

Check list

Saturday, 3 February 2024 11:10 pm

4.5 Metallic bonding

- Metallic bonds involve a lattice of cations with delocalized electrons.
 - Atom in metallic solid lose their outer electrons and form a regular lattice of positive metallic ions (electron on the outer shell of the metal are delocalized)
- Metal atoms are held together by the electrostatic attraction between a lattice of positive ions and delocalized electrons.
 - Electron delocalized - electron become free to move around (the electron from outer shell) (from other note: delocalized means that the electron are not tied down to one specific atom)
 - o Positive cation attract by the free electron and form regular arranged structure (lattice), electron is between cation and free to move around.
 - o The metallic bond form between is cation and free moved electron which is due to electrostatic attraction between them.
 - Generally - strong attraction (require more energy than ionic bonding to break the metallic bonding, which also indicate there stronger electrostatic attraction inside metallic compound than ionic)

- Volatility - how easily a substance vaporized.
- Low volatility - high boiling/melting point
 - High volatility - easier to evaporate

Noble gas - group 18 of the periodic table

- The strength of the metallic bond increases with the charge on the cation and decreases with the radius of the ion.
 - More electron in metallic delocalized, higher melting point because electrostatic attraction is stronger (strength of bond change)
 - Bigger ion radius weaker charge density which ion do not hold the electron as tight as smaller ion
 - Charge of ion increase, which leads to the strong electrostatic attraction.
- The properties of metals – electrical and thermal conductivity, malleability, ductility – are a result of the delocalized electrons.
 - Conductivity
 - o (Conduction is caused by movement of charge particles which electrons have to move in free with the electrical field) - Metallic structure is good conductor
 - Melting point and boiling point
 - o High melting point as strong attraction between cation and free electron
 - Solubility
 - o Insoluble in water - as the strong electrostatic attraction inside metallic compound, do not impact by the polar water molecule
 - Malleability - good - even hit it
 - o electron free move around, when force exerted the metal, the nuclei (electron and cation) shift, but the bond do not break.
 - o Malleability - the quality of something that can be shaped into something else without breaking
 - Ductile
 - o Good ductile - they can be stretched into wire, as the electron free to move around, the nuclei shift when force exerted - stretched into shape without breaking
 - Metallic thermal conductivity (good)
 - o When heated, the delocalized electron gain kinetic energy
 - o The delocalized electron with higher kinetic energy move towards the particle in the lattice that has a lower temperature.

Define and Explain the formation of cations and anions in terms of electron configuration

- Atom transfer the electron in order to have full valence shell
- Cation + lost electron
- Anion - gain electron
- Lattice
 - Lattice structure - regular repeating pattern

Define and Explain what an Ionic Bond is, including its state (melting point) and lattice structure

- The electrostatic attraction between cation and anion
- Ionic compound have high melting and boiling point, because the strong electrostatic attraction between the ion require more energy to break it.
- The ions in the lattice structure follow a pattern positive and negative (in ionic compound the cation and anion is arranged in regular repeating structure)
- Polyatomic is an ion that is made up of more than one atom which together gained or lost an electron and therefore carried charge.
- Transition metal - set of metallic element occupying the central block in periodic table
- Lattice enthalpy - the energy require to break (1 mole) of an ionic lattice into its individual constituent ions.

Deduce the name and formula of ionic compounds

- NO₃⁻
- NH₄⁺
- SO₄²⁻
- CO₃²⁻
- HCO₃⁻
- OH⁻
- PO₄³⁻

Recall and explain the physical properties (volatility, electrical conductivity, solubility)

- Ionic solids are brittle. This is because the ionic bond are directional (fixed in place) (Brittleness the ability of the element break into small pieces)
 - when a force is applied the same charge come together and same charge repel and therefore an ionic solid shatters into small pieces when a force is applied.
- Ionic
 - Ionic compound have high melting and boiling point, because the strong electrostatic attraction between the ion require more energy to break it. This also means they have low volatility
 - Electron difference
 - o As the electron is lost more, the negative/positive charge of ion increase, therefore the electrostatic attraction between them increase
 - Size of ion (ionic radius)
 - o weaker charge density which ion do not hold the electron as tight as smaller ion this result in the lower lattice energy so can be easier break apart. (this also happens to isoelectronic - same electron left after transfer electron) - size (mass) of electron is different
 - o Because of this, smaller ionic have higher melting/boiling point
 - Solubility - how easily something dissolve in solvent (water)
 - o Soluble in ionic or polar solvent because the ionic compound is separated by the polar solvent therefore dissolved
 - o Insoluble in non-polar solvent - do not separated by the non polar solvent
 - Conductivity
 - o Conduct when dissolved - ions are free to move - conduct electricity (do conduct electricity in molten form as the ions is free moving) (not the anion is moving, the ion is moving)
 - o Conduct thermal (good when dissolved) - the free moving anions gain kinetic energy, move toward the particle have lower temperature, result in transfer of heat.
 - o Conduct thermal (bad when solid) - the particles are fixed in place whereas the thermal relies on the free moving particle to transfer heat, whereas the strong electrostatic attraction inhibit the electron move and held in their place

Use empirical data about physical properties to determine the type of bonding present in a number of substances

- Ionic - soluble, do not conduct in solid, conduct in liquid, bad malleability
- Covalent - polar covalent solve in polar solvent, non - polar covalent solve in non - polar solvent, do not conduct when dissolved, but sometime conduct in molten form (electron is the charge particle so once there are electron delocalized, it carries charge)
- Metallic - insoluble

Create a data table to record sufficient qualitative observations in order to identify bonding

Record data for the following physical properties: Melting point (> or < 800 °C), Conductivity of Solid, Solubility in water, Conductivity of Solution

Define and explain metallic bonds

- The metallic bond form between is cation and free moved electron which is due to electrostatic attraction between them.

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- The metallic bond form between is cation and free moved electron which is due to electrostatic attraction between them.
- Generally - strong attraction (require more energy than ionic bonding to break the metallic bonding, which also indicate there stronger electrostatic attraction inside metallic compound than ionic)

Explain how the charge and radius of the ions affect the metallic bond strength and therefore melting points

- More electron in metallic delocalized, higher melting point because electrostatic attraction is stronger (strength of bond change)
- weaker charge density (bigger radius of the ion) which ion do not hold the electron as tight as smaller ion
- Charge of ion increase, which leads to the strong electrostatic attraction.

Explain electrical conductivity and malleability of metals

- Physics properties
 - o (Conduction is caused by movement of charge particles (particle can be electron or ion etc.)) which electrons have to move in free with the electrical field) - Metallic structure is good conductor
 - o High melting point as strong attraction between cation and free electron
 - o Insoluble in water - as the strong electrostatic attraction inside metallic compound, do not impact by the polar water molecule
 - o More electron in metallic delocalized, higher melting point because electrostatic attraction is stronger (strength of bond change)
 - o Malleability - good - even hit it, electron free move around, when force exerted the metal, the nuclei (electron and cation) shift, but the bond do not break.
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Recall what an alloy is and explain their properties in terms of bonding

- Metal can also be grouped into:
 - Pure metals - metal made up from only one chemical element
 - Alloys - type of metal made up from a mixture of elements eg: copper + zinc

Alloying - metals are alloyed to improve the qualities of the individual pure metals

Define and explain covalent bonding

Use knowledge of valence electrons and the octet rule to construct Lewis diagrams

- Covalent bond - forms between non - metal, electron shared be able to achieve stable structure
 - The covalent bond is the electrostatic attraction between the positive nuclei and the shared pair of electron.

- Define a covalent bond in terms of the electrostatic attraction and identify its role in allowing atoms to fulfil the octet rule by sharing valence electrons.

- Deduce the Lewis formula of molecules and ions for up to four electron pairs on each atom.

Recognise and draw molecules that are stable despite having fewer than an octet of electrons.

- Determine the name of a simple molecular compound from its chemical formula.

- Deduce the Lewis formula for organic and inorganic molecules and ions with double and triple bonds present.

- Explain the relationship between the number of bonds and bond length and bond strength.

- Identify coordination bonds in compounds.

- Describe the formation of coordination bonds and their importance.

- Describe the concept of an electron domain.

- Predict the electron domain geometry and the molecular geometry for species with up to four electron domains.

- Predict the impact of non-bonding pairs and multiple bonds on bond angles.

- Deduce the polar nature of a covalent bond from electronegativity values.

- Recognise bond dipoles represented with partial charges and vectors.

- Deduce the net dipole moment of a molecule or ion by considering bond polarity and molecular geometry.

- Describe allotropes in terms of structural and bonding patterns, and resulting impacts on chemical and physical properties.

- Describe the structures and explain the properties of silicon, silicon dioxide and carbon's allotropes: diamond, graphite, fullerenes and graphene.

- Identify the types of intermolecular forces present in molecules as dipole-dipole, dipole-induced-dipole and London (dispersion) forces, encompassed by the umbrella term 'van der Waals forces'.

- Deduce the type(s) of intermolecular force present from the electronegativity difference and structural features of molecular compounds.

- Describe the main features of each type of intermolecular force.

- Describe hydrogen bonding as a type of intermolecular force present in some molecules based on the structural features of molecular compounds.

- Compare the relative strengths of the types of intermolecular forces.

- Discuss the influence of molecular size and polarity on the strength of London (dispersion) forces.

- Explain the physical properties of covalent substances to include volatility, electrical conductivity and solubility in terms of their structure.

- Identify the difference between a dye and a pigment.

- Outline the process of chromatography in terms of phases and intermolecular forces acting.

- Explain, calculate and interpret the retardation factor values, R_f .

Higher level (HL)

- Deduce the resonance structure of molecules and ions.

- Outline the role of electron delocalisation in formation of resonance hybrid structures.

- Discuss the structure of benzene based on physical and chemical evidence.

- Represent species with five and six electron domains around a central atom using Lewis formulas.

- Deduce the electron domain geometry and the molecular geometry for these species using the VSEPR model.

- Apply formal charge to determine the preferred Lewis formula when alternative Lewis formulas exist for a species.

- Describe sigma and pi bonds in terms of atomic orbital overlap.

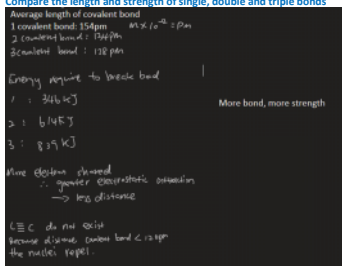
- Deduce the presence of sigma bonds and pi bonds in organic and inorganic molecular compounds and polyatomic ions.

- Analyse the hybridisation and bond formation in molecules and ions.

- Identify the relationships between Lewis formulas, electron domains, molecular geometry and type of hybridisation.

- Predict the geometry around an atom from its sp, sp² and sp³ hybridisation, and vice versa for both organic and inorganic examples.

- Compare the length and strength of single, double and triple bonds



- Recall that Be and B can break the octet rule
- B and Be is the exception of octet rule (incomplete octet)

- Define and identify molecules with coordinate (dative) covalent bonds

- Expanded octet - more than 8 electron on a atom's shell
- Coordinate bond - the bond that form where both the shared electrons originate from one atom

- Recall, Apply and explain VSEPR theory to predict the electron domain and molecular geometry of molecules
- VSEPR - As the increase of the bonding pair, each bonding pair (also called bonded pair of electron) spreads out as far as possible to each other due to repulsion of the same charge

- Predict bond angles for 3-dimensional molecules
- As the increase of bonding pairs, the bonding angle changes and their name of electron geometry and molecule geometry changes.
- Electron domain - how many pair electron surround atom
- Electron geometry: the number of electron pairs around nucleus
- Molecular geometry: number of bonding pair

- Construct 3-D models of molecules

- the repulsion of the pair electrons on the valence shell and bonding pairs is stronger than the bonding pairs and bonding pair rs.
- The order of repulsion: lone pair - lone pair > lone pair - bonding pair > bonding pair - bonding pair

- Define and explain electronegativity and link to whether a bond is ionic, ionic with covalent character, non - polar covalent, or polar covalent

Polar covalent bonds

- Sharing electron in covalent bond is not even. Some element have stronger attraction to electrons- this depends on the electronegativity - more share (electronegativity is depends on the proton) (this result in dipole and a polar bond)
- Electronegativity difference between that two element decide whether the polar bond is formed if there is a difference in electronegativity the bond is polar - however, if there's little difference between electronegativity, thus the compound non polar.

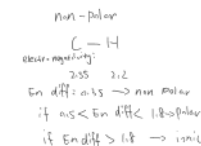
- Apply ideas of electronegativity to 3-D molecules to determine if they are polar or non-polar molecules

Polar covalent bonds

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Electronegativity of an element is a measure of its ability to attract to pair of electron

- Every element has a value from 0-4 for electronegativity



- If the molecule have polar bond, molecule could be polar as long as they are not cancel out
- Decide whether the molecule are polar - see whether there is vector canceling (polar cancelling) (or there is symmetry) (sometime the molecule can be polar with non polar end which the overall is polar)
- An asymmetrical molecule with polar bonds will have an overall net dipole movement (eg: positive end and negative end) (vector chanceless)

- Distinguish between intramolecular and intermolecular forces

Intermolecular attraction - the attraction between molecule

Intermolecular force - force formed when there are attraction between molecule

- Intermolecular force have to overcome in order to let a substance to change its state. - phase change which force of attraction between molecule is overcome (not the bond between the atom expect there are chemical changes)

Intramolecular forces - the force that holds atoms together within a molecule

- Define, explain and apply three types of intermolecular forces London (dispersion) forces (induced dipoles), dipole -dipole and hydrogen bonding

London(dispersion) forces (all molecule have it even the electronegativity difference is low)

- Force between non-polar molecule or atom
- It became stronger as the molar mass increase (bigger molar mass more electron)

as the electron are not stationary, The electron distribute temporarily in the middle of a bond between two atom, but sometime es, some electrons may slightly closer to one side than the other (this could be reversed) so there have temporarily dipole. This pres ence of these temporary dipoles causes attraction between neighboring molecule that both experience this asymmetry in electron distribution (It is the attraction between molecule which cause by the moving electron temporarily distribution in the molecule.)

Dipole - dipole - happen in the polar (permanent dipole) molecules

The attraction between the partial positive of one polar molecule and the partial negative of the another polar molecule (between permanent dipole (exist in asymmetrical molecule))

- Stronger than the dispersion forces

Hydrogen bonding (dipole - induced - dipole)

- Type of Dipole - dipole bond
- Only exist in N, O, F bond to H - molecule has special dipole - dipole force
- FON element must have a non - bonding pair of electron
- The attraction between the S+ hydrogen bonded FON with non bonding pair of electron molecule and nearby molecule that has the similar structure which electronegative element - H.
- The hydrogen bond form when the hydrogen bond with FON, in order to have intermolecular exist, the molecule must contain the lone pair

- Strong hydrogen bond form because the electronegativity of H, as element allow H have strong S+ (S+ is high as hydrogen has n o inner electron to shield the impact of its nucleus) make the S+ hydrogen very attract to non -bonding pair of electron on neighboring molecules

Strength of intermolecular force

Molecular size (molar mass) increase - the strength of intermolecular force increase

- The greater the number of electron in molecule, greater probability of electron distribute asymmetry or more susceptible the molecule develop an induced (bring about or give rise to/bring on) molecule from a charge nearby

For dipole - dipole (which has most polarity)

For hydrogen bonds (which has most nonbonding pair of electron)

For solubility

Non polar - polar (polar difference) < hydrogen bond (if all the choice are non soluble in first condition) < dipole - dipole bond < London dispersion

For isomers:

Intermolecular forces of attraction also depend upon the contact surface area between the molecules. Less surface area exposed for attraction less attraction

Solubility of general intermolecular force:

Depends on the molecule

Generally, polar molecules dissolve in polar solvent, non-polar molecule dissolve in non-polar solvent

Solubility of some polar molecule decrease as the size of the molecule increase, as the polar part of the molecule become relatively small compared to the sized of the overall molecule.

(covalent bond cannot break in the water) (therefore covalent network structure are generally insoluble in solvent)

The question sometimes says the intermolecular force in liquid state, because the molecule are free to move in liquid, so they could have the the intermolecular force, otherwise they are fixed in place...

- Recall what van der Waals forces are

Van der Waals forces - general term used to define the attraction of intermolecular forces between molecules

- Apply a molecules polarity (type of intermolecular forces) to its physical properties of volatility (melting point), electric al conductivity and solubility

For solubility

Non polar -polar (polar difference) > hydrogen bond > dipole - dipole bond > London dispersion

Compare between molecule

- Hydrogen bond>dipole-dipole>London dispersion force
- Indicate the physical properties.

Molecule are generally not electrically conductive as they do not have charged particle present

- Define the word allotrope

Allotrope: different structural forms of the same element - this leads to the element have different chemical and physical properties.

- occur when an element can exist in different crystalline forms such as in diamond, graphite, graphene, silicon, silicon dioxide, fullerenes these are also called as giant covalent structure

Allotrope of some element is the giant structure, these material strong covalent bonds join atoms together with other atoms of the same type to make giant structure rather than little group

- Giant covalent

- Strong attraction between the atom (attraction between nuclei and electron)
 - Normally Insoluble

- Recall which substances are giant covalent structures

diamond, graphite, graphene, silicon, silicon dioxide, fullerenes

- Recall and explain the bonding and structure of Si and C allotropes and SiO₂

Silicon and C Allotrope

- Have multiple allotropic forms - most stable SiO₂
 - SiO₂
 - Tetrahedral coordination are covalently bonded with four oxygen atoms around a central Si atom
 - High MP and BP, insoluble.
 - Play rule in construction, glass manufacturing
 - Silicon
 - Diamond cubic crystal structure
 - Low electrical conductivity but conduct electricity - semiconductor - the electron are not strongly held likely it is in diamond
 - Allotrope of carbon
 - Each carbon atom is covalently bonded to four other carbons in a tetrahedral geometry.
 - Diamond structure is the carbon is bonded with 4 carbon
- Crystalline silicon -
 - structure similar to diamond
 - Cubic lattice

- Explain the properties of Si and C allotropes and SiO₂ in terms of their bonding and structure

- Define, explain and apply the concept of resonance to molecules such as benzene, carbonate ion and ozone and others

Resonance hybrids

Resonance hybrids may occur when there is more than one possible position for a double bond in a molecule

Resonance - the Lewis structure of a molecule that has more than one possible position for a double bond

The molecule achieve this by delocalizing one of pair of electron - means no longer fixed to one bonding position: its spread over two positions (the position of atom is constant)

Electron delocalization is represent using a dashed line

Resonance hybrid structure have intermediate properties as they lies between these number of extreme structure - bond length in resonance is same

Strength of bond (single bond 1, double bond 2, triple bond 3)

In a resonance hybrid the bond order is 1.5

Benzene and ozone and carbonate ion is typical resonance.

In resonance structures, the electrons are able to move to help stabilize the molecule. This movement of the electrons is called delocalization. Resonance structures should have the same number of electrons, do not add or subtract any electrons.

- Deduce the formal charge for a molecule and apply to choose the most likely Lewis Structure

Formal charge

The charge of the each atom in a molecule - larger structures and more in depth explanations of bonding systems often require more sophisticated concepts and theories of bonding (eg: expanding octet)

- This can use to decide which Lewis structure is more likely to be between resonances

Formal charge = ve - (Bonding pair + nonbonding electron(or is the lone pair / 2))

B - bonding pair on atom

Ve = valence shell on neutral atom

Formal charge of molecule, all up all the formal charge on each atom

Formal charge in resonance - each hybrid resonance has the same molecular formula but they differ from the formal charge.

Formal charge = the number of valence electron around that atom in free atom - the number of nonbonding electron - 1/2 number of bonding electron

To determine the structure is more likely to be, choose the structure with formal charge is zero or the absolute value as low as possible.

- When all the overall charge is same, less formal charge on each atom is better.

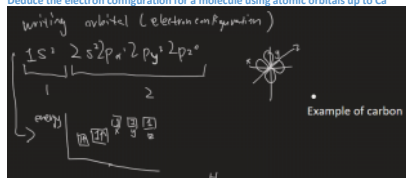
- Name and explain what s and p atomic orbitals are

Atomic orbitals a mathematical term that describes the position and wavelike behavior of an electron in an atom

Definition: Atomic orbital is a region of space around the atom nucleus where electron are likely to be found, this is characterized by s specific energy level and shapes

- Use to determine the probability of finding an electron(belonging to an atom) in a specific region around the nucleus of the atom (where an electron is most likely to be found)
- Each atomic orbital can hold a maximum of 2 electrons

- Deduce the electron configuration for a molecule using atomic orbitals up to Ca



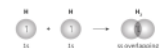
- Write configuration

- Number of shell - type of orbitals with direction (x,y,z) - number of electron on orbital
 - Each number can hold up to 2
 - Follow up each other

Electron are found in orbitals, when orbital are filled with maximum number of electron, the octet rule has been fulfilled an atom are stable, and when it is empty, it gains or lose electrons(ionic bond) or it will share electron so it can fulfill the octet rule.

- Define and explain the formation of sigma and pi molecular bonds using atomic orbitals

Single bond form between s - orbitals



Bond of atom form when two orbitals overlap(When orbital merge each other, covalent bond form)

Sigma bond form when orbital - overlap is a head - on overlap

- Electron are likely to be found in that overlapping shared space
- Single bond that atom form are always sigma bonds
- Represent in σ
- Head on overlap include axial overlap or end-to-end overlap
- Can be formed from the head on overlap of s-s, s-p, p-p

Orbitals: where an electron is most likely to be found

Area of overlap: indicate the area where the probability of finding an electron is highest.

Pi bond

For double and triple bond molecules

- The electron pair forming the double/triple bond can not lie in the same axial plane as the electron pair forming the sigma bond as they would repel each other: they must lie somewhere else between the two atom.
- Only p orbitals are used to form more bond while keeping sigma bond
 - Due to the shape of overlapping

Pi bond - if the orbital overlap is a sideways or lateral overlap

- Deduce whether sigma or pi bonds are present in a molecule

- Explain what sp, sp² and sp³ hybridisation is

Molecular shapes can be arrived at either by using VSEPR theory or by knowing the type of hybridization. Hybridization can take place between any s and p orbital in the same energy level, and is not restricted to carbon compound. If the shape and angle are known then the type of hybridization can be deduced. Vice versa

Graphite and C₆₀ fullerene structures involve each carbon atom attached to three other carbon atoms.

Diamond has no delocalised electrons, whereas graphite and C₆₀ fullerene do have delocalised electrons. Diamond has bond angles of 109, whilst graphite has bond angles of 120.

Covalent network solids are made up of atoms bonded together by strong covalent bonds, forming a continuous network throughout the structure. Here's a simple breakdown of the structures of diamond, graphite, graphene, silicon, and silica:

Diamond: Each carbon atom in diamond is bonded to four other carbon atoms in a tetrahedral arrangement, forming a three-dimensional network structure. This makes diamond extremely hard and with a high melting point.

Graphite: In graphite, carbon atoms are arranged in layers of hexagonal rings. Within each layer, carbon atoms are bonded together by strong covalent bonds, but between layers, the bonding is weak van der Waals forces, allowing the layers to slide over each other. Graphite is a good conductor of electricity in the plane of its layers.

Graphene: Graphene is a single layer of graphite, consisting of a two-dimensional hexagonal lattice of carbon atoms. It is incredibly strong, lightweight, and has excellent electrical conductivity.

Silicon: Silicon atoms in a covalent network solid are arranged similarly to carbon in diamond, forming a three-dimensional network structure. Silicon shares many properties with diamond, including hardness and high melting point.

Silica: Silica, or silicon dioxide (SiO₂), has a similar structure to diamond but with silicon atoms bonded to oxygen atoms instead of other silicon atoms. This forms a three-dimensional network structure of alternating silicon and oxygen atoms, making silica a hard and rigid material, commonly found in quartz and glass.

Normal covalent structures, have molecules held together by covalent bonds, which are strong. But in giant covalent structures, each atom is bonded to many others in a huge network. This makes the structure very strong and stable, so it takes a lot more energy to break these bonds, which is why giant covalent structures have higher melting points than normal covalent structures.

Sp3	109 degree - tetrahedral	Nitrate ...
Sp2	120 degree - trigonal planar	Ethene,diazene,methanol
sp	180 degree - linear	Nitrogen, C2H2

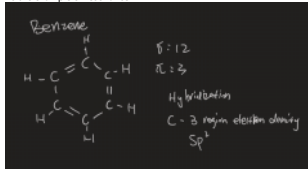
- Deciding the hybridization depends on the number of region of electron density.

Benzene

There's evidence that can show the Benzene does not behave like a molecule with single and double bonds (which single and double bond do not exist in resonance - all the bond length in resonance are same)

- Physics evidence
 - o All bond length and angles are the same
 - o Molecular geometry of each carbon atom is trigonal planar

As the example of resonance



Physical evidence of the properties of benzene can include its molecular structure, which consists of a planar hexagonal ring of carbon atoms with alternating single and double bonds. Benzene is a colorless liquid with a characteristic sweet odor. It has a relatively low boiling point of 80.1° C and melting point of 5.5° C. Additionally, benzene is less dense than water and is insoluble in water but soluble in organic solvents like alcohol and ether.

• Deduce what type of hybridisation is present in molecules

Hybrid orbital - when covalent bond form (when orbital merges (overlaps))

Hybridization - the orbitals that molecule are made up of

When look at the molecular shape (where the molecule are made up) - this also called as hybridization of molecule

- Tetrahedral sp3
- Trigonal planar sp2
- Linear sp

• Link hybridisation to electron domains and molecular geometry

• Explain the wavelength of light required to dissociate oxygen and ozone molecules.

• Describe the mechanism of ozone depletion catalysed by CFCs and NO_x.

Transitional metal

At the middle of periodic table

Metal between the left hand side

Left hand side metal are soft, low melting point, highly reactive

Transition metal has high melting point and boiling point

- No reaction or slow chemical reactivity
- Highly colored