1.38×10^{-23} J/K, or 8.6×10^{-5} eV. We may now write

$$\text{Mean kinetic energy per particle} = \frac{3}{2}kT \quad (4)$$

Ex; At room temperature, the mean energy per particle is about 0.04 eV.

The foregoing relationships hold also for a mixture of such gases; partial pressures are then used.

* The mean free path (m.f.p) λ of the particles of a gas is the average distance which any one travels between collisions.

For the gas of identical particles and the particles have an actual diamet (d), any one will collide with another which lies with its centre within a cylindrical volume of diameter ($2d$), The volume per unit length of path, and if the number density (n)

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it contains $(n\pi d^2)$ particles, which is therefore the average number of collisions per particle per unit distance travelled. Hence the average distance between collisions;

$$\lambda = \frac{1}{n\pi d^2} \quad (4)$$

For the equilibrium distribution of velocities, the m.f.p is given by;

$$\lambda = \frac{1}{\sqrt{2} n\pi d^2} \quad (5)$$

For a mixture of gases, the mean free path of a particle of type 1 between collisions with any other type of component particle is;

$$\lambda_1 = \frac{1}{\pi \sum_r n_r d_{1r}^2 \sqrt{1 + \frac{m_1}{m_r}}} \quad (6)$$

Where $d_{1r} = \frac{1}{2}(d_1 + d_r)$, and m_r is the mass of an r-type particle, for $r=1$, eq. (6) reduces to (5).

Since the mass and diameter of an electron may be ignored when compared with another particle, eq. (6) gives the m.f.p of an electron as:

$$\lambda_e = \frac{4}{\pi n d^2} \quad (7)$$

The mean free time (τ) between collisions;

$$\tau = \frac{\lambda}{\bar{c}} \quad (8)$$

Where \bar{c} average velocity.

The collision frequency is

$$\nu = \frac{c}{\lambda} \quad (9)$$



1-2 The Atomic and Molecular Structure

1-2-1 The Bohr-Rutherford Atom;

The classical kinetic theory applies to a perfect gas of spherical particles making only elastic collisions. For atomic structure we need only introduce the Bohr-Rutherford model, which was originally advanced to explain atomic spectra in terms of early quantum theory.

In Rutherford model of the atom a nucleus of heavy particles, of number and net positive charge depending on the element considered, is surrounded by electrons rotating in concentric orbits. Bohr postulated that these orbits could only exist with certain radii. The radii he defined by quantum condition:-

$$mvr = \frac{nh}{2\pi} \quad (n=1,2,3,\dots) \quad (10)$$

Where (m) is the electronic mass.

(v) is the speed of an electron in an orbit of radius (r) and (h) is Plank's constant.

The angular momentum (mvr) is said to be quantized and there is an infinite number of possible orbit radii defined by the quantum number (n).

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Bohr also postulated that, if an electron falls from one orbit to another of small radius, the resulting difference in the atom's total energy is emitted in the form of a quantum of radiated energy, that the frequency (f) of the radiation is given by the relation:

$$hf = \Delta E \quad (11)$$

where (ΔE) is the energy difference.

If the element has atomic number (Z), the nucleus has by definition a positive charge (Ze), if (e) is the electronic charge. The electrostatic force attracting the orbiting electron is then ($\frac{Ze^2}{r^2}$) and this must be balanced by the centrifugal force ($\frac{mv^2}{r}$). We may therefore write for the kinetic energy of the electron:

$$T = \frac{1}{2}mv^2 = \frac{1}{2} \frac{Ze^2}{r} \quad (12)$$

The potential energy of the atom due to the attractive force on the electron is usually calculated by:

$$P = \int_{\infty}^r \frac{Ze^2}{r^2} dr = -\frac{Ze^2}{r} \quad (13)$$

We then have the total energy of the atom - if we ignore other possible means of energy storage - as the sum;

$$T+P = \frac{1}{2} \frac{Ze^2}{r} - \frac{Ze^2}{r} = -\frac{1}{2} \frac{Ze^2}{r} \quad (14)$$

The value of (r) for a given quantum number can be found by combining eq(10) and (12) to eliminate (v)

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We find that;

$$r = \frac{n^2 h^2}{4\pi^2 m e^2 Z} \quad (15)$$

The energy of the atom with its electron in the n th orbit then becomes;

$$E_n = - \frac{2\pi^2 m e^4 Z^2}{n^2 h^2} \quad (16)$$

and according to eq(11), this means that the frequency radiated by a transition between orbit (k) and orbit (j), ($k > j$) is;

$$f_{kj} = \frac{2\pi^2 m e^4 Z^2}{h^3} \left[\frac{1}{j^2} - \frac{1}{k^2} \right] \quad (17)$$

When the electron in ($n=1$) orbit the atom is said to be in the ground state. The highest value of the energy is zero occurs when (n) and therefore (r), is infinite. It corresponds to complete removal of the electron, or ionization.

An excited atom is one where the electron is in any intermediate orbit, with ($1 < n < \infty$). It is usually unstable with very short time (10^{-8} sec), but metastable state are exception to this and have relatively long lifetime (10^{-3} sec or more). The electron of an excited atom need not fall back directly into lowest level, but may do so by a series of successive transitions. For every value of (n) there is a series of frequencies caused by transitions of the electron into the (n th) orbit from all orbits of

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greater radius. The series tends to a limit frequency for $(n \rightarrow \infty)$ given by:-

$$f_{\text{lim}} = \frac{2\pi^2 m e^4 Z^2}{n^2 h^3} \quad (18)$$

Equations (17) and (18) are confirmed experimentally in the spectrum of hydrogen. Several series of lines can be observed, each having its own value. For example the Lyman series corresponds to $(n=1)$, the Balmer series to $(n=2)$. The Balmer series is the visible region $(\lambda = 626.2 \rightarrow 344.6 \text{ nm})$ and the Lyman in the ultraviolet $(\lambda = 121.6 \rightarrow 91.2 \text{ nm})$.

The foregoing theory applies to one-electron atoms only but this does not restrict it to hydrogen, for other atoms can have all electrons except one removed by ionization. For other atoms and for molecules the Bohr-Rutherford theory fails in respect of calculated energy levels, largely because the orbiting electrons as one would expect, influence each other.

The later and more powerful methods of quantum mechanics have adequately dealt with many-electron systems as well as single-electron atoms. But it is important to note eq.(11) remains valid for all systems. The energy in (eV) required for excitation or ionization is frequently quoted merely as a potential, in (volt), and in general called a critical potential. The potential required for excitation to the first excited state,

(11)

excluding metastable states, is called the resonance potential; its value varies from about 2 to 20V. The ionization potential required to remove one electron varies from 3.9V for caesium to 24.6V for helium.

1-2-2 Molecular Structure:-

The molecules, comprising two or more atoms, have a structure to which the quantum theory also applies is given by the specific heats of polyatomic gas. From eq. (3) it is clear that the total kinetic energy of the molecules in one mole of a classical monatomic gas is $(\frac{3}{2}RT)$. If we define the molar heat at constant volume, (C_v) , as the heat required to raise one mole through one absolute degree, then:

$$C_v = \frac{3}{2} R \quad (19)$$

Experimentally this relation is found to hold for monatomic gases. For other gases C_v has larger value. Thus polyatomic molecules must absorb energy other than by their random motion in which the kinetic energy is only that of translation. First, the molecule may spin about an axis through its centre of gravity normal to its own axis and thereby possess kinetic energy of rotation. Second the distance between the two atoms may vary periodically so that the molecule can possess vibrational energy which is part kinetic and part potential.

It follows that the total energy of a particle in any gas at uniform temperature is on the average $[(x/2)kT]$ where x is the number of quadratic terms into which total energy can be resolved; of this amount $(\frac{3}{2}kT)$ is due to translational motion.

Accordingly, the molar heat C_v should have the value $[(x/2)RT]$; for a diatomic molecule $x=7$.

This theory is in part supported by measured specific heats. For diatomic gases it is found that $[C_v = \frac{3}{2}RT]$ at low temperature ($\approx 50k$) increases to $[\frac{5}{2}RT]$ and after tends to the value $[\frac{7}{2}RT]$ as the temperature is increased ($\approx 500k$). The limit of vibrational excitation for molecule is dissociation, in the same way as ionization is the limit of electronic excitation.

