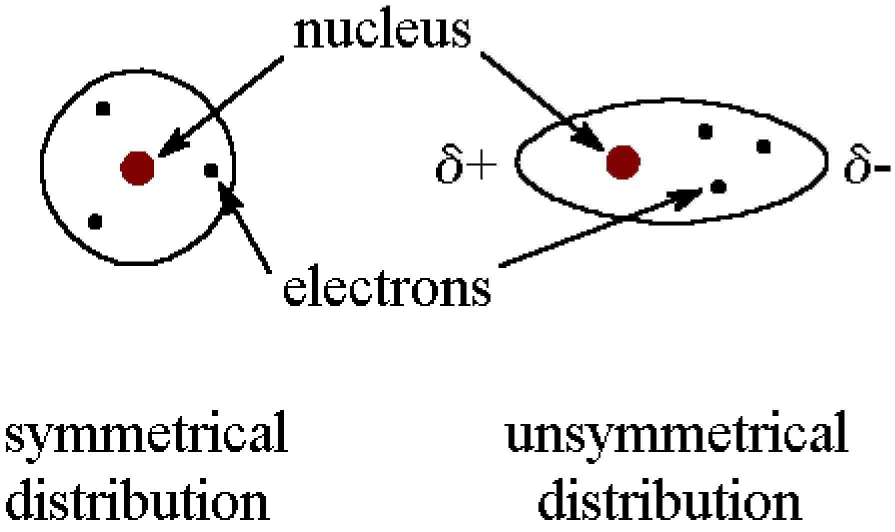
States of matter

- Intermolecular forces are the forces of attraction and repulsion between interacting particles.
- Attractive intermolecular forces are known as van der Waals forces. it is important to note that attractive forces between an ion and a dipole are known as ion-dipole forces and these are not Van der Waals forces.

- **London forces**

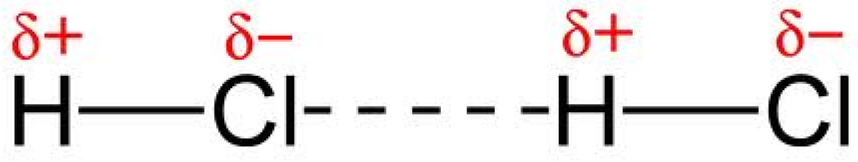
  The London dispersion force is the weakest intermolecular force. It is a temporary attractive force that results when the electrons in two adjacent atoms occupy positions that make the atoms form temporary dipoles. This force is sometimes called an induced dipole-induced dipole attraction. London forces are the attractive forces that cause non-polar substances to condense to liquids and to freeze into solids when the temperature is lowered sufficiently.



- These forces are always attractive and interaction energy is inversely proportional to the sixth power of the distance between two interacting particles (i.e., $1/r^6$ where r is the distance between two particles).

Dipole-dipole forces

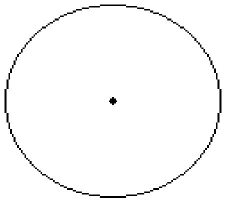
- Dipole-dipole forces are attractive forces between the positive end of one polar molecule and the negative end of another polar molecule. Dipole-dipole forces have strengths that range from 5 kJ to 20 kJ per mole. They are much weaker than ionic or covalent bonds and have a significant effect only when the molecules involved are close together (touching or almost touching).



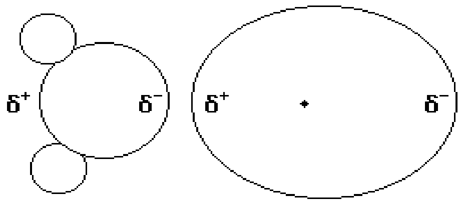
- Polar molecules have a partial negative end and a partial positive end.
- The partially positive end of a polar molecule is attracted to the partially negative end of another

Dipole Induced Dipole Forces

- A dipole-induced dipole attraction is a weak attraction that results when a polar molecule induces a dipole in an atom or in a non-polar molecule by disturbing the arrangement of electrons in the non-polar species.



Spherical atom with no dipole.
The dot indicates the location
of the nucleus.



Upon approach of a molecule with a
dipole, electrons in the atom respond
and the atom develops a dipole.

Hydrogen bonding

- The hydrogen bond is really a special case of dipole forces. A hydrogen bond is the attractive force between the hydrogen attached to an electronegative atom of one molecule and an electronegative atom of a different molecule. Usually the electronegative atom is oxygen, nitrogen, or fluorine, which has a partial negative charge. The hydrogen then has the partial positive charge. Hydrogen bonding is usually stronger than normal dipole forces between molecules.

$$\overset{\delta+}{H}-\overset{\delta-}{F}\cdots\overset{\delta+}{H}-\overset{\delta-}{F}$$

Boyle's Law

- At constant temperature, the pressure of a fixed amount (i.e., number of moles *n*) of gas varies inversely with its volume. This isknown as Boyle's law.

  $pV = K$

  $p_1V_1 = p_2V_2$

  $p_1 / p_2 = V_1/V_2$

  p- Pressure, V-volume, K-constant.

- At a constant temperature, pressure is directly proportional to the density of a fixed mass of the gas.

Charles Law

- Charles' Law describes the direct relationship of temperature and volume of a gas. Assuming that pressure does not change, a doubling in absolute temperature of a gas causes a doubling of the volume of that gas. A drop of absolute temperature sees a proportional drop in volume. The volume of a gas increases by 1/273 of its volume at 0°C for every degree Celsius that the temperature rises.

Temperature = Constant x Volume

or

Volume = Constant x Temperature

or

Volume/Temperature = Constant

- Mathematically,

$V_1/T_1 = V_2/T_2$

Gay Lussac's Law

- At constant volume, pressure of a fixed amount of a gas varies directly with the temperature.
- Mathematically,

  $P/T = \text{constant}$

Avogadro's Law

- It states that equal volumes of all gases under the same conditions of temperature and pressure contain equal number of molecules.
- Mathematically,

  $V = k*n$

  $k = \text{Avogadro number} = 6.023*10^{23}$

Ideal Gas Equation

- A gas that follows Boyle's law, Charles' law and Avogadro law strictly is called anideal gas
- Mathematically,

  $pV = n RT$.
- R is called gas constant. It is same for all gases. Therefore it is also called UniversalGas Constant and its value is = 8.314 J $K^{-1} mol^{-1}$.

Combined Gas Law

- $P_1 V_1/T_1 = P_2 V_2/T_2$

Density and Molar Mass of a Gaseous Substance

- $M = dRT / p$ (d=density)

Dalton's Law of Partial Pressures

- The total pressure exerted by the mixture of non-reactive gases is equal to the sum of the partial pressures of individual gases.
- p_{Total} = p1+p2+p3+......(at constant T, V)
- Pressure exerted by saturated water vapour is called **aqueous tension**. Aqueous tension of water at different temperatures.
- $p_{Dry\ gas} = p_{Total}$ – Aqueous tension

Partial pressure in terms of mole fraction

- $p_i = x_i\ p_{total}$

  where x_i is mole fraction.

BEHAVIOUR OF REAL GASES: DEVIATION FROM IDEAL GAS BEHAVIOUR

- Due to the failure of the following two assumptions of the Kinetic gas theory the deviation is observed.
    - There is no force of attraction between the molecules of a gas.
    - Volume of the molecules of a gas is negligibly small in comparison to the space occupied by the gas.
- The deviation from ideal behaviour can be measured in terms of compressibility factor Z, which is the ratio of product pV and nRT.
- At high pressure all the gases have $Z > 1$. These are more difficult to compress. At intermediate pressures, most gases have $Z < 1$. Thus gases show ideal behaviour when the volume occupied is large so that the volume of the molecules can be neglected in comparison to it.
- The temperature at which a real gas obeys ideal gas law over an appreciable range of pressure is called Boyle temperature or Boyle point. Boyle point of a gas depends upon its nature. Above their Boyle point, real gases show positive deviations from ideality and Z values are greater than one. The forces of attraction between the molecules are very feeble. Below Boyle temperature real gases first show decrease in Z value with increasing pressure, which reaches a minimum value.

Sample Examples

- A balloon is filled with hydrogen at roomtemperature. It will burst if pressureexceeds 0.2 bar. If at 1 bar pressure thegas occupies 2.27 L volume, upto whatvolume can the balloon be expanded ?

  Solution

  According to Boyle's Law $p_1V_1 = p_2V_2$

  If p_1 is 1 bar, V_1 will be 2.27 L

  If $p_2 = 0.2$ bar, then

  $V_2 = p_1V_1/ p_2 = 1*2.27/0.2$

  $\Rightarrow V = =11.35$ L

  Since balloon bursts at 0.2 bar pressure,the volume of balloon should be less than11.35 L.

- At 25°C and 760 mm of Hg pressure agas occupies 600 mL volume. What willbe its pressure at a height wheretemperature is 10°C and volume of thegas is 640 mL.

  Solution

  P_1 = 760 mm Hg, V_1= 600 mL

  T_1 = 25 + 273 = 298 K

  V_2 = 640 mL and T_2 = 10 + 273 = 283 K

  According to Combined gas law, $P_1 V_1/T_1 = P_2V_2/T_2$

  Substituting the values of P_1, V_1, T_1,V_2, T_2 in the above equation,

  P_2 = 676.6 mm Hg

- On a ship sailing in pacific ocean wheretemperature is 23.4 °C , a balloon is filledwith 2 L air. What will be the volume ofthe balloon when the ship reaches Indianocean, where temperature is 26.1°C ?

  Solution

  V_1 = 2 L

  T_2 = (26.1 + 273) K = 299.1 K

  T_1 = (23.4 + 273K) = 296.4 K

  From Charles law, $V_1/T_1 = V_2/T_2$

  Substituting the values of V_1, T_1 and T_2 in the above equation, we get

  V_2=2.018L.