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| **Course Title:** | Basic Physical Chemistry I |
| **Course Code:** | **CHEM211** |
| **Credit Hours:** | **2.0** |
| **Requires:** | **111** |
| **Required for:** | **honours** |
| **Course Outline:** | What are the postulates of the kinetic model of matter and how do they apply to gases? What is the difference between a real and an ideal gas and what are the requirements for a gas to be “ideal”? What is the ideal gas equation and how is it derived? How can the ideal gas equation be used to derive Boyle’s, Charles’, Gay-Lussac’s, Avogadro’s and Dalton’s Laws? How can the ideal gas equation be used to derive expressions for molecular energies, molecular velocities, molar masses, densities and heat capacities of ideal gases? How can the ideal gas equation be used to molar masses of gases?What is the Maxwell-Boltzmann distribution of molecular energies and velocities? How can this distribution be used to deduce the probability of a molecule having a certain energy or speed? What are the different ways to measure average molecular speed?How can the collision frequency in a gas be calculated? What is the mean free path of a gas and how can it be calculated? What are transport properties of gases and how can the different transport properties of ideal gases be predicted?What causes a gas not to behave like an ideal gas and how does its behaviour change as a result? What are PV isotherms and what is the compressibility factor? What is the Joule-Thompson effect and under what conditions do gases liquefy? How can the real behaviour of gases be used to calculate molecular mass accurately? How can the behaviour of real gases be predicted using equations of state?How do molecules behave in electric fields and how can this behaviour be used to determine the dipole moment of a molecule? How do molecules behave in magnetic fields and how can this behaviour be used to determine the number of unpaired electrons in a molecule? |
| **Lesson 1:** | Kinetic Theory of Matter and the Ideal Gas Equation |
| **Lesson 2:** | Derivation of the Gas Laws from first principles |
| **Lesson 3:** | Partial Pressures and Molar Heat Capacities |
| **Lesson 4:** | Maxwell-Boltzmann distribution of molecular energies and velocities |
| **Lesson 5:** | Collision Frequency and Mean Free Path |
| **Lesson 6:** | Transport Properties of Gases |
| **Lesson 7:** | The Behaviour of Real Gases |
| **Lesson 8:** | Equations of State for Real Gases |
| **Lesson 9:** | The effect of electric fields on molecules |
| **Lesson 10:** | The effect of magnetic fields on molecules |

# *items in italics are covered at senior secondary level*

# Lesson 1

# OVERVIEW OF THE KINETIC MODEL OF MATTER

* The kinetic theory of matter proposes that gases behave as follows:
* gas molecules are in constant motion in random directions
* they frequently collide with one another and these collisions are elastic, but energy can be transferred from one molecule to another as a result of these collisions
* the total energy of the particles in a closed system remains constant at a given temperature
* the average kinetic energy of the particles is directly proportional to the temperature
1. **REAL AND IDEAL GASES**
* An ideal gas is a gas in which:
* Gas particles are hard spheres which experience no attractive forces on each other
* The volume occupied by the gases is negligible compared with the volume of the container
* The behaviour of ideal gases can be predicted using simple principles of Physics
* A real gas is a gas which does not model the expected behaviour of an ideal gas, because either the intermolecular forces are significant (large molecules at high pressures and low temperatures) or because the volume of the molecules is a significant fraction of the total space available (very high pressures) or because a significant amount of energy can be absorbed by the molecule without increasing its speed (complex molecules).
* Gases are therefore most likely to demonstrate ideal behaviour if the molecules are small and simple, the pressures are low and the temperatures are relatively high.
* The behaviour of real gases is much more difficult to model than the behaviour of ideal gases.
1. **THE IDEAL GAS EQUATION**
* An equation of state is an equation which describes how the physical state of a material depends on a particular physical condition or range of physical conditions. Equations of state are used to describe the physical properties of solids, liquids and gases.
* Gases differ from solids and liquids in that many cases their behaviour does not depend on the identity of the gas – in other words, all gases behave in very similar ways and so the same equations of state can be used to describe many gases
* The ideal gas equation is an equation of state which relates the pressure, volume, temperature and amount of substance for an ideal gas. It is made by combining three of the ideal gas laws:

 Boyle’s Law: PV = k1 Charles’ Law: = k2 Avogadro’s Law: = k3

 Combining all three laws you get: = R or **PV = nRT** This is the ideal gas equation

 P = pressure in Pa, V = volume in m3, n = number of moles, T = temperature in K

 R = molar gas constant (8.31 Jmol-1K-1) and represents the value of for one mole of a gas

* The ideal gas equation can also be expressed in terms of individual molecules, by replacing the molar gas constant R with the Boltzmann constant k (1.38 x 10-23 JK-1) to give **PV = nkT**

 where L is Avogadro’s number (6.02 X 1023 mol-1)

* The ideal gas equation can be used to calculate P, V, n or T of a gas if the other three quantities are known.
* If the mass can be measured, the ideal gas equation can also be used to calculate the molar mass (mr) of a gas: n = and n = so **mr =**

The **Victor-Meyer** apparatus is designed to measure the molar mass of a gas as follows:

* A small glass bottle containing a known mass of the volatile substance under investigation is dropped into the inner part of a Victor Meyer tube, which has a bulb at the lower end and a side-arm at the upper end which leads to a trough filled with water; a liquid is placed in the outer part of the Victor Meyer tube which must have a boiling point at least 30 K higher than the substance under investigation; the outer liquid is heated until the liquid under investigation boils; when it does so it will displace a fixed volume of air, the volume of which can be measured; the ideal gas equation can be used to deduce the moles of air displaced, which will be equal to the number of moles of vapour produced by the substance

The **Dumas** method involves the use of a retort shaped bulb of known volume, with a very thin side arm

* The bulb is weighed, the mass of air inside is deduced from the volume of the bulb and the known density of air at that temperature; the mass of the bulb alone can therefore be calculated; a small quantity of the desired liquid is added to the bulb, which is then immersed in another liquid (with a higher boiling point) and heated until the liquid in the bulb has completely boiled; the tip of the bulb is then sealed and the bulb is allowed to cool to room temperature; the bulb is then weighed and the mass of the substance in the bulb can therefore be calculated; the pressure at the time of the experiment was atmospheric pressure and the temperature at the time of the experiment was the temperature of the liquid being heated
* Also, given that density (ρ) = , the molar mass of a gas an also be expressed as a function of its density: **mr =**

Note that mass, molar mass and density must all be in SI units (kg, kgmol-1 and kgm-3 respectively)

Lesson 2

1. **DERIVING THE IDEAL GAS EQUATION FROM FIRST PRINICPLES**
* Consider a cube of length l; each inner face of the cube will have an area l2 so the total area (A) of the inner faces/walls of the cube will be 6l2 and the volume (V) of the cube will be l3; this cube contains N molecules, each of mass m, moving with velocity ux in the x direction
* The momentum of each molecule before a collision with a wall is mux. The momentum of each molecule after the collision is -mux, so the change in momentum during the collision is 2mux.
* The distance between the walls is l, so the time between collisions is , so there are collisions per second
* So the total change in momentum per second (ie the force exerted F) per molecule with the x walls is:

 change in momentum per collision x number of collisions per second = 2mux x = F = 2

* This is also happening in the y and z directions, so the total force exerted on the walls of the container by each molecule is:

2 + 2 + 2 = (ux2 + uy2 + uz2) = 2

* So the total force exerted by all the molecules in the container = 2
* So the pressure in the container P = = = = = , so PV =
* Molar mass mr = mL and n = N/L, then mN = mrn so PV = (equation 1)
* The kinetic theory of matter states that the kinetic energy E of a gas is directly proportional to the temperature, so E = k1T
* The kinetic energy of one molecule can be given as , so  **=** k4T
* The total kinetic energy of all the molecules is therefore = = Nk4T = nLk4T (equation 2)
* Combining equations 1 and 2 gives the ideal gas equation:
* PV = so = = nLk4T so PV =
* If is expressed as a single constant R, then **PV = nRT**
* If k4 = then the kinetic energy for one molecule of a gas can be written **E =** or and the kinetic energy for one mole of a gas can be written E = ; E per molecule can also be expressed as , so E per mole is , so 3RT = mrv2 and **v =**
* The equations PV = and = k4T can be used to derive the three gas laws individually:
* PV = nRT so PV is constant for a given number of moles of gas at constant temperature (Boyle’s Law)
* = so is constant for a given number of moles of gas at constant pressure (Charles’ Law)
* = so is constant at constant temperature and pressure (Avogadro’s Law)

**Lesson 3**

1. **HEAT CAPACITY OF IDEAL GASES**
* The molar heat capacity of a gas is the energy required to heat one mole of the gas by 1 K. It can be measured either at constant volume (Cv) or at constant pressure (Cp).
* More energy is needed to heat a gas by 1 K at constant pressure (when the volume increases) than at constant volume (when the pressure increases). This is because extra energy is needed to increase the volume against an external pressure.
* At constant volume, E = so Cv = =
* At constant pressure, the same energy is needed to heat the gas by 1 K, but energy is also needed to expand the gas; this energy ΔE = PΔV; according to the ideal gas equation, V = so ΔV = , so ΔE = nRΔT

 So = nR and so the energy required per mole for the expansion = R

 So Cp =  **+** R **=**

* So **Cv =** and **Cp =**
* The heat capacity at constant volume and pressure for individual molecules can be written as and **Cp =** respectively

#### DALTON’S LAW OF PARTIAL PRESSURES

The ideal gas Law can be used to derive Dalton’s Law of Partial Pressures:

The pressure exerted by gas A in a container can be given as: pA =

The pressure exerted by gas A in a container can be given as: pB =

The total pressure exerted by both gases is therefore pA + pB = + = = = P

So the total pressure in a mixture of gases is the sum of the partial pressures of individual gases.

 = / = so pA = P where is the mole fraction of A in the mixture

#### Lesson 4

#### MAXWELL-BOLTZMANN DISTRIBUTION OF MOLECULAR ENERGIES AND VELOCITIES

* During collisions, energy is transferred from one molecule to another and as a result the speed of both molecules will change. As a result, not all of the molecules will have the same kinetic energy at a given time, but there will be a distribution of molecular energies and velocities.
* Maxwell and Boltzmann derived the probability distribution function for molecular velocities (v) in a single direction:

 p =

* This can be converted into a distribution function in three dimensions for molecular speeds (c):

 p =

* Using E = ½mc2, this can also be converted into a distribution function for molecular energies:

 p =

* The first part of this distribution is the Maxwell part and it shows p increasing with √E; this happens because the more energy there is, the greater the number of ways to distribute the energy amongst the molecules and so the more probable the energy (the entropy factor); the second part of this distribution is the Boltzmann part and shows p decreasing rapidly with increasing E; the number of molecules with energy much higher than the mean will be small and will decrease with increasing energy
* To calculate the probability of a molecule having a speed or energy within a particular range, it is necessary to integrate the distribution function within those two limits; however this is difficult because the function takes the form y = which is not easily integrated; an approximation can be made at high energies by assuming that the Maxwell part of the equation is approximately constant at high energies; the resulting integral then simplifies to = (; note that this approximately is only valid when the range of energies under consideration is significantly greater than the average energy
* Increasing the temperature and decreasing the value of mr increases the proportion of molecules with higher speeds; the Maxwell-Boltzmann distribution of molecular energies is best demonstrated graphically (the plots of probability against speed and probability against energy are very similar)

  

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1. **DIFFERENT MEASURES OF MOLECULAR SPEED**
* We know that the total kinetic energy of one mole of molecules = ½mrv2 =

 Therefore v2 =

 Therefore **v (or the root mean square speed) =**

* Analysis of the Maxwell-Boltzmann distribution gives the **most probable speed** of the molecules as
* Further analysis of the Maxwell-Boltzmann distribution gives the **average speed** of the molecules as
* The values of root mean square speed, average speed and most probable speed are therefore not the same value. Root mean square speed is important for kinetic energy considerations, but average speed is generally more useful in predicting individual particle behaviour

 vrms > vav > vmp in the ratio √3:: √2

  [www.hyperphysics.phy-astr.gsu.edu](http://www.hyperphysics.phy-astr.gsu.edu)

# Lesson 5

### MEAN FREE PATH AND COLLISION FREQUENCY

* The kinetic model of matter can also be used to predict two important quantities in gas kinetics – the collision frequency and the mean free path
* The effective collision area of a molecule with another of the same type can be given by A = πd2

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* If the velocity of the particles is v, the relative velocity of the two particles is v√2 (this is the relative velocity when two particles collide at right angles)
* In one second, the volume swept by this collision area will be πd2 x relative velocity = πd2v√2
* The number of collisions in this time depends on the number of other molecules in this volume of space N so the number of times a molecule will collide in this volume of space per second = Nπd2v√2

There are two useful consequences of this:

* The total number of colliding molecules in this space per second = N x Nπd2v√2 = N2πd2v√2

 so number of collisions per second, **or collision frequency Z** = =

By substituting in the expression for v = you get Z = =

* If there are two different colliding particles A and B, the total number of times a particle A will collide with particle B will be NBπ(rA + rB)2v√2so the total number of collisions between A and B will be

πv(rA + rB)2NANBv√2

The mr in the velocity equation needs to be changed to 2μ, where μ is the reduced mass

so v = so ZAB = (rA + rB)2NANB

* The total number of colliding molecules in this space per second = N x Nπd2v√2 = N2πd2v√2
* so the number of colliding molecules per unit volume per second =
* so the time interval between individual molecules colliding =
* so the distance travelled in between collisions, or the **mean free path λ**, (v x t) = , = = so **λ =**

**Lesson 6**

1. **TRANSPORT PROPERTIES OF GASES**

**Transport properties** are the properties of a compound or material associated with the transport of a particular quantity through it. Three common examples of transport properties are:

* diffusion or effusion: the transport of mass
* viscosity: the transfer of momentum
* thermal conductivity: the transfer of energy

The kinetic model of matter can be used to predict a number of transport properties of gases, including diffusion, thermal conductivity and viscosity.

1. **Diffusion and Effusion**
* Diffusion is the mixing of gases until they reach a uniform composition. Effusion is the escape of gases through a small hole into a vacuum. The rates of diffusion and effusion both follow Graham’s Law.
* The simple idea that all gases at the same temperature have the same average kinetic energy can be used to prove Graham’s Law of diffusion (or effusion):

 At a given temperature, m1v12 = m2v22, so = and =

 Since masses are directly proportional to density, this relationship can also be expressed as:

 So the relative rates of diffusion of two gases v1 and v2 is inversely proportional to the square roots of their molar masses or densities – this is **Graham’s Law**.

* The absolute rate of diffusion of a gas (fn) in molm-2s-1 is proportional to the concentration gradient :

 fn = D (this is known as Fick’s first Law)

D is the **coefficient of diffusion** (m2s-1)

* D is related to the velocity of the particles and their mean free path: D =

It can therefore be shown that D is proportional to , P-1, d-2 and

1. **Viscosity**
* Viscosity is the resistance offered by one part of a gas to the movement of another part of the gas. It results from the transfer of momentum from one part of the gas to another and is not a result of intermolecular attraction. It is therefore a feature of all gases.
* The rate of transfer of momentum in a gas (fp) in kgm-s-2 is proportional to the velocity gradient :

 fp = η

η is the **viscosity** (kgm-1s-1)

* η is related to the velocity of the particles and their mean free path: η =

It can therefore be shown that η is proportional to , d-2 and

1. **Thermal Conductivity**
* Energy flows from hot bodies to cold bodies until the temperature (or kinetic energy) of the molecules is equalised.
* The rate of heat flow fE in a gas (Jm2s-1) is proportional to the temperature gradient

 fE = κ

κ is the **thermal conductivity** (JK-1m-1s-1)

* κ is related to the velocity of the particles and their mean free path: κ =

It can therefore be shown that κ is proportional to , d-2 and

1. **Summary of transport properties of gases**

|  |  |  |
| --- | --- | --- |
| Transport Property | Rate of Transfer | Constant of Proportionality |
| Diffusion | fn = D | D =  |
| Viscosity | fp = η | η =  |
| Thermal Conductivity | fE = κ | κ =  |

### Lesson 7

### THE BEHAVIOUR OF REAL GASES

1. **Observing deviations from ideal gas behaviour**
2. **Compressibility factor**

Most gases only obey the ideal gas laws at a limited range of pressures and temperatures – generally low pressure and high temperature.

In ideal gases, = 1

If ideal gas behaviour is not observed, however, then is not equal to 1 and is given a value known as the compressibility factor (z).

At low temperatures and reasonably high pressures, the attractive forces between the particles need to be considered – these forces cause the V to be lower than expected for a particular P and cause P to be lower than expected for a particular V. These factors cause PV to be lower than expected (ie PV < nRT and z < 1); at low temperatures and relatively high pressures, z < 1.

At very high pressures, the size of the particles themselves becomes a consideration; the larger the particles, the less space there is in which for them to move. When this factor is significant, P is larger than expected for a particular V and V is larger than expected for a particular P. These factors cause PV to be higher than expected (ie PV > nRT and z > 1). At high temperatures and very high pressures, z > 1.

At low pressures, z = 1 and gases display ideal behaviour.

The stronger the intermolecular forces between gas molecules, and the larger the molecular size, the less likely a gas is to show ideal behaviour.

Non-ideal behaviour in gases can be illustrated in plots of compressibility factors against pressure (these plots are also called isotherms).



 <http://www.askiitians.com> [www.opentextbc.ca/introductorychemistry](http://www.opentextbc.ca/introductorychemistry)

1. **The Joule-Thomson effect**
* The Joule-Thomson effect describes how the temperature of a gas changes if a gas is expanded under conditions in which its total energy cannot change (ie an adiabatic process). It can be described in the form of a coefficient μ, in which μ = . Μ is positive if a decrease in pressure causes a decrease in temperature, and negative if a decrease in temperature causes an increase in temperature.
* The Joule-Thomson effect results from attractive and repulsive forces between molecules. Ideal gases do not display the Joule-Thomson effect and μ = 0.
* If the attractive forces between molecules are more significant than the repulsive forces, then work is required to expand the gas, so kinetic energy will be converted into potential energy and the gas will cool down. This is the case for most typical gases at room temperature (μ is positive).
* If the repulsive forces between molecules are more significant than the attractive ones, then the expansion of the gas will reduce these repulsions, potential energy will be converted into kinetic energy and the gas will heat up. This is typical at very high pressures and temperatures (μ is negative).
1. **Liquefaction**
* Below a certain temperature, the intermolecular forces will be strong enough to cause the molecules to stick together, meaning that the gas will begin to condense into a liquid. This will happen if the temperature is low enough for the kinetic energy of the molecules is insufficient to overcome the potential energy of attraction.
* The temperature above which a gas cannot be turned into a liquid, however high the pressure, is called the **critical temperature Tc**. The minimum pressure required to liquify a gas at the critical temperature is called the critical pressure **Pc**. The volume occupied by one mole of a gas at the critical temperature and pressure is called the critical volume **Vc**.
* Above the critical temperature, a decrease in volume will always result in an increase in pressure. Below the critical temperature, a decrease in volume will result in an increase in pressure until the pressure is sufficient to cause the gas to liquefy, at which point a decrease in volume may not cause an increase in pressure. At the critical temperature and the critical pressure, a small change in pressure will not affect the volume.



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describe results obtained when trying to verify Boyle’s Law, experiments which led to the liquefaction of CO2

**Lesson 8**

1. **Applications of real gas behaviour**
2. **Accurate determination of molecular mass**

Given that mr = , it follows that = and that the value of should be a constant at a particular temperature. However this is only true at P = 0, when all gases behave ideally. At pressures close to atmospheric pressure, z < 1 and the value of is higher than expected.

However the value of increases linearly with increasing P at relatively low pressures, which means that a plot of against P is a straight line. This line can be extrapolated back to P = 0; where the intersect with the axis will be equal to and the molecular mass can be determined very accurately.

1. **Modelling the behaviour of real gases:**

There have been various attempts to make adjustments to the ideal gas equation to make it more applicable to real gases:

1. **The Van der Waal’s Equation**
* Van der Waal proposed that the observed pressure of a real gas would be lowered by intermolecular attractions. These would be observable when particles approached each other and the likelihood of this happening depends on the collision frequency , so the reduction in pressure would be equal to , where a is a constant

So **P(ideal) = P(real) +**

* Van der Waal also proposed that the actual volume of space in a container should be reduced by the volume occupied by the particles themselves, so actual the volume of space = V – nb, where n is a constant
* Therefore Van der Waal proposed that **(P + )(V – nb) = nRT**; this is the Van der Waal’s equation
* The stronger the intermolecular forces, the larger the value of a; and in general, the larger the molecular size, the larger the value of b

|  |  |  |
| --- | --- | --- |
| Gas | a (Pam6mol-2) | b (m3mol-1) |
| Hydrogen | 0.0247 | 26.7 x 106 |
| Carbon Dioxide | 0.364 | 42.7 x 106 |
| Helium | 0.034 | 23.7 x 106 |
| Ammonia | 4.17 | 37.1 x 106 |
| Water Vapour | 5.46 | 30.5 x 106 |

* Van Der Waal also showed that when the volume, temperature and pressure of a gas a measured relative to their critical values, they all show very similar deviations from ideal behaviour.

If reduced pressure = , reduced temperature = and reduced volume = , then all gases measured at the same reduced temperature, volume and pressure will have the same compressibility factor. This is known as the Law of Corresponding States.

* As a result Van der Waal was able to predict the critical temperature and pressure of different gases by measuring a and b. As a result, if the critical pressure and temperature are known, the Van Der Waal;s constants a and b can be calculated.

It can be shown that a = and b =

1. **The Dieterici Equation**

The Dieterici equation is very similar to the Van der Waal’s equation, but instead of using an arithmetic pressure adjustment () to estimate the effect of intermolecular attraction, it uses an exponential function as follows:

* at high temperatures and volumes, = 1 and ideal gas behaviour is observed once the molecular size adjustment (V – b) is accounted for
* The value is significantly less than one at low temperatures, showing a significant intermolecular attraction
1. **The Kamerlingh-Onnes (or Virial) Equation**

The virial equation expressed PV as a power series of the pressure at a given temperature:

 = A + BP + CP2 + DP3 +….

In an ideal gas, A = RT and B, C and D are all zero. The coefficients A, B, C and D are known as virial coefficients, which are constant at a particular temperature.

A is always positive and is equal to RT

At low temperatures, B is negative (z < 1) but at high temperatures, B is positive (z > 1)

The remaining coefficients are only significant at very high pressures.

**SECTION 2 – BEHAVIOUR OF MOLECULES IN ELECTRIC AND MAGNETIC FIELDS**

**Lesson 9**

### Electric Fields and Dielectric constants

* The force of attraction between two charges in a vacuum is given by:

 F =

 q1 and q2 = charges, r = separation, εo permittivity of free space (resistance to creating an electric field)

* In a different medium, the resistance of the medium to creating an electric field must be considered:

 F = (εr = dielectric constant of the medium)

 This can be converted into potential energy as follows: E =

 The greater the dielectric constant, the smaller the electric field strength per unit of charge

* Water has a very high dielectric constant (80) and so the attraction between oppositely charged ions in water is much lower than in a vacuum, which is why water is able to break up many lattices
* The capacitance of capacitors is directly linked to the dielectric constant of the medium between the plates and this is the most convenient way to measure dielectric constants: εr =
* Dielectric constants are useful for deducing a number of useful properties of the substance:
* Permanent dipole moments
* Magnetic moments
1. **Dipole Moments of Molecules**
* A dipole moment is a separation of electrical charge resulting from the centre of positive charge (z) and the centre of negative charge (z) not being the same place, but separated by a distance l

The dipole moment on a molecule μ = zl

* All molecules can have a dipole moment induced by an electric field (induced polarization), as the molecules are distorted by the electric field, with the electrons moving closer to the positive plate and the nuclei moving closer to the negative plate. The degree of polarization depends on the electric field strength X and the polarizability of the molecule α: μi = αX
* A more useful term is the induced molar polarization of the molecule Pi = which is the distortion produced in one mole of a substance when the field strength is 1 Vm-1
* The induced molar polarization of the molecule is a quantity independent of temperature
* Some molecules have a permanent dipole moment (μ). In the presence of an electric field, these molecules will align themselves to the electric field (molar orientation polarization). This process is more difficult at high temperatures. In addition, there will be an additional dipole moment due to distortion of the nuclei and electrons by the electric field. The total molar polarization will be equal to the sum of the induced molar molarization and the molar orientation polarization: Pt = Pi + Po
* Po is linked to the permanent dipole moment of the molecule by the Debye equation: Po =
* The total molar polarization can this be written:

Pt = +

This can written in the form: Pt = A + where A = and B =

A graph of Pt against should therefore give a straight line of gradient B

From the gradient: μ =

* Pt is also related to the dielectric constant of the molecule by the Clausium-Mosotti equation

 PT =

and so can be determined by measuring the dielectric constant and the density at different temperatures

* The permanent dipole moment of non-volatile substances cannot be measured directly but can be measured by preparing solutions of them in non-polar solvents:
* Pt = nAP0A + nBP0B; if the mole fraction and P0A is known, P0B can be calculated
* the variation of P0B with temperature can give μB.

### Lesson 10

### Magnetic Properties of Molecules

* In a vacuum the attraction between two magnetic moles is given by F = ; in any other medium, this is adapted to F = ; q is the magnetic permeability of the medium
* Diamagnetic substances have no unpaired electrons and so the individual electrons cannot align themselves to the magnetic field; a magnetic field is induced in the electrons which opposes the magnetic field and q < 1 for diamagnetic substances
* Paramagnetic substances have unpaired electrons; these can align themselves to the magnetic field which enhances the magnetic field strength, so q > 1 for paramagmetic substances and q is large for ferromagnetic substances
* More useful quantities are the specific magnetic susceptibility χ = and the molar magnetic susceptibility χm = ; χ is positive for paramagnetic and ferromagnetic materials and negative for diamagnetic materials
* Specific magnetic susceptibility can be measured by Gouy’s method:
* A cylinder containing the sample is suspended by a wire from one arm of a balance until its lower part is centred between the poles of an electromagnet
* The mass of the cylinder is counterbalanced by weights in the other
* The mass needed to restore balance after the electromagnet is switched on is measured
* Δmg =
* For a substance S of known magnetic susceptibility: Δmsg =

 So =

* The total magnetic susceptibility of a substance is the sum of its diamagnetic susceptibility and its paramagnetic susceptibility; the diamagnetic component is temperature independent but the paramagnetic component decreases with increasing temperature:
* Langevin’s equation: χm = αm + ; μm = inherent magnetic moment
* A graph of χm againt give you a gradient of
* μm can be used to find the number of unpaired electrons in a molecule: μm =

 S = 0.5 x number of unpaired electrons in molecule