CHM 101

GENERAL CHEMISTRY I

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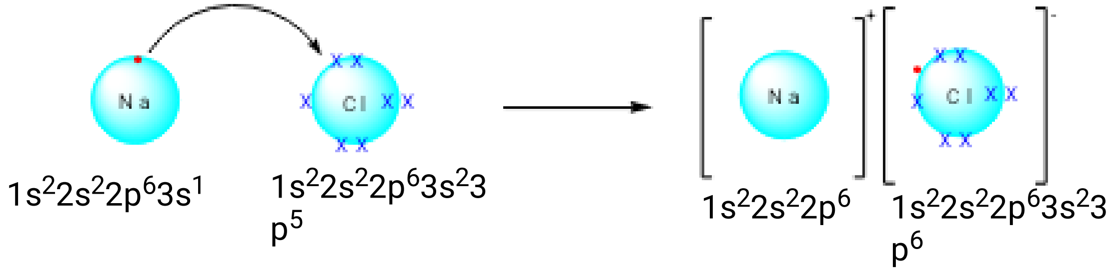
COURSE SYNOPSIS

- Chemical Bonding: Electrovalent bond between ions. Covalent bonds. Coordinate covalent bond. Metallic bonds. Intermolecular forces. Hydrogen bonding and its influence on properties. Kinetic theory of matter

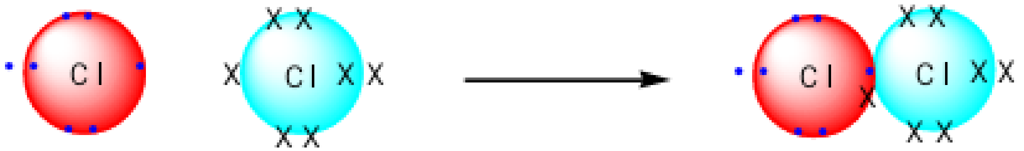
CHEMICAL BONDING

IONIC BONDING: involves transfer of electrons from metal (forming positive ion) to non-metal (forming negative ion). Example



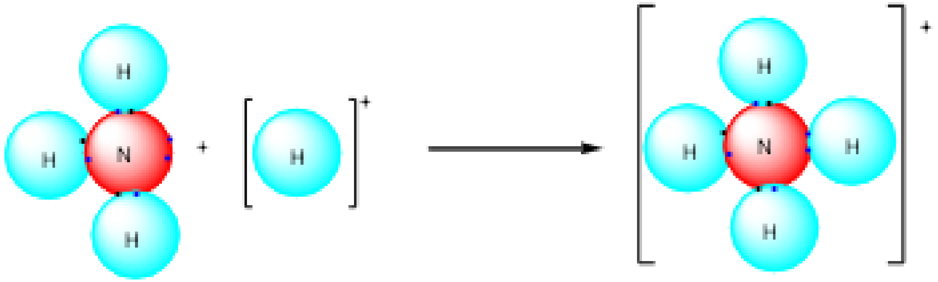
Na looses an electron while Cl gains an electron to attain octet configuration. Other examples: $MgCl_2$, Li_2O

COVALENT BONDING: involves sharing of electrons. Other examples: O_2, H_2, N_2



Chlorine molecule (Cl_2)

Coordinate Covalent Bonding: the shared pair of electrons is donated by only one of the bonded atoms



Properties of ionic compounds

- Solid with high melting and boiling points
- Conduct electricity when molten, but not as a solid
- It is made from a combination of metal and non-metal elements

Properties of covalent compounds

- Low boiling/melting points
- Soft in nature and relatively flexible
- Poor conductors of heat and electricity
- It is made from combination of non-metal elements

Metallic bonding: is the electrostatic attraction between the delocalized electron cloud and metal ions.

It holds metal ion tightly and confers high melting point.

Factors affecting metallic bonding

- Size of the metal ion: it increases with decrease in size of metal ion
- Number of valence electrons: it increases with increase in number of valence electrons

Intermolecular forces: two types – van der Waal’s forces and hydrogen bonding. Both describe electrostatic attraction between dipoles i.e. the attraction between the positive end of one molecule and negative end of another molecule

Types of Dipoles:

Permanent dipole: it exits in all polar molecules due to difference in electronegativity

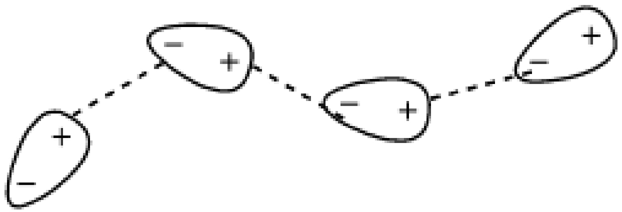
Instantaneous dipole: it occurs as a result of fluctuations in the electron cloud

Induced dipole: is created due to influence of a neighbouring dipole

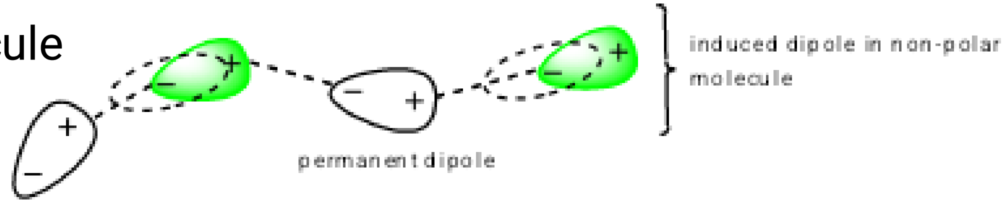
Temporary

Van der Waal's: includes dipole-dipole, dipole-induced dipole and instantaneous dipole-induced dipole interactions

Dipole – dipole interactions: it occurs when polar molecules orientate themselves in order to maximise attractive forces between molecules while repulsive forces are minimized



Dipole - induced dipole: when a non-polar molecule approaches a polar molecule, a dipole will be induced in the non-polar molecule



Instantaneous dipole - induced dipole: it occurs from fluctuations of electron cloud and can induce dipole moment in neighbouring atoms.The greater the number of electrons in a molecule, the easier it is for instantaneous dipole to be set up. The force of attraction between two temporary dipoles is known as a **London force** or **dispersion force**

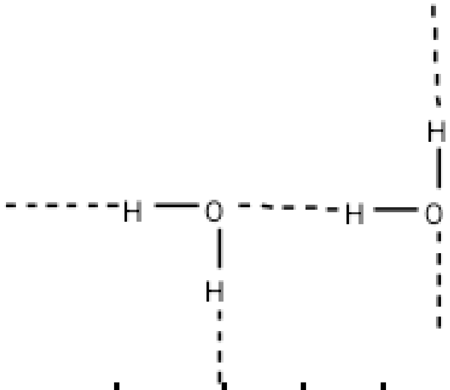
Hydrogen bonding: it exists when hydrogen is bonded to highly electronegative atom. It is the strong dipole-dipole interaction between hydrogen and electronegative atom (F, O, Cl, N). It is weaker than covalent bond but stronger than van der Waal’s forces. For hydrogen bonding to occur

- A hydrogen atom must be bonded directly to highly electronegative atom
- An unshared pair of electron on the electronegative atom

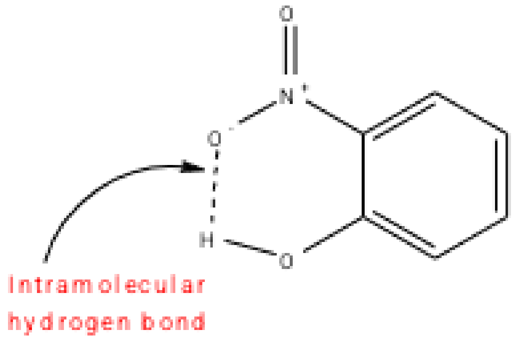
Hydrogen bonds are responsible for:

Intermolecular and intramolecular hydrogen bonding

Hydrogen bonds that occur between two different molecules are called intermolecular hydrogen bonds e.g H_2O and NH_3. Hydrogen bonds that occur within the same molecule are called intramolecular hydrogen bonds e.g 2-Nitrophenol.



(a) Intermolecular hydrogen bonds



(b) Intramolecular hydrogen bond

KINETIC THEORY OF MATTER

1) All matter is composed of atoms and molecules that are constantly in motion
2) The state of matter depends on the motion and arrangement of these particles, which is determined by temperature-higher temperatures mean higher particle motion
3) The three states of matter-solids, liquids and gases-can be distinguished based on how close together and how freely the particles can move

Note:

- At the same temperature, the heavier particles move slower than the lighter particles.
- Temperature is a measure of average kinetic energy (K.E) in a molecule

Solids, liquids & gases

Solids

- Solids have a **fixed** volume and shape and they have a high density.
- The atoms **vibrate** in position but can't change location
- The particles are packed very closely together in a fixed and regular pattern

Liquids

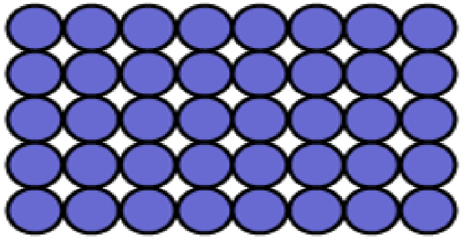
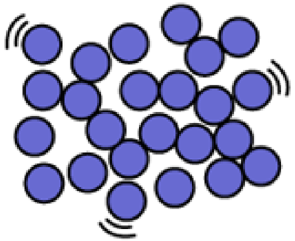
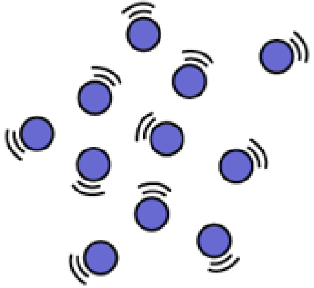
- Liquids also have a fixed volume but adopt the shape of the container
- They are generally less dense than solids (an exception is water), but much denser than gases
- The particles **move** and **slide** past each other which is why liquids adopt the shape of the container and also why they are able to flow freely

Solids, liquids & gases

Gases

- Gases do not have a fixed volume, and, like liquids, take up the shape of the container (Indefinite shape and volume)
- Gases have a very low density
- Since there is a lot of space between the particles, gases can be compressed into a much smaller volume
- The particles are far apart and move randomly and quickly (around 500 m/s) in all directions
- They **collide** with each other and with the sides of the container (this is how **pressure** is created inside a can of gas)

Summary of the properties of solids, liquids and gases

	Solid	Liquid	Gas
Diagram			
Arrangement of particles	Regular arrangement	Randomly arranged	Randomly arranged
Movement of particles	Vibrate about a fixed position	Move around each other	Move quickly in all directions
Closeness of particles	Very close	Close	Far apart
Density	High	Medium	Low
Energy of particles	Low energy	Greater energy	Highest energy

Phase Changes

- Melting/Freezing
- Boiling (vaporization)/Condensing
- Sublimation
- Evaporation

Melting/Freezing Point

- Change from solid to liquid and liquid to solid.

- Same temperature.; if melting, particles are gaining energy; if freezing, particles are losing energy.

- The stronger the intermolecular forces (IF’s), the more energy needed to weaken the IF’s, therefore higher melting point temperature.

Melting/Freezing Continued

- During the phase change, the temp. remains constant
- After all the sample has changed phase, the temp. will change.
- During the phase change, potential energy (P.E.) is changing, but K.E. is constant.

Boiling (Vaporization)/Condensation Point

- Change from liquid to gas and gas to liquid.
- Same temp.; if boiling, particles are gaining energy; if condensing, particles are losing energy
- The stronger the IF’s, the more energy needed to break the IF’s, therefore higher boiling point temperature

Boiling (Vaporization)/Condensation Point

- During the phase change, the temp. remains constant
- After all the sample has changed phase, the temp. will change.
- During the phase change, potential energy (P.E.) is changing, but K.E. is constant.

Sublimation

- Changing directly from a solid to a gas.
- Also, changing directly from a gas to a solid
- Skipping the liquid state.

Evaporation

- Liquid to gas but not necessarily at the boiling point temperature.
- Some particles gain enough K.E. to overcome the IF’s and become a gas.
- Remember, temperature is a measure of the average K.E.

Reading List:

- Abass A. Olajire. *Fundamental University Organic and Inorganic Chemistry*. Sina 2tees publication. 2008.
- Sachin Kr. Ghosh. *Advanced General Organic Chemistry (A modern Approach)*. New Central Book Agency (P) Limited. 2007
- Wong Y.C., Wong C.T., Onyinruka S.O., Akpanisi L.E. University General Chemistry Inorganic and Physical. Africana-FEP Publishers LTD