

Unit 2: Chemical Bonding and Organic Chemistry

Chapter 7: Quantum Theory and the Electronic Structure of Atoms

pg. 303 to 306

2. The units for wavelength are metre (m) and nanometre (nm). The units of frequency are reciprocal second (s^{-1}) and Hertz (Hz). The speed of light is 3.00×10^8 m/s or 1.86×10^5 miles/hr.
4. The range of wavelength for visible light is between 380 nm (violet) to 750 nm (red).
14. Photons are light particles that have no mass. Einstein's photoelectric effect indicates that light can behave like particles as well as a wave. When we describe reflection, refraction and interference patterns due to diffraction of light, it is necessary to view light as a wave. However, with instances like photoelectric effect, light should be viewed as a particle called photon. This is because these photons exhibit particle properties such as collisions and momentum necessary to knock a valence electron off a metal surface of a cathode ray tube to allow the completion of a circuit.
22. An energy level is a discrete quantized amount of energy an electron possess in an atom. According to Bohr and other quantum physicists, an electron can exist at a specific energy level and can never exist between these levels. When an electron gained a discrete amount of energy, it can "jump" to another higher energy level. We describe this electron being in an **excited state**. When an electron is at its lowest possible energy level, we refer to it as being in a **ground state**.
24. The negative sign in the equation $E_n = -R_H \left(\frac{1}{n^2}\right)$ signifies that as electron is dropped into various energy levels of an atom from the outside, it gives off energy to the surrounding (exothermic). This is because the electron is moving into a more stable ground state and achieving a lower energy level.
36. de Broglie hypothesized that if light can behave like wave and particles; perhaps electrons can behave like particles and waves as well. These waves are quantized because they are specific standing waves at specific distances from the nucleus as prescribed by the Bohr's energy levels. As electrons move from a higher energy level to a lower energy level, the number of nodes of these standing wave changes and the energy emits appear quantized.
38. All particles emit waves as they move. The wavelengths of these waves originating from particles (called particle waves) are largely determined by the mass and speed of the particles (what we commonly refer to as momentum). The formula, $\lambda = \frac{h}{m\nu}$ or $\lambda = \frac{h}{p}$ allows us to find the wavelength of particle wave. As we can see, the wavelength is very large for everyday objects with little mass and low speed. Hence, these particle waves are hard to detect.
43. Bohr's theory predicted the behaviors of a single electron accurately. But, it cannot account for the emission spectra of atoms containing more than one electron.
44. The Heisenberg Uncertainty Principle states that we cannot measure the momentum and the position of the particle precisely at the same time. This is because these quantities of a particle can be influence by the very act of simply measuring them. The Schrödinger Equation evaluate if a function is indeed a quantum standing wave function for any specific energy level. It also can determine the Eigen-values (Energy Values) of any qualifying wave functions.
45. The wave function is a mathematical equation that describes the subshells (atomic orbitals) of any specific energy level. Solving the wave function allows us to visualize the probability cloud (electron density) of the atomic orbital (subshell) it represents.
46. Electron density indicates the probability where an electron with a specific energy level and satisfy a specific wave function can exist in an atom. Since we cannot observe the momentum and the position of an electron with certainty (Uncertainty Principle), we can only approximate their whereabouts using an "electron probability cloud" which we call atomic orbital.

47. The atomic orbital is sometimes known as an energy subshell. It is the result of a specific wave function of an energy level. It is usually drawn as an electron density or probability cloud or a specific line on an energy diagram. An orbit is commonly used in the Bohr's model to represent the path electrons move around the nucleus and the maximum number of electrons for each orbit is governed by the number of atoms in each row of the Periodic Table of Elements.
48. The shape of an s -orbital is a sphere. The shape of a p -orbital is a dumbbell. The shape of a d -orbital is a four-leaf clover or a dumbbell with a donut in the middle. Using quantum numbers, n is the principle quantum number and it represents the energy level of an electron. ℓ is the angular momentum and it represents the type of atomic orbitals, such as s , p , d , and f subshells. m_ℓ is the magnetic angular momentum and it represents the orientation of these atomic orbitals (the p -orbital has three orientations, p_x , p_y and p_z , all pointing at different direction in a three-dimensional space).
50. According to the rules set out by quantum numbers, s -orbitals can exist in all principle quantum numbers and it has only one orientation (that of the sphere). p -orbitals can exist in principle quantum number, $n = 2$ and above and at each n , it can have three orientations. d -orbitals can exist in principle quantum number, $n = 3$ and above and at each n , it can have five orientations. The orbitals that cannot exist are $1p$, $2d$, $3f$ and $4g$.
52. n is the principle quantum number and it represents the energy level of an electron. ℓ is the angular momentum and it represents the type of atomic orbitals, such as s , p , d , and f subshells. m_ℓ is the magnetic angular momentum and it represents the orientation of these atomic orbitals (the p -orbital has three orientations, p_x , p_y and p_z , all pointing at different direction in a three-dimensional space). m_s is the magnetic spin quantum number and it represents the direction of a spin of an electron in a specific atomic orbital of a particular orientation.
54. n determines the energy of an electron in a hydrogen atom and in many-electron atom. It also determines the size of an orbital. ℓ determines the shape of an orbital. m_ℓ determines the orientation of an orbital in space.
72. $4d^6$ indicated that there are 6 electrons in the d -orbitals of the 4th energy level.
74. "Shielding of electrons" means that the nucleus (specifically the protons in the nucleus) is shielded from the valence electrons by the core electrons. For example, Li atom's nucleus is shielded from its one valence electron in the second energy level by its two core electrons in the first energy level. As the shielding effect of an atom decreases (like when the number of core electrons are the same and the valence electrons and the number of protons are increasing – as one moves across the period of the Table of Elements), the valence electrons can come in closer towards the nucleus. This causes a decrease in atomic radius, an increase in ionization energy and electron affinity.
75. (a) $(1, 0, \frac{1}{2}, \frac{1}{2})$; $m_\ell = \frac{1}{2}$ is not acceptable (m_ℓ has to be an integer).
(c) $(2, 2, 1, +\frac{1}{2})$; $0 < \ell < (n - 1)$ and since $n = 2$, ℓ can only be 0 and 1.
(e) $(3, 2, 1, 1)$; $m_s = 1$ is not acceptable ($m_s = \pm\frac{1}{2}$).
82. Osmium is in 6th period and group 8 of the Table of Elements.
84. Cr has a ground-state electron configuration of $[\text{Ar}] 4s^1 3d^5$ because the $4s$ and $3d$ orbitals have very close energy levels. In order to maximize Hund's rule of unpaired electrons, an electron that was in a $4s^2$ configuration moves into an empty $3d$ orbital, creating 6 unpaired electrons arrangement instead of 4 unpaired electrons. In the case of Cu, its ground-state electron configuration of $[\text{Ar}] 4s^1 3d^{10}$ is more favourable than $[\text{Ar}] 4s^2 3d^9$. This is because at this point of filling electrons in the atomic orbitals, the $4s$ orbital has a slightly higher energy than the $3d$ orbital. Since we have to fill in the lower energy orbital first, one of the $4s^2$ electron migrate to the $3d^9$ orbitals to complete them to $3d^{10}$.
86. The statement is correct. Even if two electrons exist in the same orientation of the same atomic orbital of a specific energy level, they will be of different spin direction making them never to be assigned the same four quantum numbers.

91. There are a total of twelve electrons:
The element is magnesium (Mg)

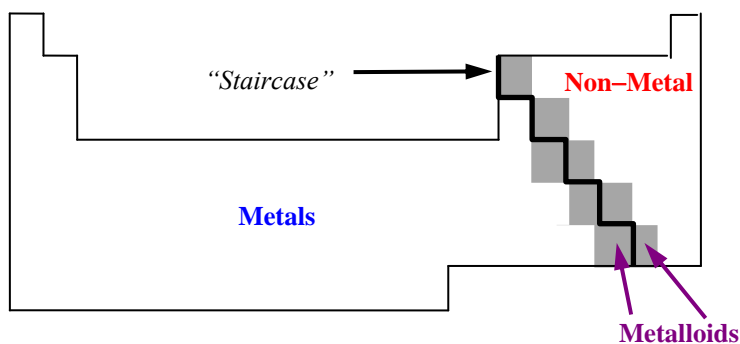
Orbital	n	ℓ	m_ℓ	m_s
1s	1	0	0	+1/2
1s	1	0	0	-1/2
2s	2	0	0	+1/2
2s	2	0	0	-1/2
2p	2	1	1	+1/2
2p	2	1	1	-1/2
2p	2	1	0	+1/2
2p	2	1	0	-1/2
2p	2	1	-1	+1/2
2p	2	1	-1	-1/2
3s	3	0	0	+1/2
3s	3	0	0	-1/2

Chapter 8: Periodic Relationships Among the Elements

pg. 348 to 350

- Mendeleev's Periodic Table is significant because it organizes the elements based on their chemical properties and it is possible to make predictions on elements that had yet been discovered.
- Moseley discovered a relationship between atomic number and the frequency of X-rays generated by bombarding an element with high energy electrons.
- The modern layout of the Periodic Table of Elements not only retains the spirit of Mendeleev's work of grouping the elements with similar chemical properties in columns, it also blocked the elements by their atomic orbitals (s - and p -orbitals into the representative elements – A groups, d -orbitals into the transition metals – B groups and the inner-transition metals of the f -orbitals as outlined by the lanthanides and actinides series at the bottom of the table).
- The most important relationship of the elements in the same group of the table is their similar chemical properties as they have relatively the same electron configuration at the valence energy level.

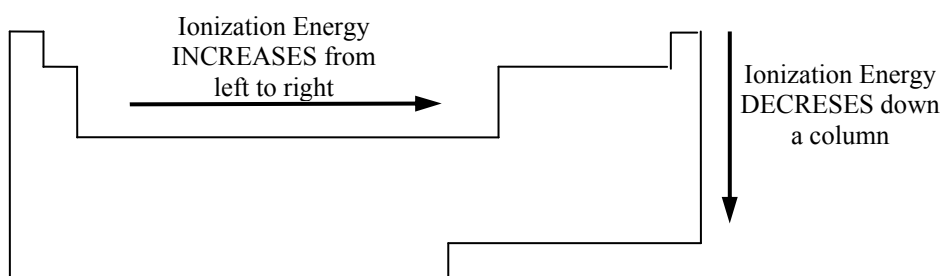
7.



- We can test whether the substance conducts electricity. As iodine is a non-metal, it would not conduct any electricity.
- (a) Alkali Metals ns^1 (b) Alkaline Earths ns^2 (c) Halogens ns^2np^5 (d) Noble Gases ns^2np^6
- The electron configuration for representative metal ions is the same as their previous noble gas. While for representative non-metal ions is the same as the subsequent noble gases. This is due to the fact that noble gases are the most stable element of each period because they have the least shielding effects and highest Z_{eff} .
- Isoelectronic means that ions or an atom having the same electron configuration.
- No two atoms can have the same electron configuration.

18. Using Figure 4.11 on pg. 136, the first-row transition metal ions that have the same electron configuration as the argon core are Sc^{3+} , Ti^{4+} , V^{5+} , Cr^{6+} and Mn^{7+} .
19. Hydrogen forms the H^+ ion (resembles the alkali metals) and the H^- ion (resembles the halogens).
23. (a) $1s^2 2s^2 2p^6 3s^2 3p^5$ (b) halogen
34. (a) Atomic radius generally decreases from left to right across the period.
(b) Atomic radius increases top to bottom in a group.
36. For isoelectronic ions, the cations tend to have more protons in their nuclei than their isoelectronic anions while they all have the same number of electrons. The higher number of protons allows these cations to pull the electrons more effectively than the anions. Hence, isoelectronic anions are larger than cations.
43. (a) Cl is smaller than Cl^- . An atom gets bigger when more electrons are added.
(b) Na^+ is smaller than Na . An atom gets smaller when electrons are removed.
(c) O^{2-} is smaller than S^{2-} . Both elements belong to the same group, and ionic radius increases going down a group.
(d) Al^{3+} is smaller than Mg^{2+} . The two ions are and in such cases the radius gets smaller as the charge becomes more positive (more protons in the nucleus of Al^{3+}).
(e) Au^{3+} is smaller than Au^+ for the same reason as part (b). Besides they both have the same number of protons and the pull of the electrons is stronger on Au^{3+} has two less electrons.
45. The Cu^+ ion is larger than Cu^{2+} because it has one more electron. They both have the same number of protons and the pull of the electrons is stronger on Cu^{2+} has one less electron.
47. Bromine is liquid; all the others are solids.
49. Ionization energy is the minimal amount of energy needed to completely remove a valence electron of a gaseous atom from its ground state. Gaseous atoms are specified in the definition because atoms in gas phase do not have any intermolecular interactions with adjacent atoms. Hence, the amount of energy measured for the ionization process is more accurate. The Second Ionization Energy is always greater than the First Ionization Energy because the number of protons remain the same at each stages of ionization. Therefore, as more electrons are removed, the difference between the electric charges increases and more energy are required to remove any successive electrons.

50.



51. Ionization energy increases across a row of the periodic table and decreases down a column or group. The correct order of increasing ionization energy is $\text{Cs} < \text{Na} < \text{Al} < \text{S} < \text{Cl}$.
59. (a) Electron Affinity is the negative of the energy change that occurs when an electron is accepted by a gaseous atom.
(b) Electron affinity is made with gaseous atom because like ionization energy, we want to minimize effect of intermolecular forces when measuring electron affinity.
(c) Ionization is always a positive energy because electrons absorb energy to move to higher orbitals. Since ionization is an act of remove an electron of an atom from its ground state, it will always be associated with positive energy. Electron affinity is a process when an atom accepts an electron. This is the reverse of ionization, and hence energy is usually given off in this process. Therefore, electron affinity is usually associated with negative energy. Its energy can be positive in cases when placing an electron in a filled energy level – such as putting an extra electron into the valence energy level of a noble gas atom.

60. There is a generally increasing trend for electron affinity from aluminum to chlorine, with the exception of Si. This is because Si has 2 unpaired electrons in its $3p$ orbital and an extra electron would be an arrangement with minimum energy according to Hund's rule. Hence, Si has a higher electron affinity than P despite the general increasing trend.
61. (a) $K < Na < Li$
 (b) $I < Br < F < Cl$ (F has only $2p$ electrons and it is too small and hence electron repulsion is more pronounced than Cl.)
63. Based on electron affinity values, we would not expect the alkali metals to form anions. A few years ago most chemists would have answered this question with a loud "No"! In the early seventies a chemist named J.L. Dye at Michigan State University discovered that under very special circumstances alkali metals could be coaxed into accepting an electron to form negative ions! These ions are called alkalide ions.

Chapter 9: Chemical Bonding I: Basic Concepts

pg. 390 to 393

1. Lewis dot symbol shows the valence electrons of the s and p orbitals. Lewis dot symbol is effective when it applies to representative (A) group elements.
5. (a) $\cdot\ddot{\text{I}}\cdot$ (b) $[\cdot\ddot{\text{I}}\cdot]^-$ (c) $\cdot\ddot{\text{S}}\cdot$ (d) $[\cdot\ddot{\text{S}}\cdot]^{2-}$ (e) $\cdot\ddot{\text{P}}\cdot$ (f) $[\cdot\ddot{\text{P}}\cdot]^{3-}$ (g) $\text{Na}\cdot$
 (h) $[\text{Na}]^+$ (i) $\cdot\text{Mg}\cdot$ (j) $[\text{Mg}]^{2+}$ (k) $\cdot\overset{\cdot}{\text{Al}}\cdot$ (l) $[\text{Al}]^{3+}$ (m) $\cdot\text{Pb}\cdot$ (n) $[\text{Pb}]^{2+}$
7. When atoms with small ionization energy (especially groups 1A and 2A metals) combine with atoms that have large electron affinity (especially groups 6A and 7A nonmetals), they will form ionic compound. This is because it would be easier for the electron to leave the metal atom and join the nonmetal atom where large energy is given off. Hence, it would achieve a more favorable energy state.
9. Any ionic compounds with ammonia as cations. Example: NH_4Cl
10. Answers may vary. Examples: NH_4OH , $(\text{NH}_4)_2\text{CO}_3$
13. (a) $\text{NaCl}_{(s)}$ would not conduct electricity because its ions are locked in a lattice structure.
 (b) $\text{NaCl}_{(l)}$ would conduct electricity because the lattice structure is broken down and the ions are free to move about.
 (c) $\text{NaCl}_{(aq)}$ would conduct electricity because water molecules would carry the ions to move around in the solution.
15. We use Coulomb's law to answer this question: $E_{\text{lattice}} = \frac{kQ_{\text{cation}}Q_{\text{anion}}}{r}$
 (a) Doubling the radius of the cation would increase the distance, r , between the centers of the ions. A larger value of r results in a smaller energy, E , of the ionic bond.
 (b) Tripling the charge on the cation will result in tripling of the energy, E , of the ionic bond, since the energy of the bond is directly proportional to the charge on the cation, Q_{cation} .
 (c) Doubling the charge on both the cation and anion will result in quadrupling the energy, E , of the ionic bond.
 (d) Decreasing the radius of both the cation and the anion to half of their original values is the same as halving the distance, r , between the centers of the ions. Halving the distance results in doubling the energy.
21. Lattice energy is the energy required to completely separate one mole of solid ionic compound into gaseous state. The greater the lattice energy, the more stable is the ionic compound.

22. Base on the conservation of energy, we can use the Born-Haber process to determine the lattice energy of KCl.

$$\Delta H_f \text{ of KCl} = \Delta E_{\text{sub}} + \frac{1}{2}\Delta E_{\text{bond}} \text{ of Cl}_2 + I_1 + E_{\text{affinity}} + E_{\text{lattice}}$$

ΔH_f = Molar Heat of Formation of KCl

ΔE_{sub} = Heat of Sublimation of K atoms (endothermic)

ΔE_{bond} = Molar Bond Energy of Cl₂ (endothermic)

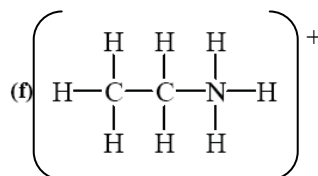
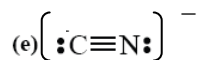
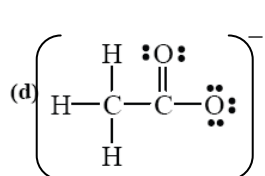
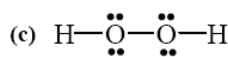
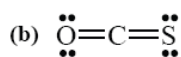
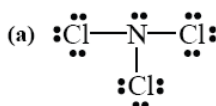
I_1 = First Ionization Energy of K atoms (endothermic)

E_{affinity} = Electron Affinity of Cl atoms (exothermic)

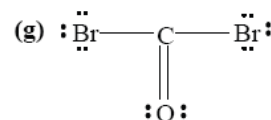
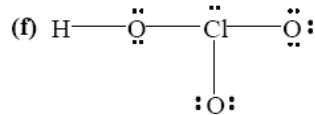
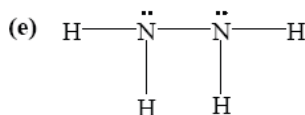
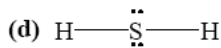
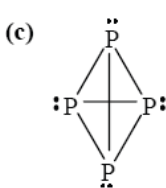
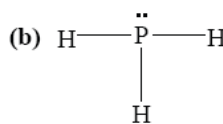
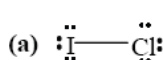
E_{lattice} = Lattice Energy of KCl (exothermic)

24. (a) LiF is more stable than LiF₂ because LiF₂ (which contains Li²⁺ ion) would require a significant higher ionization energy overall ($I_1 + I_2$ vs. just I_1).
25. 787 kJ/mol
30. For HBr, bromine has 3 lone pairs. For H₂S, sulfur has two lone pairs. For CH₄, carbon has no lone pairs.
31. Single bond has two electrons between the atoms of the covalent bond, as in the case of CH₄. Double bond has four electrons between the atoms of the covalent bond, as in the case of CO₂. Triple bond has six electrons between the atoms of the covalent bond, as in the case of CO. Bond length decreases with increasing multiple bonds. Hence, triple bond is shorter than double bond, and double bond is shorter than single bond.
32. Ionic compounds have crystalline lattice structure, but covalent compound have no lattice structure in its crystal. When ionic compounds dissolve in water, it can conduct electricity because their ions are now free to move. When molecular compounds dissolve in water, there are no ions and the water molecule and the solute molecules are attracted because of their polarity. Hence, no electricity can be conducted in a solution of molecular compounds. Since lattice energy is a stronger bond than energy in a covalent bond, ionic compounds tend to be solids in room temperature and they have very high melting points. Molecular compounds can be solids, liquids or gases in room temperature as they tend to have relative low melting points.
34. A polar covalent bond is a covalent bond with two atoms of different electronegativity. This can cause electrons to be closer to one atom instead of the other, creating a non-uniform electron density within the covalent bond. As such, one side of the covalent bond is more negatively charged than the other side.
37. DG < EG < DF < DE
39. (a) The two carbon atoms are the same. The bond is covalent.
 (b) The electronegativity difference between K and I is 2.5 - 0.8 = 1.7. The bond is polar covalent.
 (c) The electronegativity difference between N and B is 3.0 - 2.0 = 1.0. The bond is polar covalent.
 (d) The electronegativity difference between C and F is 4.0 - 2.5 = 1.5. The bond is polar covalent.
42. The formal charge of an atom is the electrical charge difference between the valence electrons in an isolated atom and the number of electron assigned to that atom in a Lewis structure.
 Formal Charge = # of valence electron - # of lone pair electrons - 1/2 (# of bonding electrons)

43.



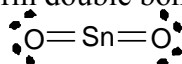
45.



47.

(a) The structure has too many electrons. The correct structure is $\text{H}-\text{C}\equiv\text{N}$:(b) Hydrogen atoms do not form double bond. The correct structure is $\text{H}-\text{C}\equiv\text{C}-\text{H}$

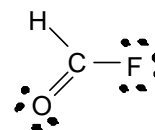
(c) Too few electrons.



The correct structure is

(f) Oxygen does not have an octet.

The correct structure is



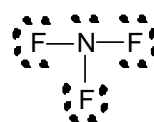
(d) Too many electrons.

The correct structure is



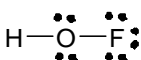
(g) Too few electrons.

The correct structure is



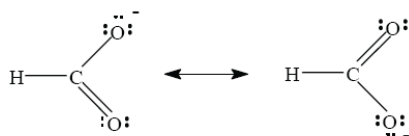
(e) Fluorine has more than an octet.

The correct structure is

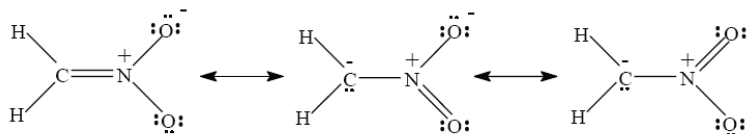


51. The resonance structures are:

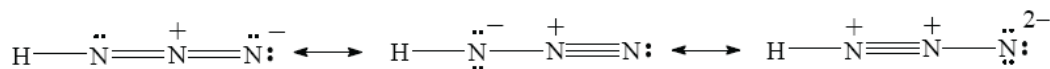
(a)



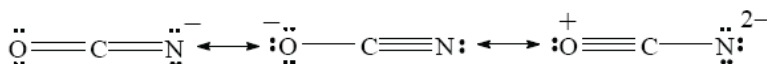
(b)



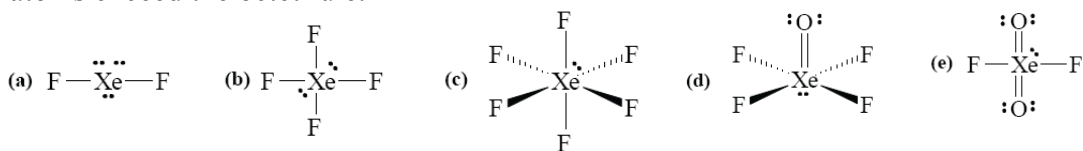
53. The structures of the most important resonance forms are:



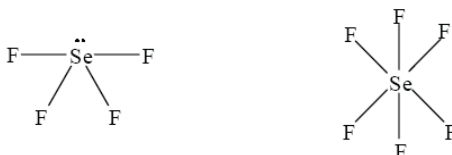
55. Three reasonable resonance structures are:

57. Atoms in the third period and beyond tend to exceed the octet rule. This is because the adjacent d -orbitals are available for borrowing starting at $n = 3$. Hence, these central atoms can have dsp^3 and d^2sp^3 hybridizations (sometimes even d^3sp^3).59. F can only form a single bond. In order to exceed the octet rule, the atom must be able to borrow from any adjacent d -orbitals. Since fluorine has an electron configuration of $[\text{He}]2s^22p^5$, and d -orbitals start appearing in $n = 3$, it cannot exceed the octet rule. Hence, compounds like FH_7 and FCl_7 (both would mean F follow the 14-electron rule) cannot be synthesized.

63. For simplicity, the three lone-pairs around the fluorine atoms are omitted. Note that all central atoms exceed the octet rule.



65. For simplicity, the three lone-pairs around the fluorine atoms are omitted. In each case the octet rule did not satisfy for Se.



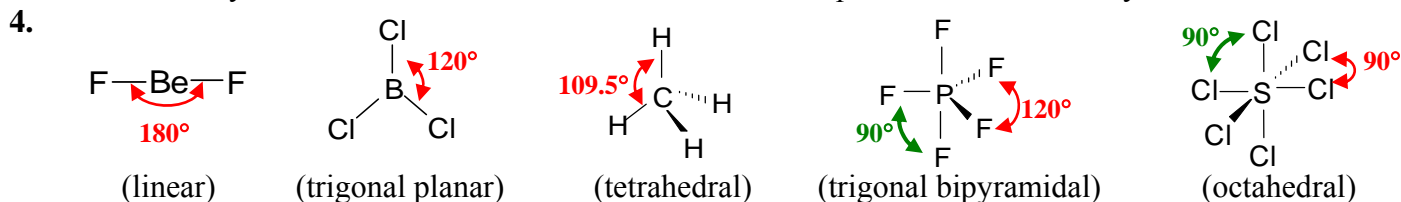
69. ΔH for average (N-H) bond = 392 kJ/mol

71. ΔH_f for $\text{F}_{(g)}$ = 78.5 kJ/mol

Chapter 10: Chemical Bonding II: Molecular Geometry & Hybridization of Atomic Orbitals

pg. 442 to 443

2. Lewis dot symbol shows the valence electrons of the *s* and *p* orbitals. Lewis dot symbol is effective



5. If there are any lone-pairs on a trigonal bipyramidal central atom, it would occupy the equatorial position because the angle there is bigger than the axial bond angle (120° compares to 90°).

7. (a) PCl_3 (trigonal pyramid) (b) CHCl_3 (tetrahedral)

- (c) SiH_4 (tetrahedral) (d) TeCl_4 (see-saw)

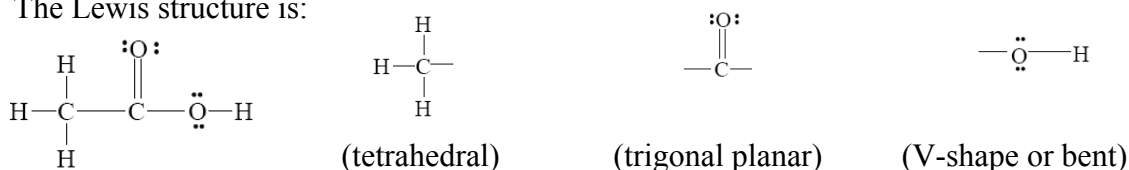
9. (a) CBr_4 (tetrahedral) (b) BCl_3 (trigonal planar)

- (c) NF_3 (trigonal pyramid) (d) H_2Se (bent or V-shape) (e) NO_2^- (bent or V-shape)

11. (a) The lone-pairs on Br atoms have been omitted for simplicity. $\text{Br}-\text{Hg}-\text{Br}$ (HgBr_2 is linear).

- (b) N_2O is linear. $:\text{N}\equiv\text{N}^+-\ddot{\text{O}}^-$ (c) SCN^- is linear. $:\ddot{\text{S}}-\text{C}\equiv\text{N}^-$

13. The Lewis structure is:



16. Dipole Moment and Bond Moment are essentially the same thing. It is the relative difference between the electronegativities of the two atoms that form a chemical bond. The greater this difference, the more polar is the bond. It is necessary to know that even though a molecule has many dipole moments, it can still be a non-polar molecule. This is true when all the dipole moment vectors cancel each other out when added.

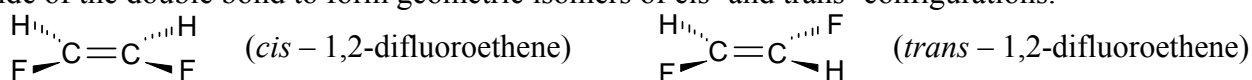
17. It takes two atoms bonding together to form a dipole. A single atom cannot have a permanent dipole moment. (It can, however, have a temporary dipole moment because electrons can momentarily gathered on one side of the atom, making that side temporary more negative. It is called London-Dispersion Force – more of this in Unit 3.)

21. $\text{CO}_2 = \text{CBr}_4$ (non-polar for both) < H_2S < NH_3 < H_2O < HF

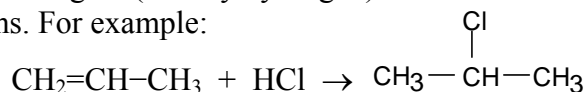
Chapter 24: Organic Chemistry

pg. 1029 to 1031

- Carbon can form more compounds than any other element because carbon atoms are able not only to form single, double, and triple carbon-carbon bonds, but also to link up with each other in chains and ring structures
- Aliphatic hydrocarbons do not contain the benzene group, or the benzene ring, whereas aromatic hydrocarbons contain one or more benzene rings.
- Saturated hydrocarbons are hydrocarbon molecules that consist of single covalent bonds. An example of saturated hydrocarbons is propane, which belongs to a class of hydrocarbons called alkanes, C_nH_{2n+2} . Unsaturated hydrocarbons are hydrocarbon molecules that consist of one or more multiple bonds. An example of unsaturated hydrocarbon is ethane, C_2H_4 , which belongs to a class of hydrocarbons called alkenes, C_nH_{2n} .
- Methane is a natural product of the anaerobic bacterial decomposition of vegetable matter under water, such as marshes. It can also be produced by termites as they digest they cellulose in wood. Some methane is produced in sewage treatment processes. Commercially, methane is obtained from natural gas.
- For C=C bond, the pi-bond offer by the $2p_z$ orbitals do not allow the double bond to rotate freely as single C-C bond, which lack the pi-bond restriction. This allows the special substituents of either side of the double bond to form geometric isomers of cis- and trans- configurations.



- Alkanes consist of only single bonds that can rotate freely. Hence, their special substituents on either side of the adjacent carbon atoms cannot have permanent cis- or trans- configurations. Alkynes, on the other hand, have a triple bond. This means that it can allow only one substituent on either side of the triple bonded carbons. Since it is necessary to have two substituents on either side of these adjacent carbons to have cis- or trans- arrangements, alkynes can never have geometric isomers.
- Markovnikov's rule states that in the addition of unsymmetrical (that is, polar) reagents to alkenes, the positive portion if the reagent (usually hydrogen) adds to the carbon atom that has already has the most hydrogen atoms. For example:



- Hydrocarbon Combustion Reactions: Alkane + Oxygen \rightarrow Carbon Dioxide + Water Vapour

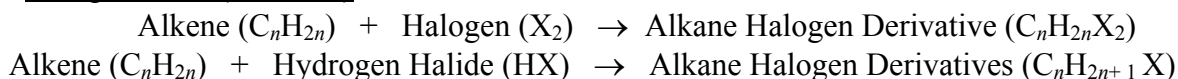
Reactions with Alkanes:

- Substitution: Alkane + Halogen (X_2) $\xrightarrow{h\nu}$ Halogen Derivate + HX
- Dehydrogenation: Alkane $\xrightarrow{\text{Pt}}$ Alkene (double bond) + Hydrogen

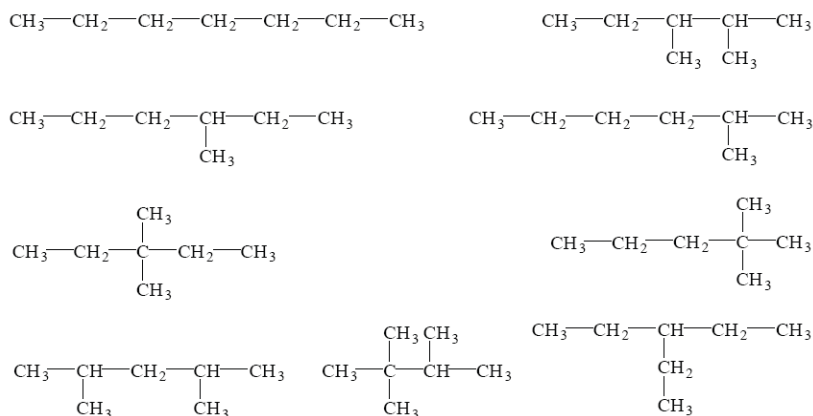
Reactions with Alkenes and Alkynes:

- Hydrogenation: Alkyne + Hydrogen $\xrightarrow{\text{catalyst}}$ Alkene
Alkene + Hydrogen $\xrightarrow{\text{catalyst}}$ Alkane

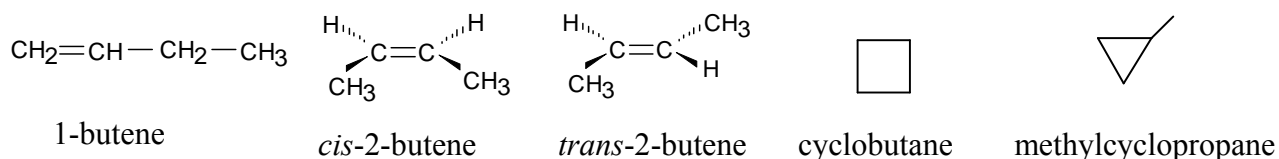
- Halogenations (Addition):



11. The structural isomers of C_7H_{16} are

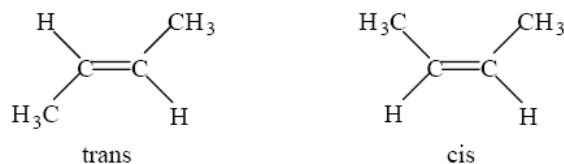


13. The structural isomers of C_4H_8 are



15. The straight chain molecules have the highest boiling points and therefore the strongest intermolecular attractions. These chains can pack together more closely and efficiently than highly branched, cluster structures. This allows intermolecular forces to operate more effectively and cause stronger attractions. The stronger the intermolecular force, the more energy is needed to pull the molecules from each other when they change from liquid phase to gaseous phase. Hence, the higher the boiling points for these compounds.

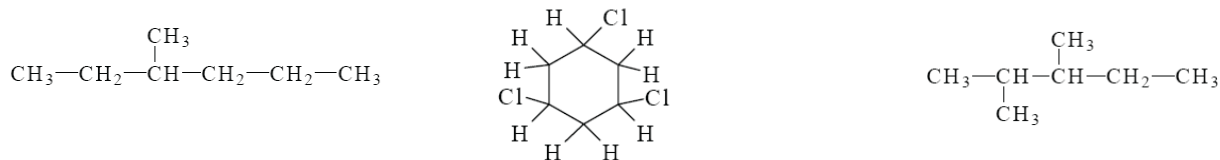
17. The two isomers are



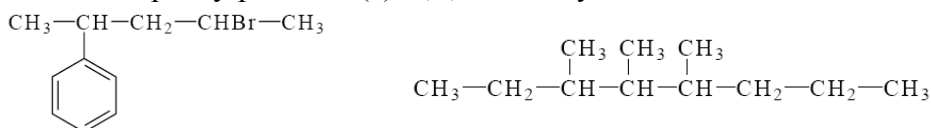
The *cis*- structure is more crowded and a little less stable. As a result, slightly more heat (energy) would be released when the alkene adds a molecule of hydrogen to form butane, C_4H_{10} . Note that butane is the product when either alkene is hydrogenated.

23. (a) $\text{CH}_2=\text{CH}-\text{CH}_2-\text{CH}_3 + \text{HBr} \rightarrow \text{CH}_3-\text{CHBr}-\text{CH}_2-\text{CH}_3$
 (b) $\text{CH}_3-\text{CH}=\text{CH}-\text{CH}_3 + \text{HBr} \rightarrow \text{CH}_3-\text{CH}_2-\text{CHBr}-\text{CH}_3$ (a) and (b) have the same product.

27. (a) 3-methylhexane (b) 1,3,5-trichlorocyclohexane (c) 2,3-dimethylpentane

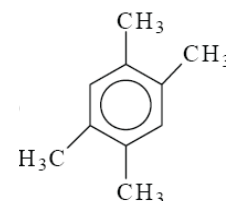
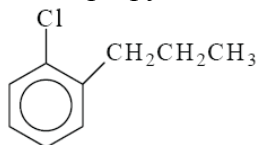
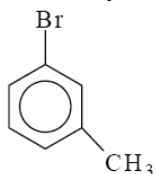


(d) 2-bromo-4-phenylpentane (e) 3,4,5-trimethyloctane

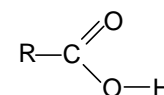
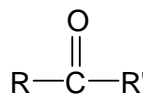
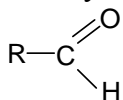
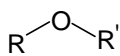
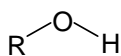


29. Benzene is relatively stable due to the electron delocalization of the alternate double bonds in the ring structure. Because it is difficult to break the double bonds in benzene, the aromatic compound generally undergo substitute reactions instead of addition reactions. On the other hand, ethylene (ethene) does not have the alternate cyclic double bond arrangement. Therefore, its double bond tends to allow addition reactions with other elements or compounds.
30. Benzene is planar because each carbon of the ring has a sp^2 hybridization, which makes each of their bond angles 120° . Since the interior angles of a regular hexagon is also 120° , benzene ring is a planar structure. Cyclohexane is non-planar because all the carbons has a sp^3 hybridization, which makes each of their bond angles 109.5° . As this does not fit the interior angles of a regular hexagon, cyclohexane is non-planar.

31. (a) 1-bromo-3-methylbenzene (*m*-bromomethylbenzene) (b) 1-chloro-2-propylbenzene (*o*-chloropropylbenzene) (c) 1,2,4,5-tetramethylbenzene

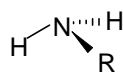
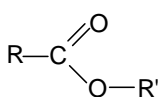


34. Alcohol Ether Aldehyde Ketone Carboxylic Acid

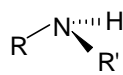


Ester

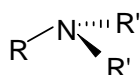
Amine



or



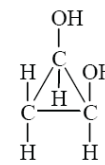
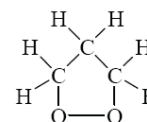
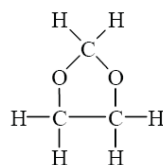
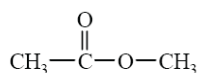
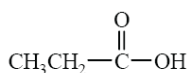
or



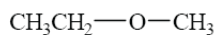
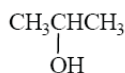
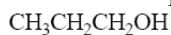
35. (a) There is only one isomer: CH_3OH

(b) There are two structures with this molecular formula: CH_3-CH_2-OH and CH_3-O-CH_3

(c) The cyclic di-alcohol has geometric isomers.

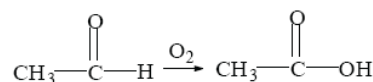


(d) There are two possible alcohols and one ether.



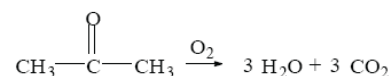
37. Aldehydes can be oxidized easily to carboxylic acids.

The oxidation reaction is:

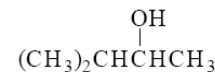
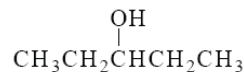
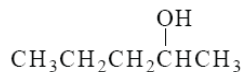


Oxidation of a ketone requires that the carbon chain be broken.

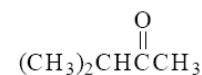
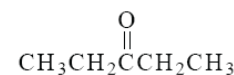
Essentially, it becomes a hydrocarbon combustion.



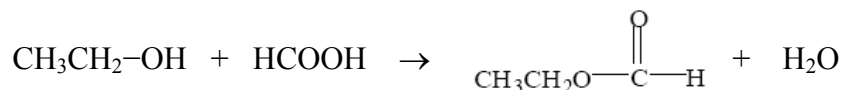
39. Secondary alcohols can be oxidized to ketones under controlled conditions. The possible starting compounds are:



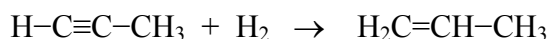
The corresponding products are:



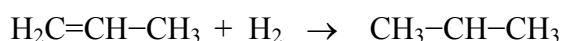
41. (a) Esterification of ethanol and methanoic acid produces ethyl methanoate and water



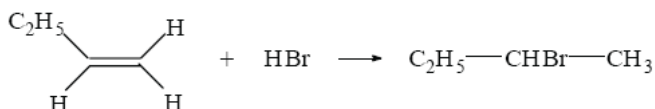
- (b) Addition of equal mole of hydrogen to an equivalent mole alkyne gives an alkene.



With another equal mole of hydrogen, the alkene can turn into alkane.



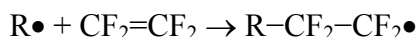
- (c) HBr will add to the alkene as shown. (Note: the carbon atoms at the double bond have been omitted for simplicity.)



Chapter 25: Synthetic Organic Polymers

pg. 1058 to 1059

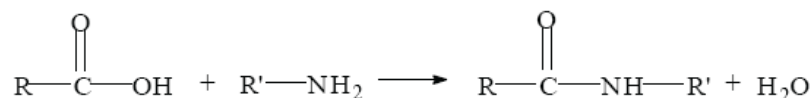
3. $(-\text{CH}_2-\text{CH}_2-)_n = n(28.06 \text{ g/mol})$ For $n = 4600$, molar mass = $1.291 \times 10^3 \text{ g/mol} = 129.1 \text{ kg/mol}$
4. Synthetic organic polymers can be made via addition reactions across monomers with double bonds, where same monomers can add onto themselves (homopolymer), or different monomers can be added (copolymer). Synthetic polymers can also be made by condensation reaction (esterification) between alcohol containing monomer with a carboxylic acid containing monomer. The resulting material is a polyester class of polymer.
7. The reaction is initiated by a radical $\text{R}\bullet$,



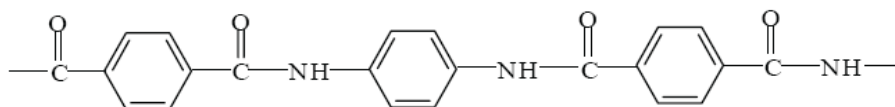
The product is also a radical, and the reaction continues.



9. The general reaction is a condensation to form an amide (carboxylic acid + amine \rightarrow amide + water).



The polymer chain looks like:



Note that both reactants are distributed benzene derivatives with the substituent in the para- or 1,4-positions.

11. The structures of the monomers are as shown.

