Chemistry guide 1: Covalent bonds

How to use: This guide won't explain everything but supplement the syllabus, so read the syllabus things first and then the things it omits. It's good to try and answer each section from the syllabus beforehand and then check the answer. If you already understand a certain aspect of the syllabus then skip it (or read it just to make sure you know it). Don't forget to do a lot of practice questions.

STANDARD AND HIGHER LEVEL CONTENT

Structure 2.2—The covalent model

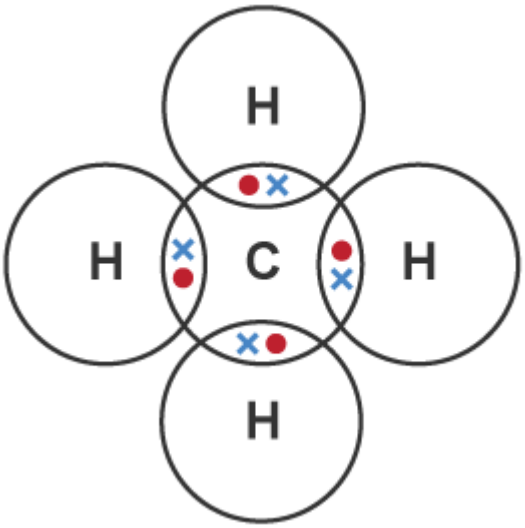
Guiding question: What determines the covalent nature and properties of a substance?

Standard level and higher level: 10 hours

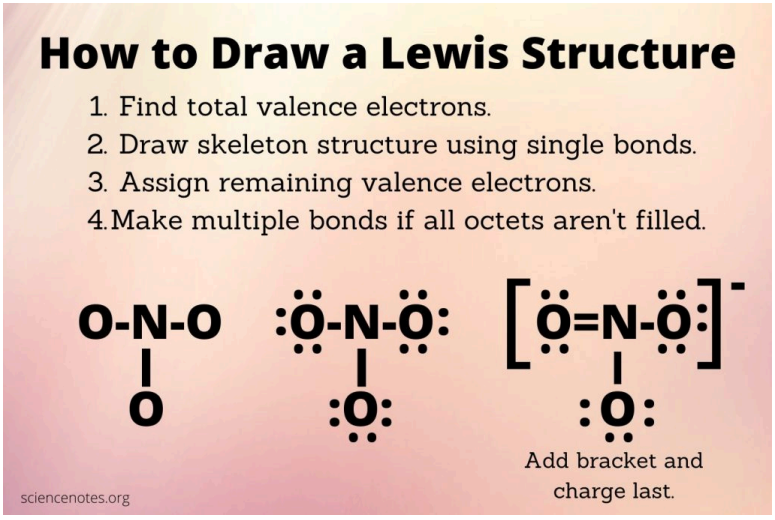
<table>
<tr><td colspan="2">Structure 2.2.1—A covalent bond is formed by the electrostatic attraction between a shared pair of electrons and the positively charged nuclei.
The octet rule refers to the tendency of atoms to gain a valence shell with a total of 8 electrons.
Deduce the Lewis formula of molecules and ions for up to four electron pairs on each atom.</td></tr>
<tr><td>Lewis formulas (also known as electron dot or Lewis structures) show all the valence electrons (bonding and non-bonding pairs) in a covalently bonded species.
Electron pairs in a Lewis formula can be shown as dots, crosses or dashes.
Molecules containing atoms with fewer than an octet of electrons should be covered.
Organic and inorganic examples should be used.</td><td>Nature of science—What are some of the limitations of the octet rule?
Structure 1.3—Why do noble gases form covalent bonds less readily than other elements?
Structure 2.1—Why do ionic bonds only form between different elements while covalent bonds can form between atoms of the same element?</td></tr>
</table>

Explanation:

Here's how to draw covalent bonds, they share electrons. This is a dot and cross diagram (of methane) which differentiates the source of the electrons. The bond is the electrostatic attraction between the shared pair of electrons and the nuclei between each species. So methane has four covalent bonds. Bonding pairs are simply the pair of electrons being shared and non-bonding pairs are not bonded.



This is how to draw Lewis structures:



Drawing them can be confusing. Above the bonded nitrogen has 4 bonds but base nitrogen has 5 electrons. This is to satisfy the octet rule.

Limitations of the octet rule:

Incomplete octets:

Beryllium: only 4 electrons
Boron: only 6 electrons
Hydrogen: only 2 electrons

Expanded octets occur from period 3 down.

Noble gases form bonds less readily because they already have full octet shells.

Ionic bonds require a large electronegativity difference (typically between metals and non-metals) so electrons are transferred. Thus, identical atoms cannot form ionic bonds.

Covalent bonds form when electron transfer is not favorable (similar or identical electronegativities).

Structure 2.2.2—Single, double and triple bonds involve one, two and three shared pairs of electrons respectively. Explain the relationship between the number of bonds, bond length and bond strength.	
	Reactivity 2.2—How does the presence of double and triple bonds in molecules influence their reactivity?

Relationship between number of bonds, bond length and bond strength:

As the number of bonds increases, the electrostatic attraction pulling the nuclei together grows stronger with a greater charge. As the strength increases, the length decreases because the attraction is stronger, decreasing the distance between the pair of electrons and nuclei.

Reactivity, the influence of double and triple bonds:

In organic chemistry, molecules with multiple bonds (like alkenes or alkynes) are much more reactive towards electron-seeking species (electrophiles) because the extra electron pairs (pi bonds) are more accessible and easier to break than single (sigma) bonds.

Double and triple bonds cause areas of increased electron density. But the pi bonds are still weaker and more accessible (they are below and above the internuclear axis) so double and triple bonds increase reactivity

Structure 2.2.3—A coordination bond is a covalent bond in which both the electrons of the shared pair originate from the same atom. Identify coordination bonds in compounds.	
AHL—Include coverage of transition element complexes.	AHL Reactivity 3.4—Why do Lewis acid–base reactions lead to the formation of coordination bonds?

Coordination Bond: A covalent bond where both electrons in the shared pair are contributed by the same atom. Once formed, a coordinate bond is identical in strength and behavior to a regular covalent bond.

Lewis acid-base reactions:

Lewis base: Electron pair donor (2 extra electrons)
Lewis acid: Electron pair acceptor (needs 2 electrons)

The lewis base gives the lewis acid a pair of electrons which is how coordinate bonds form.

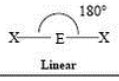
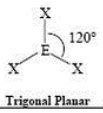
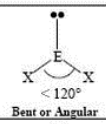
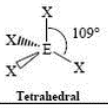
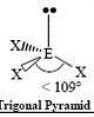
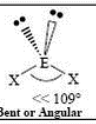
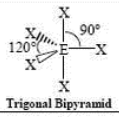
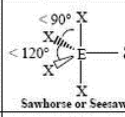
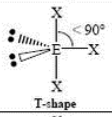
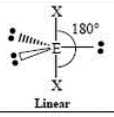
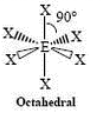
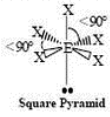
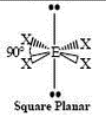
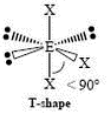
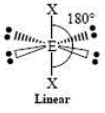
Transition element complexes: Transition metals act as lewis acids, they have a free d-orbital. They accept electrons from ligands (molecules or ions like H_2O, acting like lewis bases) to form complex ions.

Structure 2.2.4—The valence shell electron pair repulsion (VSEPR) model enables the shapes of molecules to be predicted from the repulsion of electron domains around a central atom.	
Predict the electron domain geometry and the molecular geometry for species with up to four electron domains.	
Include predicting how non-bonding pairs and multiple bonds affect bond angles.	Nature of science—How useful is the VSEPR model at predicting molecular geometry?

VSEPR Theory: Valence Shell Electron Pair Repulsion theory states that electron domains around a central atom repel each other and position themselves as far apart as possible to minimize repulsion.

Electron Domain: Any region of electron density around a central atom.
Note: A single bond, double bond, triple bond, or lone pair each counts as one single electron domain

VSEPR Geometries

Steric No.	Basic Geometry 0 lone pair	1 lone pair	2 lone pairs	3 lone pairs	4 lone pairs
2	 Linear				
3	 Trigonal Planar	 Bent or Angular			
4	 Tetrahedral	 Trigonal Pyramid	 Bent or Angular		
5	 Trigonal Bipyramid	 Sawhorse or Seesaw	 T-shape	 Linear	
6	 Octahedral	 Square Pyramid	 Square Planar	 T-shape	 Linear

Electron Geometry	Molecular Geometry
It refers to the name of the geometry of the electron pair/groups/domains on the central atom, whether they are bonding electrons or non-bonding electrons.	It is the three-dimensional structure of atoms in a molecule and it determines the shape of a molecule around the central atom.
It helps understand how different electron groups are arranged in a molecule.	It helps understand the entire atom and its arrangement.
It is determined on the basis of both bonding electron pairs and lone electron pairs.	It is determined on the basis of only bonding electron pairs, not the number of electron pairs.

DB Difference Between.net

Factors Affecting Bond Angles: Lone Pair Repulsion: Lone pairs are held closer to the central nucleus than bonding pairs, so they exert greater repulsion

Lone pair–Lone pair >Lone pair–Bonding pair >Bonding pair–Bonding pair

Each lone pair reduces the ideal bond angle by approximately 2° to 2.5°.

Multiple Bonds: Multiple bonds contain high electron density and occupy more space than single bonds, slightly increasing repulsion and squeezing adjacent bond angles.

Limitations: Does not account for exact orbital hybridization, fails to accurately predict transition metal complex shapes, and assumes equal repulsion for all single bonds regardless of electronegativity differences.

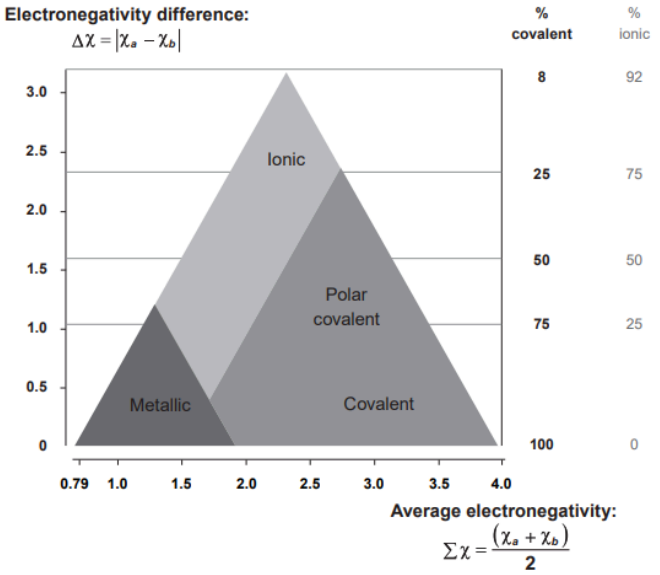
Strengths:

- Distinguishes between electron domain geometry and the physical molecular geometry
- Accounts for lone pair distortions
- Works for standard molecules and expanded octets.

Structure 2.2.5—Bond polarity results from the difference in electronegativities of the bonded atoms. Deduce the polar nature of a covalent bond from electronegativity values.	
Bond dipoles can be shown either with partial charges or vectors. Electronegativity values are given in the data booklet.	Structure 2.1—What properties of ionic compounds might be expected in compounds with polar covalent bonding?

You will use this diagram:

17. Triangular bonding diagram (van Arkel–Ketelaar triangle)



And the electronegativity values from here:

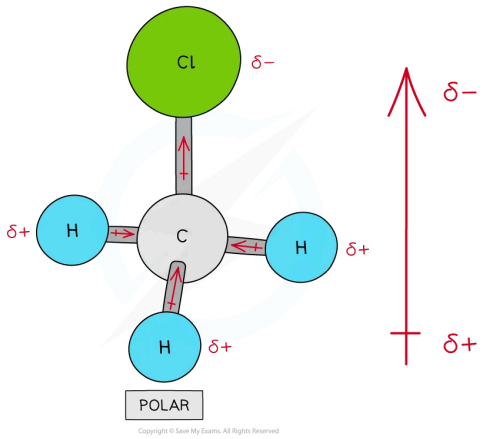
9. First ionization energy, electron affinity and electronegativity of the elements

1312 −73 **H** 2.2																	2372 **He**
520 −60 **Li** 1.0	900 **Be** 1.6											801 −27 **B** 2.0	1086 −122 **C** 2.6	1402 **N** 3.0	1314 −141 (+753) **O** 3.4	1681 −328 **F** 4.0	2081 **Ne**
496 −53 **Na** 0.9	738 **Mg** 1.3											578 −42 **Al** 1.6	787 −134 **Si** 1.9	1012 −72 **P** 2.2	1000 −200 (+545) **S** 2.6	1251 −349 **Cl** 3.2	1520 **Ar**
419 −48 **K** 0.8	590 −2 **Ca** 1.0	633 −18 **Sc** 1.4	659 −8 **Ti** 1.5	651 −51 **V** 1.6	653 −64 **Cr** 1.7	717 **Mn** 1.6	762 −15 **Fe** 1.8	760 −64 **Co** 1.9	737 −112 **Ni** 1.9	745 −119 **Cu** 1.9	906 **Zn** 1.6	579 −41 **Ga** 1.8	762 −119 **Ge** 2.0	944 −78 **As** 2.2	941 −195 **Se** 2.6	1140 −325 **Br** 3.0	1351 **Kr**
403 −47 **Rb** 0.8	549 −5 **Sr** 1.0	600 −30 **Y** 1.2	640 −41 **Zr** 1.3	652 −88 **Nb** 1.6	684 −72 **Mo** 2.2	702 −53 **Tc** 2.1	710 −101 **Ru** 2.2	720 −110 **Rh** 2.3	804 −54 **Pd** 2.2	731 −126 **Ag** 1.9	868 **Cd** 1.7	558 −29 **In** 1.8	709 −107 **Sn** 2.0	831 −101 **Sb** 2.0	869 −190 **Te** 2.1	1008 −295 **I** 2.7	1170 **Xe**
376 −46 **Cs** 0.8	503 −14 **Ba** 0.9	538 −45 **La** † 1.1	659 −1 **Hf** 1.3	728 −31 **Ta** 1.5	759 −79 **W** 1.7	756 −14 **Re** 1.9	814 −106 **Os** 2.2	865 −151 **Ir** 2.2	864 −205 **Pt** 2.2	890 −223 **Au** 2.4	1007 **Hg** 1.9	589 −36 **Tl** 1.8	716 −35 **Pb** 1.8	703 −91 **Bi** 1.9	812 −183 **Po** 2.0	−270 **At** 2.2	1037 **Rn**
393 −47 **Fr** 0.7	509 −10 **Ra** 0.9	499 −34 **Ac** ‡ 1.1															

†	534 −63 **Ce** 1.1	528 −83 **Pr** 1.1	533 −185 **Nd** 1.1	539 **Pm**	545 **Sm** 1.2	547 −83 **Eu**	593 **Gd** 1.2	566 −112 **Tb**	573 < 0 **Dy** 1.2	581 **Ho** 1.2	589 **Er** 1.2	597 −99 **Tm** 1.3	603 2 **Yb**	524 −33 **Lu** 1.0
‡	609 **Th** 1.3	568 **Pa** 1.5	598 **U** 1.7	605 **Np** 1.3	581 **Pu** 1.3	576 **Am**	578 **Cm**	598 **Bk**	606 **Cf**	619 **Es**	627 **Fm**	635 **Md**	642 **No**	473 **Lr**

Take the electronegativity from the periodic table and then plug it into both formulas of average electronegativity and electronegativity difference.

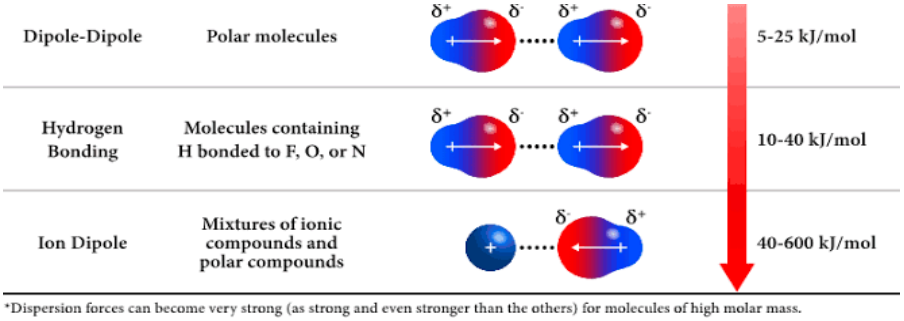
Here’s how to represent dipoles with partial charges:



The chlorine is more electronegative therefore it pulls more electrons towards itself and has a more negative charge than hydrogen. The squiggly line with a charge represents that dipole.

Ionic compounds and polar covalently bonded similarities:

- They both have areas of charge, positive and negative regions. Ionic compounds have full integer charges like 2- or 2+ whereas polar compounds have partial charges as seen above.
- Both dissolve in water and polar solvents. This is done through dipole-dipole or ion-dipole interactions.



- Both form to fulfil the octet rule and lower potential chemical energy (make it more stable)

Structure 2.2.6—Molecular polarity depends on both bond polarity and molecular geometry. Deduce the net dipole moment of a molecule or ion by considering bond polarity and molecular geometry.	
Examples should include species in which bond dipoles do and do not cancel each other.	AHL Structure 3.2—What features of a molecule make it "infrared (IR) active"?

Molecular polarity depends on bond polarity and molecular geometry. Sometimes the geometry of a molecule can cancel polarity. Remember that a difference in electronegativity makes a molecule polar, the difference creates a "net dipole moment", this is a vector from the less electronegative atom towards the more electronegative one. Though due to some VSEPR configurations, that polarity is canceled out if the vector sum is zero. If two vectors

point in opposite directions for example. Examples include CO_2 where oxygen is very electronegative but due to its linear structure, both C=O bonds pull in the same direction leading to a net zero vector and therefore the molecule is nonpolar. On the other hand, a molecule like H_2O is polar with its bent shaped and electronegative oxygen. Vectors aren't necessary in IB descriptions though.

Structure 2.2.7—Carbon and silicon form covalent network structures. Describe the structures and explain the properties of silicon, silicon dioxide and carbon's allotropes: diamond, graphite, fullerenes and graphene.	
Allotropes of the same element have different bonding and structural patterns, and so have different chemical and physical properties.	Structure 3.1—Why are silicon–silicon bonds generally weaker than carbon–carbon bonds?

	Properties	Structure
Silicon /Si	High thermal conductivity, a semiconductor, high thermal stability, low density.	Silicon has four covalent bonds with neighboring atoms. However it doesn't have the pyramid shape which diamond has.
Silicon dioxide /SiO_2	High melting point, very hard, electric insulator	Giant covalent network structure, each oxygen is bonded to 2 silicons but each silicon has 4 bonds with different oxygens.
Diamond / C	Very hard, very high melting points,	Carbons covalently bonded with 4 other carbons. Giant covalent crystal structure.
Graphite /C	Soft, conducts electricity (due to electrons) and high melting points	Layered sheets (grapene) of carbon atoms covalently bonded to 3 other carbon atoms with declocalised electrons
Fullerenes / C	Buckminster fullerene: a semi-conductor, individual molecules are very hard but a collection of them are not because molecules are held together by van der Waals forces. Nanotubes: very high tensile strength, semiconductor being very conductive along its length, good heat conductor as well	There are different types of fullerenes such as the buckminster ball or nanotubes. The Buckminister ball has 60 carbons, 20 hexagons and 12 pentagons.

Graphene / C	Very good electrical conductor, strongest material in the world per unit thickness so very high tensile strength, very good heat conductor.	One sheet of graphite with a hexagon structure.

Nature of science: Silicon-silicon bonds are weaker because silicon atoms are larger and have longer bond lengths. It has an extra inner electron shell whereas carbon has valence electrons on its second shell.

Structure 2.2.8—The nature of the force that exists between molecules is determined by the size and polarity of the molecules. Intermolecular forces include London (dispersion), dipole-induced dipole, dipole–dipole and hydrogen bonding. Deduce the types of intermolecular force present from the structural features of covalent molecules.	
The term "van der Waals forces" should be used as an inclusive term to include dipole–dipole, dipole-induced dipole, and London (dispersion) forces. Hydrogen bonds occur when hydrogen, being covalently bonded to an electronegative atom, has an attractive interaction on a neighbouring electronegative atom.	Structure 1.5—To what extent can intermolecular forces explain the deviation of real gases from ideal behaviour? Nature of science, Structure 1.1, 2.1, 2.3—How do the terms "bonds" and "forces" compare? Nature of science—How can advances in technology lead to changes in scientific definitions, e.g. the updated International Union of Pure and Applied Chemistry (IUPAC) definition of the hydrogen bond?

Van der Waals forces:

London dispersion forces: Present in all molecules. They are the only forces between non-polar molecules. Strength increases with larger molecular size / more electrons.

Dipole-induced dipole forces: Force between one dipole and another molecule which isn't a dipole. So they are found in mixtures containing polar (dipoles) and non-polar (no dipoles) molecules. The permanent dipole from the polar molecule alters the non-polar molecule electron cloud.

Dipole-dipole forces: Forces between mixtures of polar molecules. The permanent dipoles align, opposite charges attracting.

Hydrogen bonding: A specific dipole-dipole force which is a lot stronger than other van der waals forces. It happens when hydrogen bonds to fluorine, oxygen or nitrogen. This strength comes from the large difference in electronegativity.

In order to determine which forces apply, find whether each molecule is polar and not and then deduce based on the criteria above. Remember that each molecule also has forces with itself along with the other molecules in a mixture.

Nature of science:

Ideal gases: ideal gases carry the assumption that there are no attractive forces between molecules. Real gases do and this is evident at low temperatures and high pressure where the intermolecular forces pull the gases together. This reduces the frequency of collision with the container's wall causing the pressure to be lower than expected by the ideal gas laws.

Bonds and forces: Bonds include the transfer or sharing of electrons and occur within molecules whereas forces refer to electrostatic attraction between molecules and do not involve transfer/sharing of electrons. Forces are also a lot weaker.

Structure 2.2.9—Given comparable molar mass, the relative strengths of intermolecular forces are generally: London (dispersion) forces < dipole–dipole forces < hydrogen bonding. Explain the physical properties of covalent substances to include volatility, electrical conductivity and solubility in terms of their structure.	
	Tool 1, Inquiry 2—What experimental data demonstrate the physical properties of covalent substances? Structure 3.2—To what extent does a functional group determine the nature of the intermolecular forces?

Volatility:
Definition: How easily a substance turns into a gas
Strong intermolecular forces hold molecules tightly together meaning more energy is required to turn the substance into a gas. Therefore volatility decreases with increasing strength in forces. So structures with stronger forces, such as hydrogen bonding will generally be less volatile than those with weaker forces such as London dispersion forces.

Electrical conductivity:
Definition: The ability to carry an electrical charge
Covalent substances generally don't conduct electricity because they consist of neutral molecules. An exception is graphite which has delocalized electrons (and acids in water)

Solubility:
Definition: How well substances dissolve: "like dissolves like", polar solvents dissolve polar solutes and vice versa with non polar.
Solubility depends on whether the substance is polar and non-polar and if the solute has the same property: like dissolves like.

Functional groups: They are important because they affect the polarity and bonding capabilities. E.g. Hydroxyl, -OH, functional groups introduce strong hydrogen bonding. Carbonyl groups create dipole-dipole forces. However, hydrocarbon chains for example don't provide anything other than weak London forces. Also remember that symmetry is still important, CO_2 for example has two carbonyl groups though they cancel out due to symmetry: just because a molecule consists of polar functional groups does not mean they are polar.

Structure 2.2.10—Chromatography is a technique used to separate the components of a mixture based on their relative attractions involving intermolecular forces to mobile and stationary phases. Explain, calculate and interpret the retardation factor values, R_F.	
Knowledge of the use of locating agents in chromatography is not required. The technical and operational details of a gas chromatograph or high-performance liquid chromatograph will not be assessed.	Tool 1—How can a mixture be separated using paper chromatography or thin layer chromatography (TLC)?

Stationary phase: The fixed substance that doesn't move such as paper in paper chromatography or a thin layer of silica/alumina in thin layer chromatography
Mobile phase: The fluid, solvent, that moves through or over the stationary phase. It carries the components of the mixture with it.

Separation is "based on their relative attractions involving intermolecular forces to mobile and stationary phases". The better attracted the component is towards the mobile phase the further it will move up whereas if it's stronger attracted to the stationary phase it will stay nearer the baseline.

Retardation factor, Rf value: $\frac{Distance\ traveled\ by\ component}{Distance\ traveled\ by\ the\ solvent}$

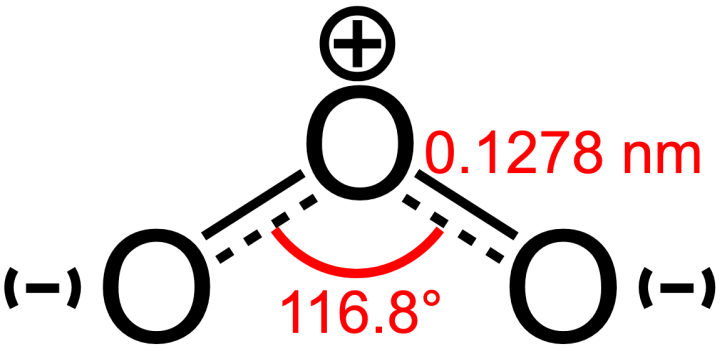
Lower Rf values mean the component is better attracted towards the stationary phases and higher Rf values mean the component is better attracted towards and dissolves better in the mobile phase.

HIGHER LEVEL CONTENT

Additional higher level: 8 hours

Structure 2.2.11—Resonance structures occur when there is more than one possible position for a double bond in a molecule. Deduce resonance structures of molecules and ions.	
Include the term "delocalization".	Structure 1.3—Why are oxygen and ozone dissociated by different wavelengths of light?

Resonance structures show electrons being shared across multiple bonds rather than locked into a fixed position. When a molecule has resonance structures, the electronic structure is a hybrid, an average, of all the possible Lewis structures. Electrons in pi bonds are not localized (they are delocalized) and instead spread out over multiple atoms. With delocalization, the bonds are identical in length and strength. The bond order could therefore be e.g. 1.5.

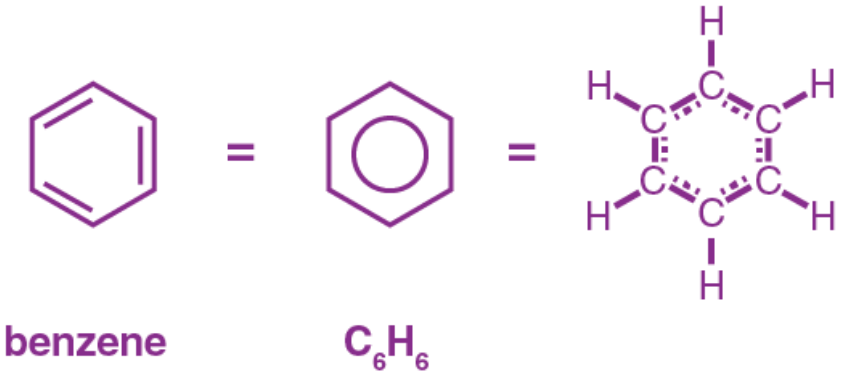


Deducing, e.g. ozone (above)
Add the dotted lines when there are multiple potential places for the double bond to be.

Oxygen and ozone: oxygen is dissociated by shorter wavelengths of light (higher energy) because they have stronger bonds compared to ozone which has delocalized bonds.

Structure 2.2.12—Benzene, C_6H_6, is an important example of a molecule that has resonance. Discuss the structure of benzene from physical and chemical evidence.	
	Reactivity 2.1, 2.2—How does the resonance energy in benzene explain its relative unreactivity? Reactivity 3.4—What are the structural features of benzene that favour it undergoing electrophilic substitution reactions?

This is the structure of benzene:



It's a ring. There can be 3 double bonds but the structure is a resonance structure because of the potential places in which it can be. Therefore it is shown as a hexagon with a ring for the delocalised bonds.

Evidence for this structure:

Physical evidence:

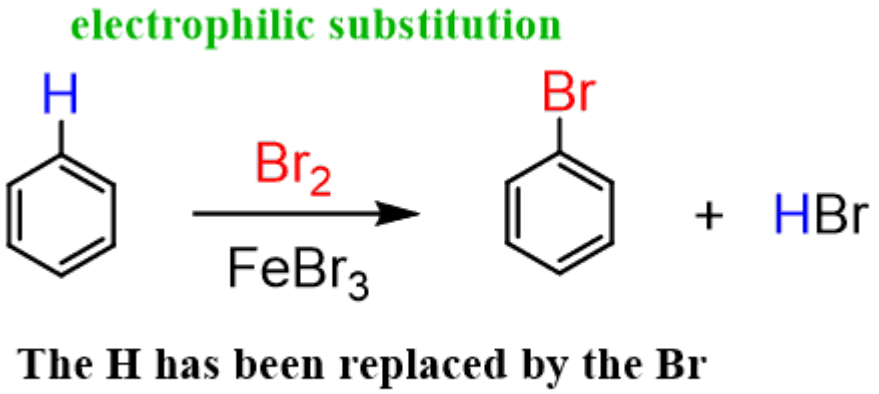
- All the C-C bond lengths are the same, between a double bond and a single bond
- Bond strengths are the same throughout, they have the same bond enthalpies between the enthalpies of single and double carbon bonds.
- Benzene is planar and a hexagon with internal bond angles of 120°

Chemical evidence:

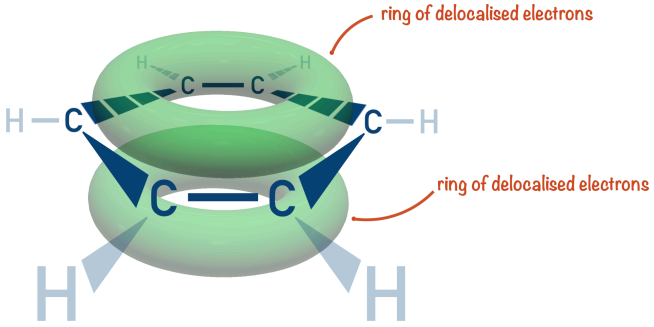
- Benzene resists additional reactions unlike alkenes.

Unreactivity: Benzene is relatively unreactive because it has a stable structure due to the delocalized bonds. There is a high activation energy to overcome the already stable structure.

Electrophilic substitution reactions:



A H atom is replaced by an electrophile (electron deficient species). They are attracted by the delocalized pi electron ring above and below the carbon ring.

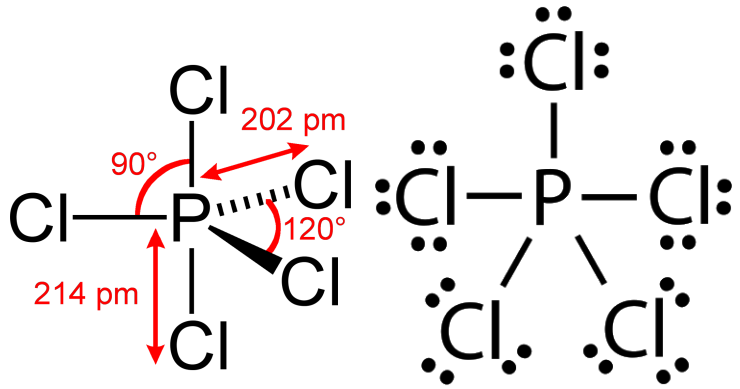


This restores the stable aromatic (word for special cyclic structures such as benzenes) structure of benzene.

Structure 2.2.13—Some atoms can form molecules in which they have an expanded octet of electrons. Visually represent Lewis formulas for species with five and six electron domains around the central atom. Deduce the electron domain geometry and the molecular geometry for these species using the VSEPR model.	
	Structure 3.1—How does the ability of some atoms to expand their octet relate to their position in the periodic table?

All elements in or below period three can form expanded octets (More than 8 electrons) but that doesn't mean they always do. Period 2 and above lack d orbitals to create expanded octets.

Example: Phosphorus pentachloride



VSEPR Geometries

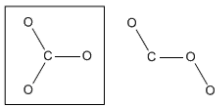
Steric No.	Basic Geometry 0 lone pair	1 lone pair	2 lone pairs	3 lone pairs	4 lone pairs
2	180° Linear				
3	120° Trigonal Planar	< 120° Bent or Angular			
4	109° Tetrahedral	< 109° Trigonal Pyramid	<< 109° Bent or Angular		
5	90°, 120° Trigonal Bipyramid	< 90°, < 120° Sawhorse or Seesaw	< 90° T-shape	180° Linear	
6	90° Octahedral	< 90° Square Pyramid	90° Square Planar	< 90° T-shape	180° Linear

Using the table above, it's apparent that the electron domain geometry is trigonal bipyramidal as the central atom has five domains. The molecular geometry is also trigonal bipyramidal as there are no lone pairs.

Structure 2.2.14—Formal charge values can be calculated for each atom in a species and used to determine which of several possible Lewis formulas is preferred. Apply formal charge to determine a preferred Lewis formula from different Lewis formulas for a species.	
	Structure 3.1, Reactivity 3.2—What are the different assumptions made in the calculation of formal charge and of oxidation states for atoms in a species?

Strategy for Constructing Lewis Structures: Another Example with Carbonate Ion, CO_3^{2-}

1. How are the atoms connected?



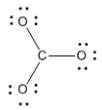
- four atoms could be connected multiple ways
- the possibility that comes closer to satisfying valences may be more likely

2. How many electrons are there?



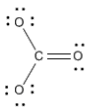
- C has 4 e⁻; **each** O has 6 e⁻; **2 extra e⁻ give the 2- charge**
- total of 24 e⁻
- used 2 e⁻ in each bond
- have 18 e⁻ left (9 pairs)

3. Fill in electrons



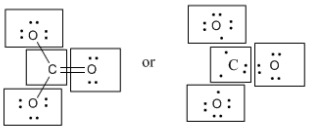
- used 10 e⁻
- have 0 e⁻ left
- have octet on each oxygen
- have no octet on carbon

4. Rearrange electrons to fill octets on each atom



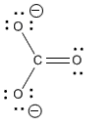
- octets complete
- this is a reasonable structure

5. Are there formal charges?



- looked at this way, two of the oxygens appear to be bringing **7 electrons** to the structure, **not 6**
- carbon brings 4e⁻, as expected
- the third oxygen brings 6 e⁻, as expect
- what about valences? Do you expect formal charge here?

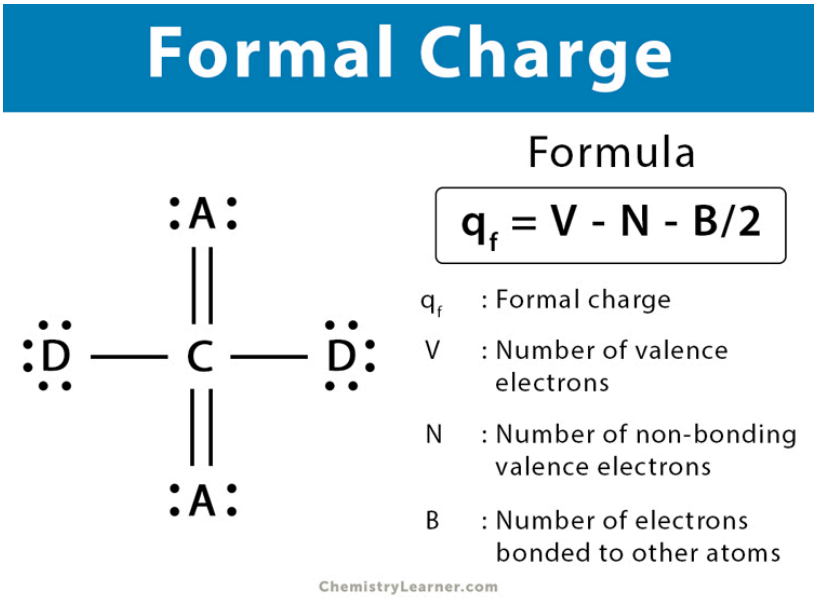
5. Draw in formal charge and circle it



-this is a **complete** Lewis structure

Go back and look at earlier examples of Lewis structures. Should there be formal charges on any of the atoms?

Compare the formal charges of different structures. The closer to 0 the better the structure is. Use this formula to find the formal charge:



Assumption:

- All shared bonding electrons (in covalent bonds) are split equally, neglecting electronegativity. (B divided by 2)
- All non-bonding electrons belong entirely to the atom they are attached to. (-N, entire charge is applied)

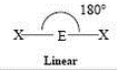
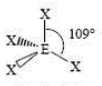
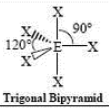
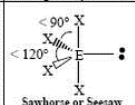
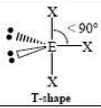
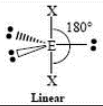
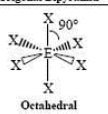
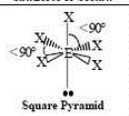
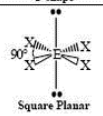
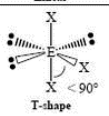
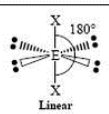
- All the individual formal charges equal the overall net charge of the molecule or ion. (represented by the summing in the equation)

Structure 2.2.15—Sigma bonds (σ) form by the head-on combination of atomic orbitals where the electron density is concentrated along the bond axis. **Pi bonds (π) form by the lateral combination of p-orbitals where the electron density is concentrated on opposite sides of the bond axis.** Deduce the presence of sigma bonds and pi bonds in molecules and ions.

Sigma bonds are the first bonds, the pi bonds are the additional ones atop the first one For example, CO_2 has two double bonds. This means that there are two sigma bonds and two pi bonds. Water on the other hand only has single bonds so it only has sigma bonds.

Structure 2.2.16—Hybridization is the concept of mixing atomic orbitals to form new hybrid orbitals for bonding. Analyse the hybridization and bond formation in molecules and ions. Identify the relationships between Lewis formulas, electron domains, molecular geometry and type of hybridization. Predict the geometry around an atom from its hybridization, and vice versa.	
Include both organic and inorganic examples. Only sp, sp^2 and sp^3 hybridization need to be covered.	

VSEPR Geometries

Steric No.	Basic Geometry 0 lone pair	1 lone pair	2 lone pairs	3 lone pairs	4 lone pairs
2	 Linear				
3	 Trigonal Planar	 Bent or Angular			
4	 Tetrahedral	 Trigonal Pyramid	 Bent or Angular		
5	 Trigonal Bipyramid	 Sawhorse or Seesaw	 T-shape	 Linear	
6	 Octahedral	 Square Pyramid	 Square Planar	 T-shape	 Linear

Domains include bonds and lone pairs. If a molecule has 4 domains it will do sp^3 (tetrahedral) hybridization, if it has 3 domains it will do sp^2(trigonal planar), two domains and it will do sp (linear). Hybridization enables bonding.