

Chapter 1

Problem 1P

How many electrons does carbon have? How many are valence electrons? What third-row element has the same number of valence electrons as carbon?

Step-by-step solution

step 1 of 1

Carbon belongs to 4A group. Its atomic number is 6. So, the carbon has 6 electrons. Its electronic configuration is $1s^2 2s^2 2p^2$. In the carbon atom, the outer most orbitals (valence orbitals) are $2s^2 2p^2$. These are second-row orbitals. Therefore, the valence electrons present in outer most orbitals ($2s^2 2p^2$) are 4.

The electronic configurations for 4A group elements are as shown below:

Element	Atomic number	Electronic configuration
C	6	$[\text{He}] 2s^2 2p^2$
Si	14	$[\text{Ne}] 3s^2 3p^2$
Ge	32	$[\text{Ar}] 3d^{10} 4s^2 4p^2$
Sn	50	$[\text{Kr}] 4d^{10} 5s^2 5p^2$
Pb	82	$[\text{Xe}] 4f^{14} 5d^{10} 6s^2 6p^2$

From the above table, the third-row of silicon (Si) has the same number of valence electrons ($3s^2 3p^2$) as carbon ($2s^2 2p^2$).

Problem 2P

Referring to the periodic table as needed, write electron configurations for all the elements in the third period.

Step-by-step solution

step 1 of 1

The third period begins with sodium (Na) and ends with argon (Ar).

The atomic number (Z) of sodium is 11, so a sodium atom has 11 electrons. The electron configuration of sodium is $1s^2 2s^2 2p^6 3s^1$.

Sodium (Na) - $1s^2 2s^2 2p_x^2 2p_y^2 2p_z^2 3s^1$

The electron configurations of all other elements of the third period are given below.

Atomic Number	Element	Electronic configuration
12	Magnesium (Mg)	$1s^2 2s^2 2p^6 3s^2$
13	Aluminum (Al)	$1s^2 2s^2 2p^6 3s^2 3p_x^1$
14	Silicon (Si)	$1s^2 2s^2 2p^6 3s^2 3p_x^1 3p_y^1$ $1s^2 2s^2 2p^6 3s^2 3p_x^1 3p_y^1$
15	Phosphorus (P)	$1s^2 2s^2 2p^6 3s^2 3p_x^1 3p_y^1 3p_z^1$
16	Sulfur (S)	$1s^2 2s^2 2p^6 3s^2 3p_x^2 3p_y^1 3p_z^1$
17	Chlorine (Cl)	$1s^2 2s^2 2p^6 3s^2 3p_x^2 3p_y^2 3p_z^1$
18	Argon (Ar)	$1s^2 2s^2 2p^6 3s^2 3p_x^2 3p_y^2 3p_z^2$

Problem 3P

Species that have the same number of electrons are described as isoelectronic. What + 2 ion is isoelectronic with Na+? What -2 ion?

Step-by-step solution

step 1 of 1

Species that have the same numbers of electrons are described as isoelectronic.

The atomic number of sodium (Na) is 11, and a sodium atom has 11 electrons. The Na^+ ion has 10 electrons.

Thus, a +2 ion which is isoelectronic with Na^+ has 10 electrons.

Again, the atomic number of that +2 ion is $(10+2)=12$. So it is a magnesium ion (Mg^{2+}) which is isoelectronic with Na^+ .

Ion	Electron configuration
Na^+	$1s^2 2s^2 2p^6$
Mg^{2+}	$1s^2 2s^2 2p^6$

A -2 ion which is isoelectronic with Na^+ also contains 10 electrons. The atomic number of that -2 ion is $(10-2)=8$.

Thus, it is oxygen. The electron configuration of both Na^+ and O^{2-} are $1s^2 2s^2 2p^6$.

So Na^+ and O^{2-} are isoelectronic.

Note:

Element	Number of electrons	Electronic configuration
Na	11	$1s^2 2s^2 2p^6 3s^1$
Na^+	$(11-1)=10$	$1s^2 2s^2 2p^6$
Mg	12	$1s^2 2s^2 2p^6 3s^2$
Mg^{2+}	$(12-2)=10$	$1s^2 2s^2 2p^6$

Problem 4P

Which of the following ions possess a noble gas electron configuration? Which ions are isoelectronic?

(a) K^-

(b) He^+

(c) H^-

(d) 0

(e) F^-

(f) Ca^{2+}

Step-by-step solution

step 1 of 6

(a)

The atomic number of potassium is 19. It possesses 19 electrons in its atom. The K^+ ion has one electron deficient than its atom. Therefore, the K^+ ion has 18 electrons in it which is the same as the number of the electrons in Argon, noble gas.

The electronic configuration of K^+ ion and Argon are the same - $1s^2 2s^2 2p^6 3s^2 3p^6$.

Therefore, K^+ ion possesses the noble gas configuration.

step 2 of 6

(b)

The atomic number of Helium is 2. It possesses 2 electrons in its atom. The He^+ ion has one electron deficient than its atom. Therefore, the He^+ ion has 1 electron in it. The electronic configuration of He^+ ion is - $1s^1$.

Therefore, He^+ ion does not possess the noble gas configuration.

step 3 of 6

(c)

The atomic number of Hydrogen is 1. It possesses 1 electron in its atom. The H^- ion has one electron more than its atom. Therefore, the H^- ion has 2 electrons in it which is the same as the number of electrons in Helium, noble gas.

The electronic configuration of H^- ion and Helium are the same - $1s^2$.

Therefore, H^- ion possesses the noble gas configuration.

step 4 of 6

(d)

The atomic number of Oxygen is 8. It possesses 8 electrons in its atom. The O^- ion has one electron more than its atom. Therefore, the O^- ion has 9 electrons in it. The electronic configuration of O^- ion is - $1s^2 2s^2 2p^5$.

Therefore, O^- ion does not possess the noble gas configuration.

step 5 of 6

(e)

The atomic number of Fluorine is 9. It possesses 9 electrons in its atom. The F^- ion has one electron more than its atom. Therefore, the F^- ion has 10 electrons in it which is the same as the number of electrons in Neon, noble gas.

The electronic configuration of F^- ion and Ne are the same - $1s^2 2s^2 2p^6$.

Therefore, F^- ion possesses the noble gas configuration.

step 6 of 6

(f)

The atomic number of Calcium is 20. It possesses 20 electrons in its atom. The Ca^{2+} ion has two electrons deficient than its atom. Therefore, the Ca^{2+} ion has 18 electrons in it which is the same as the number of electrons in Argon, noble gas.

The electronic configuration of Ca^{2+} ion and Argon are the same - $1s^2 2s^2 2p^6 3s^2 3p^6$.

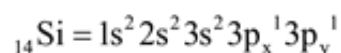
Therefore, Ca^{2+} ion possesses the noble gas configuration.

There are six electrons in carbon.

The electron configuration of carbon is $1s^2 2s^2 2p_x^1 2p_y^1$

There are four valence electrons $(2s^2 2p_x^1 2p_y^1)$ in carbon.

Silicon (Si, atomic number 14) is the third row element which has the same number of valence electrons as carbon.



The four valence electrons are $3s^2, 3p_x^1$ and $3p_y^1$.

[Key to solve: The third row element which has the same number of valence

electrons must have the valence electrons in the configuration $3s^2, 3p_x^1$ and $3p_y^1$.

The maximum number of inner orbital electrons in a $2s, 2s^1$, and $2p$ orbital is ten.

Thus, the atomic number of the element is $(10+4) = 14$. So it is silicon.]

Problem 5P

What is the electron, configuration of C^+ ? Of C^- ? Does either one of these ions have a noble gas (closed-shell) electron configuration?

Step-by-step solution

step 1 of 1

Carbon has atomic number 6, and so a carbon atom has six electrons.

Element	Number of electrons	Electronic configuration
C^+	$(6-1) = 5$	$1s^2 2s^2 2p^1$
C^-	$(6+1) = 7$	$1s^2 2s^2 2p^3$ or $1s^2 2s^2 2p_x^1 2p_y^1 2p_z^1$

The noble gases are characterized by an extremely stable “closed-shell” electron configuration. However, neither C^+ nor C^- possess such a type of “closed-shell” electron configuration.

Thus, C^+ and C^- do not have noble gas configurations.

Problem 6P

Write Lewis structures, including unshared pairs, for each of the following. Carbon has four bonds in each compound.

(a) Propane (C_3H_8)

(b) Methanol (CH_4O)

(c) Methyl fluoride (CH_3F)

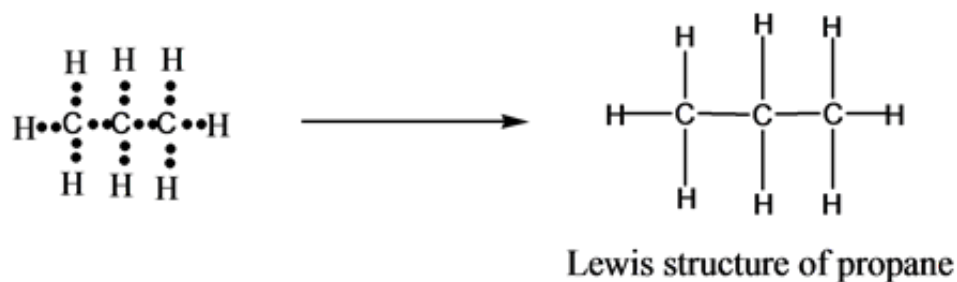
(d) Ethyl fluoride (C_2H_5F)

Step-by-step solution

step 1 of 5

(a)

Lewis dot structure of propane (C_3H_8) is as follows:



The ten covalent bonds, shown in the Lewis structure, account for 20 valence electrons, which is the same as that calculated from the molecular formula (C_3H_8).

The eight hydrogens of (C_3H_8) contribute 1 electron each and the three carbons contribute 4 each, for a total of 20 (8 from the hydrogen and 12 from the carbon).

Therefore, we say that the valence electrons are in covalent bonds; propane has no unshared pairs.

step 2 of 5

(b)

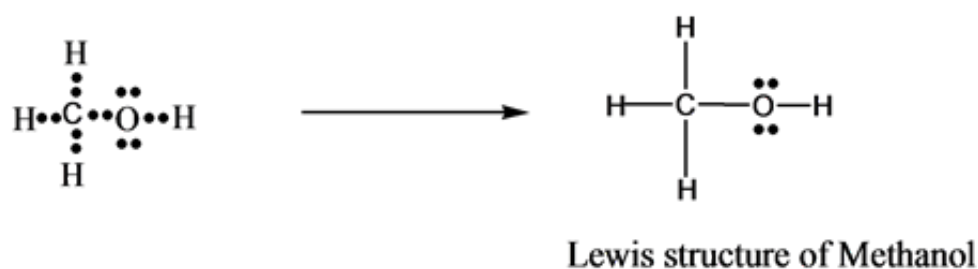
Lewis dot structure of methanol (CH_4O)

Number of valence electrons in methanol (CH_4O)

Atom	Valence electrons
Four hydrogens	4
One carbon	4
One oxygen	6
	Total = 14

The total number of covalent bonds is 5. So, ten electrons ($5 \times 2 = 10$) form 5 covalent bonds.

Thus, the $(14 - 10) = 4$ electrons of the oxygen atom that are not involved in bonding comprise two unshared pairs.



step 3 of 5

(c)

Lewis-dot structure of methyl fluoride (CH_3F)

Number of valence electrons in methyl fluoride (CH_3F)

Atom	Valence electrons
C	4
F	7
H	3 (1 from each hydrogen)
	Total = 14

The number of covalent bonds in (CH_3F) is four. Thus, $(4 \times 2) = 8$ valence electrons form four covalent bonds.

Thus, the $(14 - 8) = 6$ valence electrons of the fluorine atom that are not involved in bonding comprise three unshared pairs.



Lewis structure of Methyl fluoride

step 4 of 5

(d)

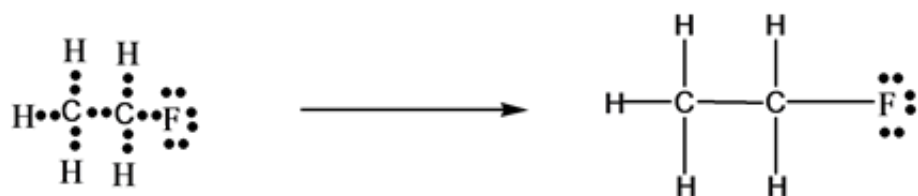
Lewis-dot structure of ethyl fluoride ($\text{C}_2\text{H}_5\text{F}$)

Atom	Valence electrons
Two carbons	8 (4 electrons from each of the two carbons)
Five hydrogens	5 (1 electron from each hydrogen)
One fluorine	7
	Total = 20

The number of covalent bonds in ethyl fluoride ($\text{C}_2\text{H}_5\text{F}$) is seven. So, $(7 \times 2) = 14$ electrons form seven covalent bonds.

step 5 of 5

Therefore, the $(20 - 14) = 6$ electrons of fluorine atom that are not involved in bonding comprise three unshared pairs.



Lewis structure of Ethyl fluoride

Problem 7P

All of the hydrogens are bonded to carbon in both of the following. Write a Lewis structure that satisfies the octet rule for each.

(a) Formaldehyde (CH₂O)

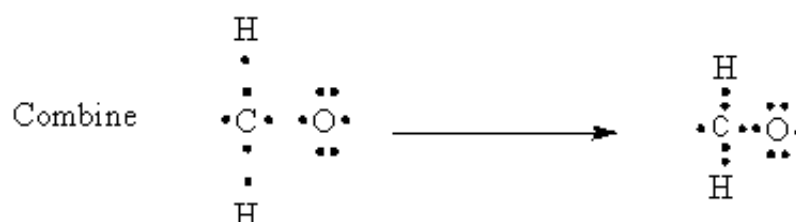
(b) Hydrogen cyanide (HCN)

Step-by-step solution

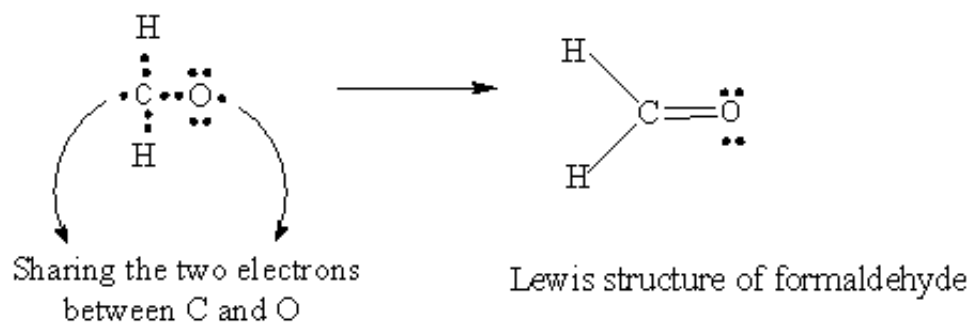
step 1 of 2

(a) Lewis structure for formaldehyde (CH_2O)

Atom	Valence electrons
One carbon	4
One oxygen	6
Two hydrogens	2 (one from each hydrogen)
	Total = 12

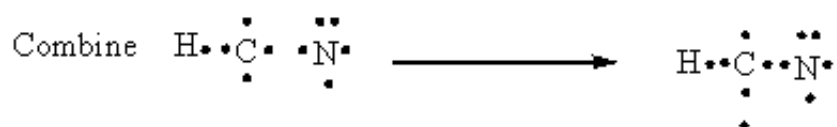


Now, the unpaired electron on carbon gets paired with the unpaired electron on oxygen to give a carbon oxygen double bond. The resulting structural formula satisfies the octet rule.

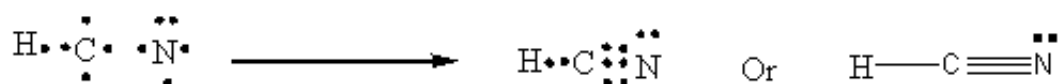


(b) Lewis structure of hydrogen cyanide (HCN)

Atom	Valence electrons
One hydrogen	1
One carbon	4
One nitrogen	5
	Total = 10



Now, the two unpaired electrons on carbon get paired with the two unpaired electrons on nitrogen to give a carbon nitrogen triple bond. The resulting structural formula satisfies the octet rule.



Lewis structure of HCN

Problem 8P

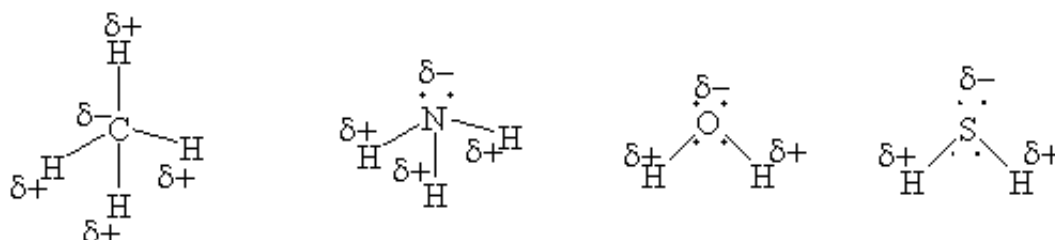
In which of the compounds CH_4 , NH_3 , H_2O , SiH_4 , or H_2S is δ^+ for hydrogen the greatest? In which one does hydrogen bear a partial negative charge?

Step-by-step solution

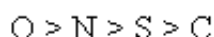
The greater the electronegativity difference between two elements, the more polar the bond between them is, and consequently the bonded atoms carry more partial charge.

In the case of compounds CH_4 , CH_3 , H_2O , SiH_4 , and H_2S , due to the electronegativity differences between the bonded atoms, each compound is polar.

Among them, hydrogen carries δ^+ in CH_4 , NH_3 , H_2O , and H_2S .



The order of electronegativity of the central atom is



Since oxygen is the most electronegative atom among them, then in compound H_2O , δ^+ for hydrogen is the greatest.

In the case of compound SiH_4 , hydrogen is more electronegative than Si. So here, hydrogen bears a partial negative charge.

Problem 9P

Indicate the direction of the dipole for the following bonds using the symbol \longleftrightarrow and δ^+ , δ^- notation.



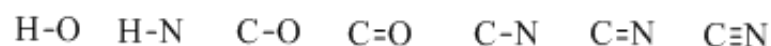
Step-by-step solution

step 1 of 2

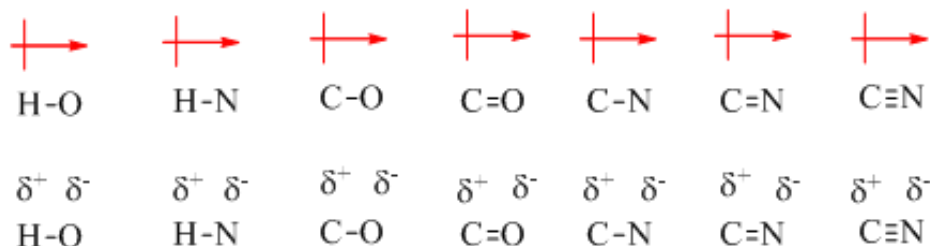
An electronic dipole exists whenever two dissimilar atoms are connected by a chemical bond. The unequal electronegativity values of atoms states that the atom which is more electronegative will attract more of the electron density than the other atom. This builds up a partial negative charge on the atom with higher electronegativity, and a partial positive charge on the atom with lower electronegativity.

The existence and direction of an electronic dipole is designated by a special red arrow, with the head of the arrow pointing towards the atom possessing higher electronegativity.

For example, consider the following chemical bonds:



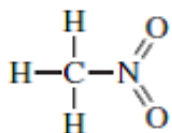
We can use the special red arrow along with Greek special characters to show the existence and direction of the dipole, along with the atoms which possess partial positive and partial negative charges.



Oxygen is more electronegative than hydrogen, and so the oxygen atom attracts more electron density from the bond than hydrogen. The dipole points towards the oxygen atom, and as expected, the oxygen develops a partial positive charge (leaving a partial negative charge on the hydrogen). Similar logic is used to explain the remaining examples. Nitrogen is more electronegative than hydrogen, thus oxygen has more electronegative than carbon and nitrogen is more electronegative than carbon.

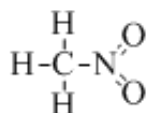
Problem 10P

Why is the formula shown for nitromethane incorrect?



Step-by-step solution

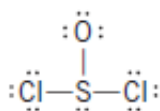
Being able to recognize whether a structural formula is correct or incorrect is an important skill in organic chemistry. Consider the following structural for nitromethane molecule.



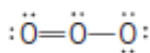
In the above structure, the nitrogen atom contains five bonds with other atoms in which two double bonds and one single bond. As we know that each bond has two electrons total 10 valence electrons must be contain for five bonds formed by nitrogen atom. But this is a violation of an octet rule. Here, we remember that the nitrogen atom is not so far enough down in the periodic table it has to access electrons from others to get an octet number as like d or f orbitals. It must have most total eight electrons in its valence shells to get high stability as like inert gases. Therefore, the given structure is not correct to the nitromethane molecule.

Problem 11P

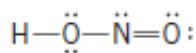
The following inorganic species will be encountered in this text. Calculate the formal charge on each of the atoms in the Lewis formulas given.



(a) Thionyl chloride



(b) Ozone



(c) Nitrous acid

Step-by-step solution

step 1 of 6

The formal charge is the difference between the number of valence electrons in the neutral atom and the electron count in the Lewis formula.

(a)

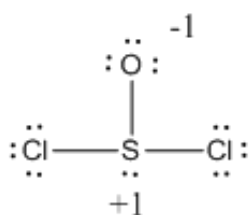
For the thionyl chloride molecule the formal charges on the atoms are counted as following:

Atom	Valence electrons on the neutral atom	Electron count	Formal charge
Sulfur	6	$\frac{1}{2} (6) + 2 = 5$	+1
Oxygen	6	$\frac{1}{2} (2) + 6 = 7$	-1
Chlorine	7	$\frac{1}{2} (2) + 6 = 7$	0

step 2 of 6

The formal charges are shown in the Lewis structure of thionyl chloride is as follows.

step 3 of 6



step 4 of 6

(b)

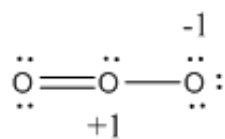
For the ozone molecule the formal charges on the atoms are counted as following:

Atom	Valence electrons on the neutral atom	Electron count	Formal charge
Oxygen (Left)	6	$\frac{1}{2} (4) + 4 = 6$	0
Oxygen (Mid)	6	$\frac{1}{2} (6) + 2 = 5$	+1
Oxygen (Right)	6	$\frac{1}{2} (2) + 6 = 7$	-1

step 5 of 6

The formal charges are shown in the Lewis structure of ozone is as follows.

step 6 of 6

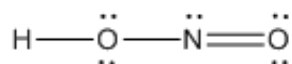


(c)

For then nitrous acid molecule the formal charges on the atoms are counted as following:

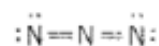
Atom	Valence electrons on the neutral atom	Electron count	Formal charge
Hydrogen	1	$\frac{1}{2} (2) + 0 = 1$	0
Oxygen (Left)	6	$\frac{1}{2} (4) + 4 = 6$	0
Nitrogen	5	$\frac{1}{2} (6) + 2 = 5$	0
Oxygen (Right)	6	$\frac{1}{2} (4) + 4 = 6$	0

The formal charges are shown in the Lewis structure of nitrous acid is as follows.



Problem 12P

Calculate the formal charge on each nitrogen in the following Lewis structure (azide ion) and the net charge on the species.



Step-by-step solution

step 1 of 1

The formal charge is the difference between the number of valence electrons in the neutral atom and the electron count in the Lewis formula.

For the Azide ion the formal charges on the atoms are counted as following:

Atom	Valence Electrons on the neutral atom	Electron count	Formal charge
Nitrogen (Left)	5	$\frac{1}{2} (4) + 4 = 6$	-1
Nitrogen (Mid)	5	$\frac{1}{2} (8) + 0 = 4$	+1
Nitrogen (Right)	5	$\frac{1}{2} (4) + 4 = 6$	-1

So, the net charge of the species is calculated as

$$(-1)+1+(-1)=-1$$

Problem 13P

Write structural formulas for all the constitutional isomers that have the given molecular formula.

(a) C_2H_7N

(b) C_3H_7Cl

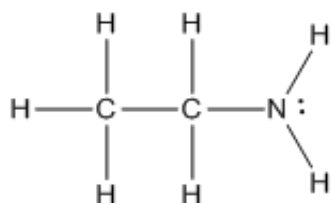
(c) C_3H_8O

Step-by-step solution

step 1 of 3

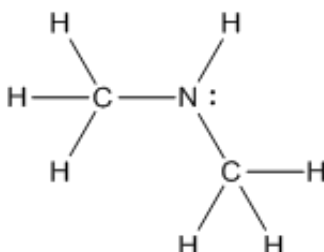
(a)

The structural formulas for all the constitutional isomers with molecular formula C_2H_7N are shown below:



ethanamine

Chemical Formula: C_2H_7N



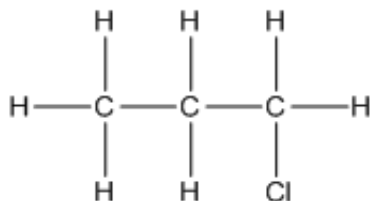
dimethylamine

Chemical Formula: C_2H_7N

step 2 of 3

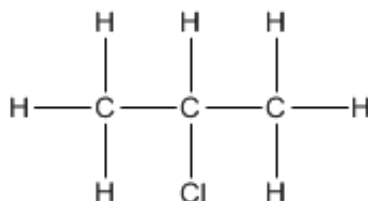
(b)

The structural formulas for all the constitutional isomers with molecular formula C_3H_7Cl are shown below:



1-chloropropane

Chemical Formula: C_3H_7Cl

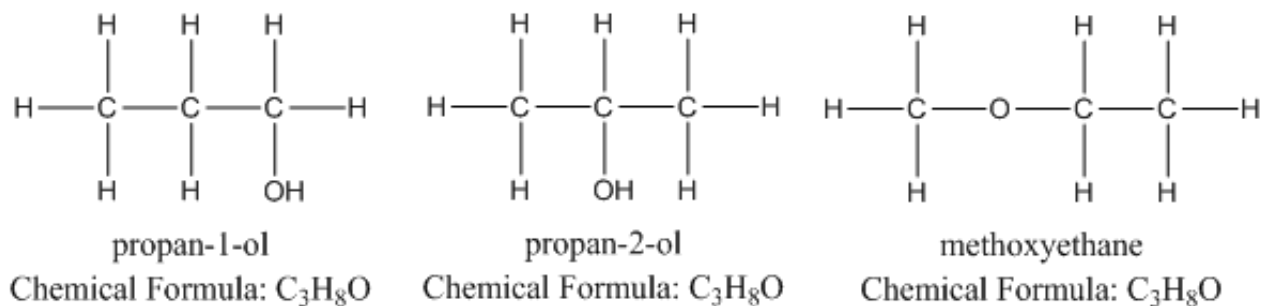


2-chloropropane

Chemical Formula: C_3H_7Cl

(c)

The structural formulas for all the constitutional isomers with molecular formula $\text{C}_3\text{H}_8\text{O}$ are shown below:



Problem 14P

Nitrosomethane and formaldoxime both have the molecular formula CH_3NO and the connectivity CN0 . All of the hydrogens are bonded to carbon in nitrosomethane. In formaldoxime, two of the hydrogens are bonded to carbon and one to oxygen. Write Lewis formulas for (a) nitrosomethane and (b) formaldoxime that satisfy the octet rule and are free of charge separation.

Sample Solution:

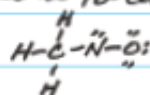
(a) • The connectivity is $\text{H}-\overset{\text{H}}{\underset{\text{H}}{\text{C}}}-\text{N}-\text{O}$ which accounts for 10 valence electrons.

• Total number of valence electrons

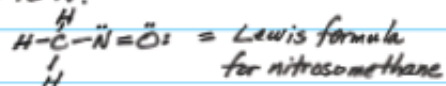
$$\left. \begin{array}{l} 3\text{H} = 3 \\ \text{C} = 4 \\ \text{N} = 5 \\ \text{O} = 6 \end{array} \right\} = 18$$

• 5 Bonds account for 10 electrons.

• Assign remaining 8 electrons to O and N in pairs so as to complete octets. Begin with O.



• N has only 6 electrons. Use one of the pairs assigned to O to form a double bond to N.

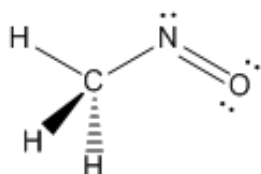


Step-by-step solution

step 1 of 2

(a)

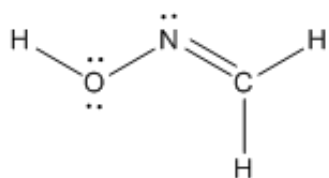
Lewis dot structure of nitrosomethane is shown below:



step 2 of 2

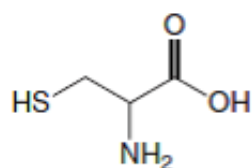
(b)

Lewis dot structure of formaldoxime is shown below:

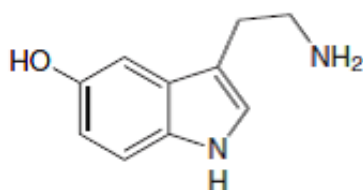


Problem 15P

Expand the bond-line formulas of the amino acid cysteine and the neurotransmitter serotonin to show all the unshared electron pairs. Molecular formulas of organic compounds are customarily presented in the fashion $C_aH_bX_cY_d$. Carbon and hydrogen are cited first, followed by the other atoms in alphabetical order. What are the molecular formulas of cysteine and serotonin?



Cysteine



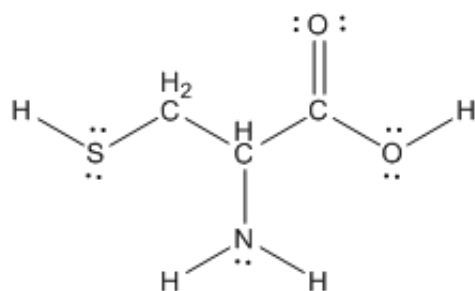
Serotonin

Step-by-step solution

step 1 of 2

(a)

Lewis dot structure of Cysteine is shown below:

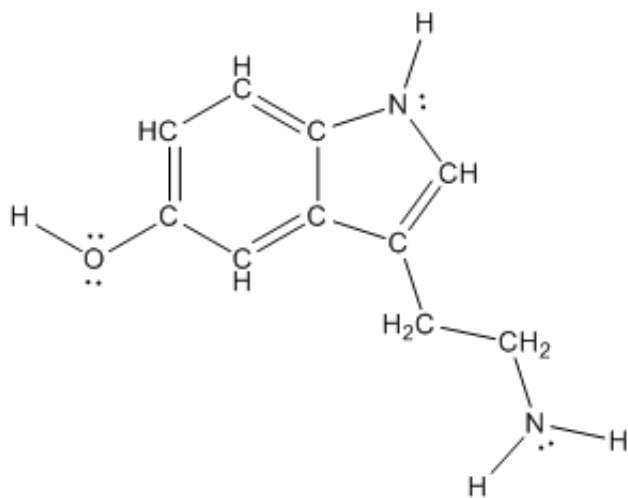


Chemical Formula: $C_3H_7NO_2S$

step 2 of 2

(b)

Lewis dot structure of Serotonin is shown below:



Chemical Formula: $C_{10}H_{12}N_2O$

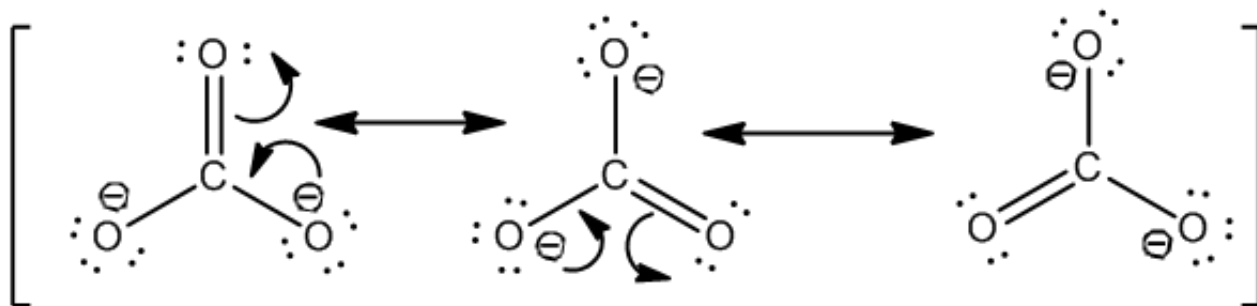
Problem 16P

All of the bonds in the carbonate ion (CO_3^{2-}) are between C and O. Write Lewis structures for the major resonance contributors, and use curved arrows to show their relationship. Apply the resonance concept to explain why all of the C—O bond distances in carbonate are equal.

Step-by-step solution

step 1 of 2

Lewis dot structures of the contributing canonical forms of carbonate anion are shown below:



step 2 of 2

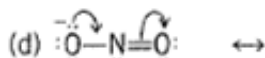
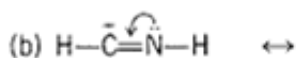
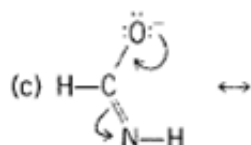
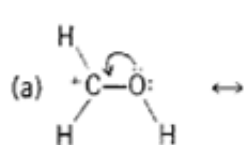
For all the Carbon-Oxygen bonds the resultant bond order considering the canonical forms is

$$\frac{2+1+1}{3} = \frac{4}{3}$$

As the bond orders of all the Carbon-Oxygen bonds are equal the bond lengths will also be identical.

Problem 17P

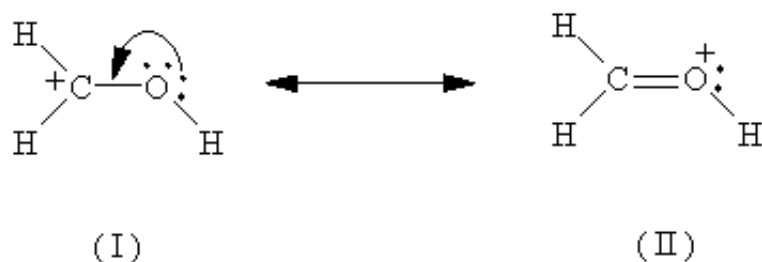
Write the resonance structure obtained by moving electrons as indicated by the curved arrows. Compare the stabilities of the two Lewis structures according to the guidelines in Table 1.6. Are the two Lewis structures equally stable, or is one more stable than the other? Why?



Step-by-step solution

step 1 of 4

(a)



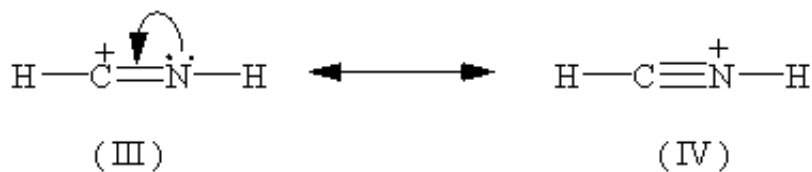
Structure (II) is more stable because it has one more covalent bond than the original structure (I).

In structure (I), carbon did not have an octet of electrons, but the octet rule is satisfied for both carbon and oxygen in structure (II).

Therefore, structure (II) is more stable than structure (I).

step 2 of 4

(b)



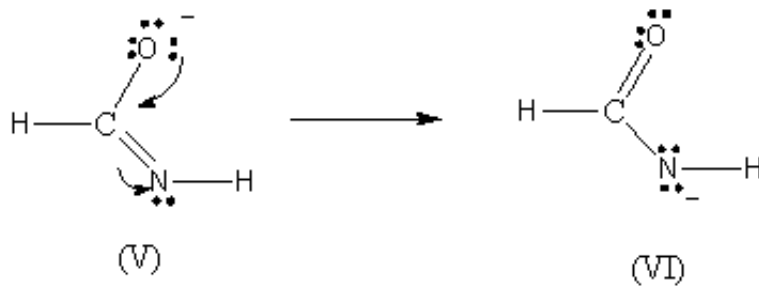
In structure (III), carbon did not have an octet of electrons, but the octet rule is satisfied for both carbon and nitrogen in structure (IV).

Moreover, structure (IV) has one more covalent bond than the original structure (III).

Due to these two factors, structure (IV) is more stable and a more contributing resonating structure than that of structure (III).

step 3 of 4

(c)



In both structures (V) and (VI), the octet rule is satisfied for the O, C, and N atoms. Both structures (V) and (VI) contain equal numbers of covalent bonds.

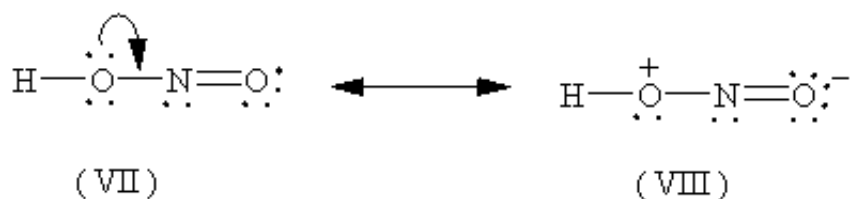
However, in structure V, there is a negative charge on oxygen, whereas in structure (VI), there is a negative charge on nitrogen.

Oxygen is more electronegative than nitrogen and can better support a negative charge. According to rules of resonance, the major contributing structure is the one in which the negative charge resides on the most electronegative atom.

Thus, structure (V) is more stable and more contributing than that of structure (VI).

step 4 of 4

(d)



Both structures (VII) and (VIII) satisfy the octet rule, and contain equal numbers of covalent bonds.

However, structure (VII) is the major contributor because it lacks the separation of positive and negative charges that characterize structure (VIII).

Thus, structure (VII) is more stable than structure (VIII).

Problem 18P

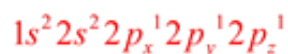
Of the four structural formulas shown, three are permissible and one is not. Which one is not a permissible structure? Why?



Step-by-step solution

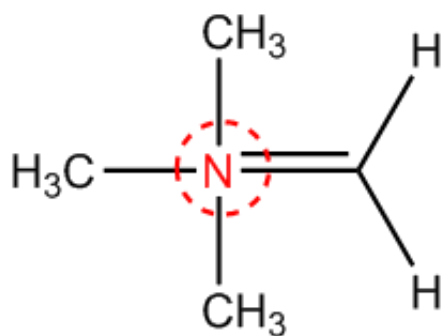
step 1 of 1

Nitrogen atom has the electronic configuration



Due to absence of d orbital in the second period element nitrogen cannot exceed its valency beyond 4.

The impossible molecular structure of is shown in the following diagram



Problem 19P

The salt sodium borohydride, NaBH_4 has an ionic bond between Na^+ and the anion BH_4^- . What are the $\text{H}-\text{B}-\text{H}$ angles in borohydride anion?

Step-by-step solution

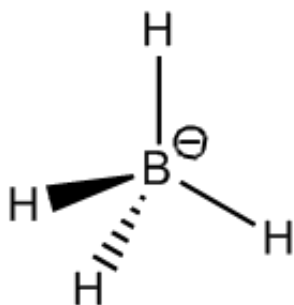
step 1 of 1

Boron atom has the electronic configuration



In the anion BH_4^- there are 4 pairs of electron on the central Boron atom. Hence, the hybridization of the central Boron atom is sp^3 and the geometry of the molecule is tetrahedral.

The molecular structure of the anion is shown in the following diagram



So, the $\text{H}-\text{B}-\text{H}$ angle in the anion is $109^\circ 28'$ as predicted for a tetrahedral geometry.

Problem 20P

Specify the shape of the following:

(a) $\text{H}-\text{C}\equiv\text{N}$: (Hydrogen cyanide)

(b) H_4N^+ (Ammonium ion)

(c) N_3^- (Azide ion)

(d) CO_3^{2-} (Carbonate ion)

Step-by-step solution

step 1 of 5

The three dimensional arrangement of atoms in a molecule is called molecular geometry. Number of electrons around the central atom in its Lewis structure helps to predict the geometry of

the molecule or ion. Valence shell is the outermost shell of an atom occupied with electrons which are generally involved in bonding. In a covalent bond, a pair of electrons is present called bonding pair, the electrons in different bonding pairs should be as far as apart as possible to prevent repulsion between them.

The geometry that the molecule finally attains has minimum repulsion, study through this approach is known as valence-shell electron pair **(VSEPR)**. The two general rules govern the use of VSEPR model as follows:

For the electron-pair repulsion, double and triple bonds are assumed as single bonds although former are larger than single bond but this good for qualitative purposes.

For two or more resonance structures of a molecule, VSEPR model is applied to any one of them.

For convenience molecules are divided in two categories as one with lone pairs on central atom and one which doesn't have lone pair on central atom. When the central atom has one or more than one lone pairs then repulsive forces decrease as follows:

Lone pair-lone pair repulsion > lone pair-bonding pair repulsion > bonding pair-bonding pair repulsion

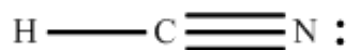
Shape of a molecule is predicted by excluding the lone pairs. The shape is not given by any geometry but it can be anything from T-shape to linear.

step 2 of 5

(a)

In the molecule **HCN** there is a carbon-nitrogen triple bond which contains 2 π bonds. 2 p orbitals of carbon are required to form these π bonds. As 2 p orbitals are kept aside for π bonding the central carbon atom is sp hybridized. Therefore, the molecule is linear in shape.

The shape of the molecule is as follows:

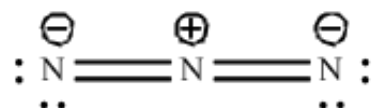


step 3 of 5

(b)

In the anion **N_3^-** there are two nitrogen-nitrogen double bonds which contain 2 π bonds. 2 p orbitals of nitrogen are required to form these π bonds. As 2 p orbitals are kept aside for π bonding the central nitrogen atom is sp hybridized. Therefore, the anion is linear in shape.

The shape of the molecule is as follows:

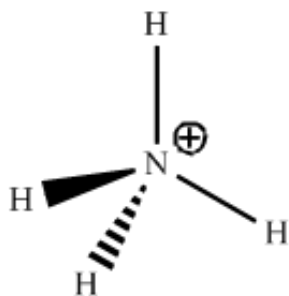


step 4 of 5

(c)

In the cation **NH_4^+** there are 4 pairs of electron on the central nitrogen atom. Hence, the hybridization of the central nitrogen atom is sp^3 and the shape of the cation is tetrahedral.

The shape of the molecule is as follows:

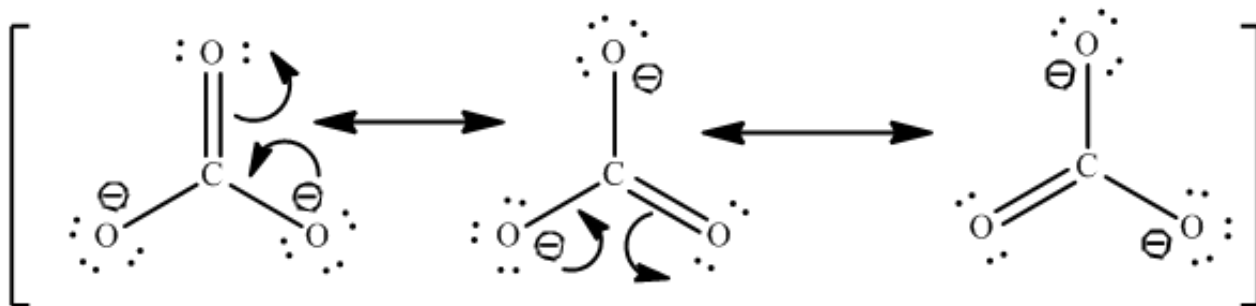


step 5 of 5

(d)

In the carbonate anion there is a carbon-oxygen double bond which contains a π bond. 1 p orbital of carbon is required to form this π bond. As 1 p orbital is kept aside for π bonding the central carbon atom is sp^2 hybridized. Therefore, the anion is trigonal planar in shape.

The shape of the molecule is as follows:



Problem 21P

Which of the following compounds would you expect to have a dipole moment? If the molecule has a dipole moment, specify its direction.

(a) BF_3

(b) H_2O

(c) CH_4

(d) CH_3Cl

(e) CH_2O

(f) HCN

Step-by-step solution

step 1 of 7

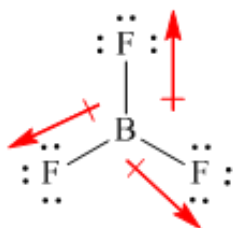
An electronic dipole exists when two dissimilar atoms are connected by a chemical bond.

The existence and direction of an electronic dipole is designated by a special red arrow, with the head of the arrow pointing towards the atom possessing higher electronegativity.

step 2 of 7

(a)

In boron trifluoride (BF_3), fluorine is more electronegative than boron. So, the direction of dipole is as follows:



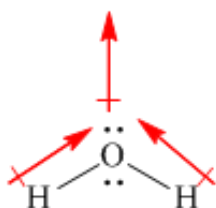
The dipole moment of a molecule is the vector sum of the bond dipoles and includes the dipoles due to the presence of lone pairs on the central atom.

Therefore, the molecule has no dipole moment because of the mutual cancellation of individual bond dipoles.

step 3 of 7

(b)

In water molecule (H_2O), oxygen is more electronegative than hydrogen. So, the direction of dipole is as follows:



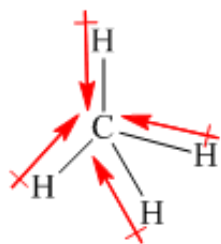
The dipole moment of a molecule is the vector sum of the bond dipoles and includes the dipoles due to the presence of lone pairs on the central atom.

Thus, the water molecule is expected to have a dipole moment.

step 4 of 7

(c)

In methane (CH_4), carbon is more electronegative than hydrogen. So, the direction of dipole is as follows:

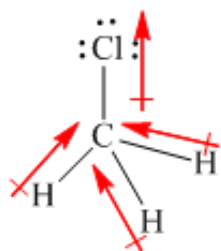


Therefore, the molecule has no dipole moment because of the mutual cancellation of individual bond dipoles.

step 5 of 7

(d)

In chloroform (CHCl_3), chlorine is more electronegative than carbon; carbon is more electronegative than hydrogen. So, the direction of dipole is as follows:



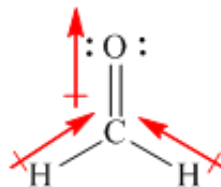
The dipole moment of a molecule is the vector sum of the individual bond dipole moments of the molecule.

Therefore, the molecule has a dipole moment as there is no mutual cancellation of individual bond dipoles.

step 6 of 7

(e)

In formaldehyde, oxygen is more electronegative than carbon atom; carbon is more electronegative than hydrogen atom. So, the direction of dipole is as follows:

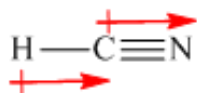


Therefore, the molecule has a dipole moment as there is no mutual cancellation of individual bond dipoles.

step 7 of 7

(f)

In hydrogen cyanide, carbon is more electronegative than hydrogen and nitrogen is more electronegative than carbon. So, the direction of dipole is as follows:

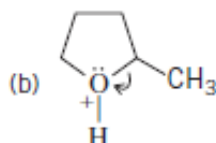
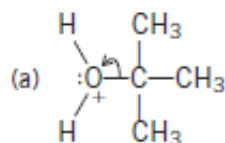


Therefore, the molecule has a dipole moment as there is no mutual cancellation of individual

bond dipoles.

Problem 22P

Using the curved arrow to guide your reasoning, show the products of the following dissociations. Include formal charges and unshared electron pairs. Check your answers to ensure that charge is conserved.

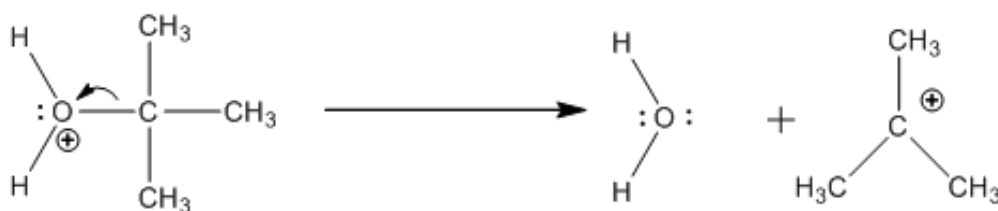


Step-by-step solution

step 1 of 2

(a)

The curved arrow tells us that the $\text{C}-\text{O}$ bond is cleaved and the electron pair involved in the σ bond moves to the more electronegative oxygen atom.

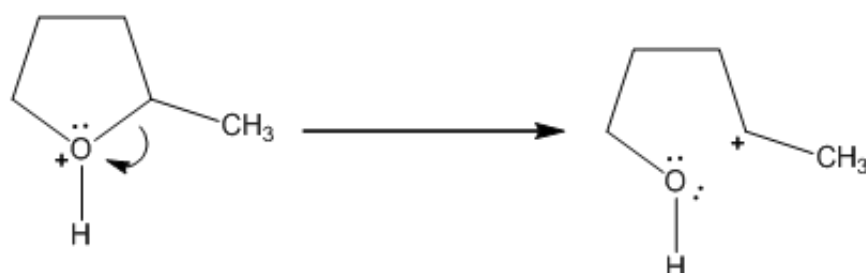


Water is eliminated as the leaving group. The carbocation generated is a stable tertiary carbocation. The central carbon atom contains six electrons and formal charge on the carbon atom is +1. So, the net charge is conserved in the reaction.

step 2 of 2

(b)

The curved arrow tells us that the $\text{C}-\text{O}$ bond is cleaved and the electron pair involved in the σ bond moves to the more electronegative oxygen atom.

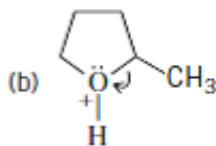
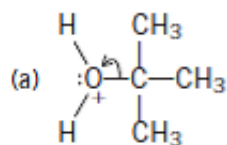


The protonated cyclic ether is broken down in the reaction. The carbocation generated is a stable secondary carbocation. The positive carbon atom contains six electrons and formal charge on the carbon atom is +1. So, the net charge is conserved in the reaction.

Problem 23P

Write equations, including curved arrows, describing the reverse reactions of Problem 1.22.

Using the curved arrow to guide your reasoning, show the products of the following dissociations. Include formal charges and unshared electron pairs. Check your answers to ensure that charge is conserved.

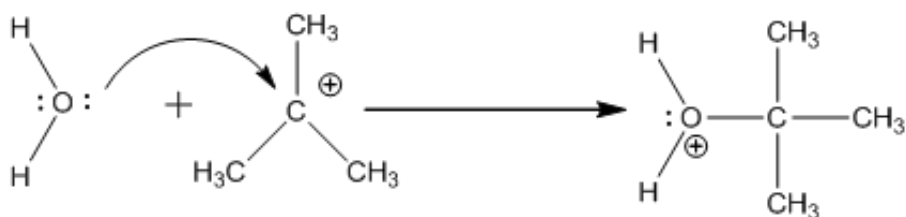


Step-by-step solution

step 1 of 2

(a)

The curved arrow tells us that the $\text{C}-\text{O}$ bond is formed as the lone pair of oxygen atom of the water molecule attacks the electron deficient carbocation.

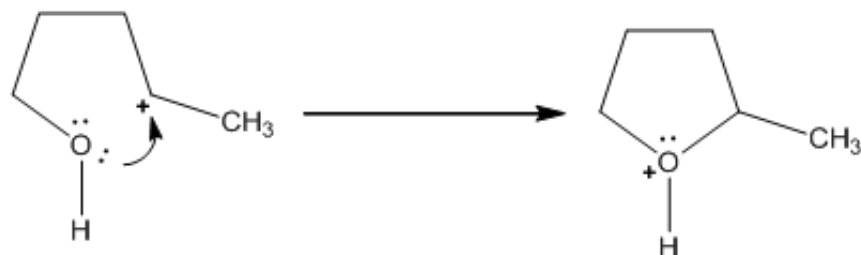


The cation generated is an oxonium cation. The oxygen atom contains five electrons and formal charge on the oxygen atom is +1. So, the net charge is conserved in the reaction.

step 2 of 2

(b)

The curved arrow tells us that the $\text{C}-\text{O}$ bond is formed as the $-\text{OH}$ group acts as the internal nucleophile and attacks the electron deficient carbocation.

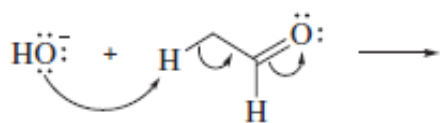


The cation generated is an oxonium cation. The oxygen atom contains five electrons and formal charge on the oxygen atom is +1. So, the net charge is conserved in the reaction.

Problem 24P

Reactions of the type shown are an important part of

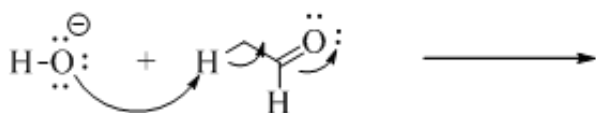
Chapter 20. Follow the arrows to predict the products. Show formal charges and include all unshared electron pairs.



Step-by-step solution

step 1 of 2

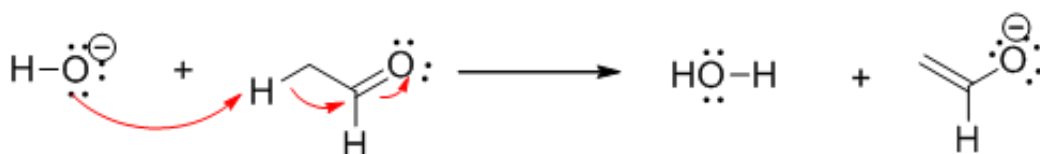
Consider the following chemical reaction.



This reaction shows hydroxide anion (a moderately strong base) reaching out and abstracting a proton from the organic compound, using one of the lone pairs on the oxygen atom. As the hydrogen leaves, the two electrons that were previously in the carbon-hydrogen bond swing over and settle on the electrophilic carbonyl carbon. This carbon is the target because the broken bond's electrons are negatively charged and nucleophilic. The carbonyl carbon is electrophilic because of the electronegative oxygen atom, which leads to a partial positive charge building up on the carbonyl carbon. It presents an attractive target for the nearby electrons.

step 2 of 2

The five-bonded carbon violates the octet rule because its valence is four. The double bond presents between carbon and oxygen has become to a single bond in the product. One of the lone pair electrons in the double bond will move up and settles at the high electronegative oxygen atom. The negative charge on oxygen atom can stabilize the extra charge as a third lone pair of electrons. The structure of species in the completed reaction is shown as follows:



The overall result of the reaction is a molecule of water as well as a negatively charged Alkene was formed.

Problem 25P

Write an equation for proton transfer from hydrogen chloride (HCl) to

(a) Ammonia (:NH_3)

(b) Trimethylamine [$(\text{CH}_3)_3\text{N:}$]

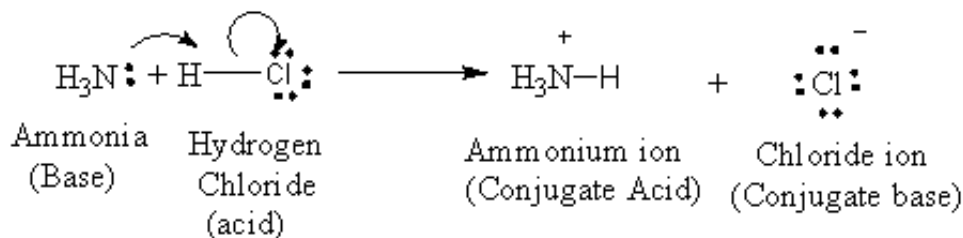
Identify the acid, base, conjugate acid, and conjugate base and use curved arrows to track electron movement.

Step-by-step solution

step 1 of 2

(a) Ammonia (:NH_3)

When ammonia reacts with HCl , a proton is transferred from HCl to :NH_3 leading to the formation of the ammonium ion and chloride ion. Hence, HCl is the Bronsted acid and :NH_3 is the Bronsted base.

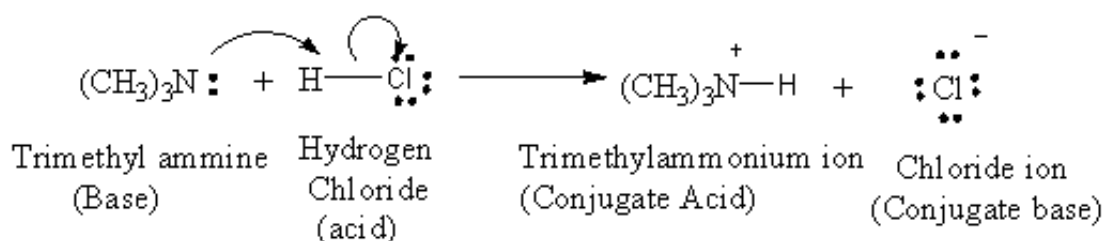


step 2 of 2

(b)

Trimethylamine $\text{[(CH}_3\text{)}_3\text{N:]}$

When trimethylamine reacts with HCl , a proton is transferred from HCl to $\text{(CH}_3\text{)}_3\text{N:}$. Therefore, HCl is the Bronsted acid and $\text{(CH}_3\text{)}_3\text{N:}$ is the Bronsted base.



Problem 26P

Potassium hydride (KH) is a source of the strongly basic hydride ion (:H^-).

Using curved arrows to track electron movement, write an equation for the reaction of hydride ion with water. What is the conjugate acid of hydride ion?

Step-by-step solution

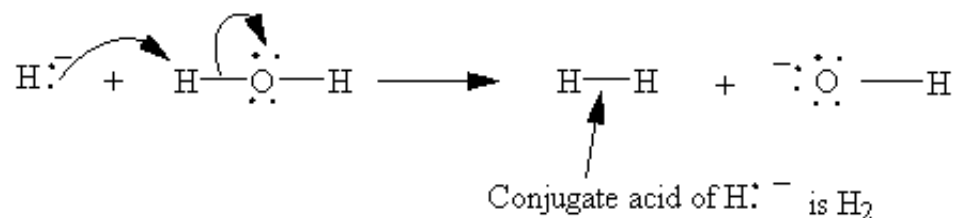
step 1 of 1

Potassium hydride (KH) is a source of the strongly base hydride ion (H^-).

We are told that H^- reacts with water.

We are told that H^- is a strong base.

Therefore, H^- removes a proton from water.



Problem 27P

Salicylic acid, the starting material for the preparation of aspirin, has a K_a of 1.06×10^{-3} . What is its $\text{p}K_a$?

Step-by-step solution

step 1 of 1

We know that $\text{p}K_a = -\log_{10} K_a$

Salicylic acid has a K_a of 1.06×10^{-3}

So

$$\text{p}K_a = -\log_{10} (1.06 \times 10^{-3})$$

$$\Rightarrow \text{p}K_a = 3 - \log 1.06$$

$$= 3 - 0.02530$$

$$= 2.974$$

Thus, $\text{p}K_a$ of salicylic acid is 2.974.

Problem 28P

Hydrogen cyanide (HCN) has a $\text{p}K_a$ of 9.1. What is its K_a ?

Step-by-step solution

step 1 of 1

Hydrogen cyanide has a pK_a of 9.1

We know that $pK_a = -\log_{10} K_a$

$$\Rightarrow 9.1 = -\log_{10} K_a$$

$$\Rightarrow \log_{10} K_a = -9.1$$

$$K_a = 10^{-9.1}$$

$$= 7.943 \times 10^{-10}$$

K_a of hydrogen cyanide is 7.943×10^{-10} .

Problem 29P

Which is the stronger base in each of the following pairs?
(Note: This information will prove useful when you get to Chapter 9.)

(a) Sodium ethoxide ($\text{NaOCH}_2\text{CH}_3$) or sodium amide (NaNH_2)

(b) Sodium acetylide ($\text{NaC}\equiv\text{CH}$) or sodium amide (NaNH_2)

(c) Sodium acetylide ($\text{NaC}\equiv\text{CH}$) or sodium ethoxide ($\text{NaOCH}_2\text{CH}_3$)

Step-by-step solution

step 1 of 3

(a)

$\text{NaOCH}_2\text{CH}_3$ contains the ions Na^+ and $\text{CH}_3\text{CH}_2\text{O}^-$. NaNH_2 contains the ions Na^+ and H_2N^- . $\text{CH}_3\text{CH}_2\text{O}^-$ is a conjugate base of ethanol, and H_2N^- is the conjugate base of ammonia.

Base	$\text{CH}_3\text{CH}_2\text{O}^-$	H_2N^-
Conjugate acid	$\text{CH}_3\text{CH}_2\text{OH}$	NH_3
pK_a of conjugate	16	36

The conjugate acid of $\text{CH}_3\text{CH}_2\text{O}^-$ is stronger than the conjugate acid of H_2N^- .

Therefore, H_2N^- is a stronger base than $\text{CH}_3\text{CH}_2\text{O}^-$.

step 2 of 3

(b) $\text{NaC}\equiv\text{CH}$ contains Na^+ and $\text{CH}\equiv\text{C}^-$. NaNH_2 contains the ions Na^+ and H_2N^- .

Base	Conjugate acid	$\text{p}K_{\text{a}}$ of conjugate
$\text{HC}\equiv\text{C}^-$	$\text{HC}\equiv\text{CH}$	26
NH_2^-	NH_3	36

The conjugate acid of $\text{HC}\equiv\text{C}^-$ is stronger than that of NH_2^- . Therefore, NH_2^- is a stronger base than $\text{HC}\equiv\text{C}^-$.

step 3 of 3

(c)

Sodium acetylide contains the ions Na^+ and $\text{HC}\equiv\text{C}^-$

$\text{NaOCH}_2\text{CH}_3$ contains the ions Na^+ and $\text{CH}_3\text{CH}_2\text{O}^-$

$\text{CH}\equiv\text{C}^-$ is the conjugate base of $\text{CH}\equiv\text{CH}$, and $\text{CH}_3\text{CH}_2\text{O}^-$ is a conjugate base of ethanol.

Base	$\text{CH}\equiv\text{C}^-$	$\text{CH}_3\text{CH}_2\text{O}^-$
Conjugate acid	$\text{CH}\equiv\text{CH}$	$\text{CH}_3\text{CH}_2\text{OH}$
$\text{p}K_{\text{a}}$ of conjugate acid	26	16

The conjugate acid of $\text{CH}_3\text{CH}_2\text{O}^-$ is stronger than the conjugate acid of $\text{CH}\equiv\text{C}^-$, so $\text{CH}\equiv\text{C}^-$ is a stronger base than $\text{CH}_3\text{CH}_2\text{O}^-$.

Problem 30P

Which is the stronger acid, H_2O or H_2S ? Which is the stronger base, HO^- or HS^- ? Check your predictions against the data in Table 1.8.

Step-by-step solution

step 1 of 2

A conjugate base is a base that is formed after an acid loses a proton, and a conjugate acid is an acid that is formed after a base accepts a proton. A Bronsted-Lowry acid is a proton donor, and a base is a proton acceptor.

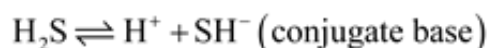
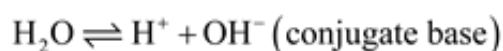
Weaker is the acid stronger is its conjugate base and Stronger is the acid weaker is its conjugate base.

The direction of equilibrium in an acid-base reaction is always in the favor of the reaction of the stronger acid (lower value pK_a) with the stronger base to form the weaker acid (higher value pK_a) and the weaker base. The quantity pK_a is the measure of the acidity of the acids and bases quantitatively.

More is the value of pK_a , stronger is the base that is weaker is the acid and lesser is the value of pK_a , weaker is the base that is stronger is the acid.

step 2 of 2

The reaction of H_2O and H_2S is as follows:



The acidity of an acid depends on its tendency to lose proton (H^+). The oxygen atom is smaller in size than sulfur atom; therefore the bond length between $O-H$ is smaller than $S-H$. Therefore H_2S is stronger acid than H_2O .

Weaker is the acid stronger is its conjugate base and Stronger is the acid weaker is its conjugate base. The molecule H_2O is weaker acid; therefore its conjugate base OH^- is stronger than SH^- .

Therefore, OH^- is a stronger base than SH^- .

Problem 31P

Try to do this problem without consulting Table 1.8.

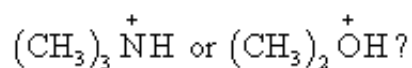
(a) Which is the stronger acid: $(CH_3)_3NH^+$ or $(CH_3)_2\ddot{O}H^+$?

(b) Which is the stronger base: $(CH_3)_3N:$ or $(CH_3)_2\ddot{O}:$?

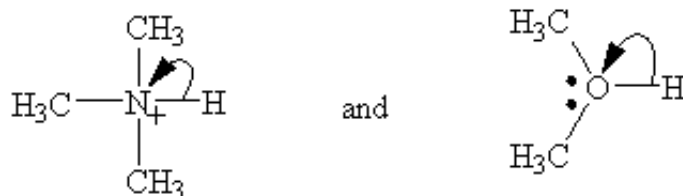
Step-by-step solution

step 1 of 2

(a)



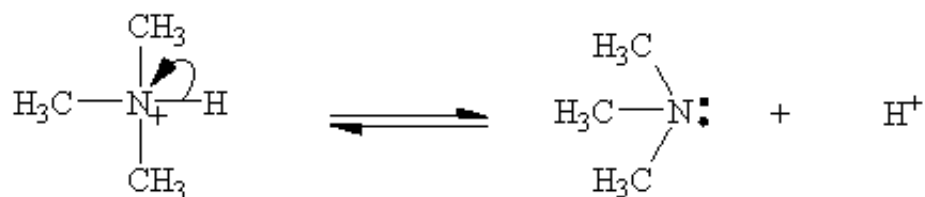
The ionizable proton is bonded to N in $(\text{CH}_3)_3\text{NH}^+$ and to O in $(\text{CH}_3)_2\text{OH}^+$.



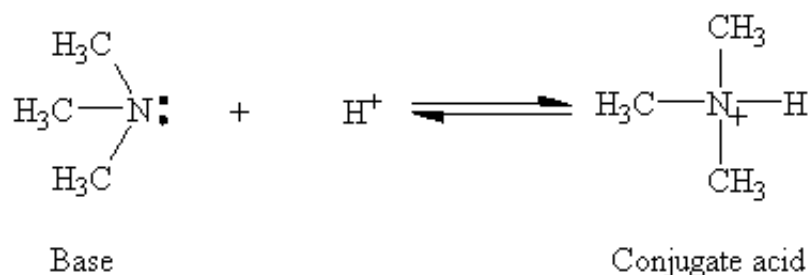
Nitrogen and oxygen are in the same row of the periodic table, so their relative electronegativities are the determining factor. Oxygen is more electronegative than nitrogen.

Therefore, $(\text{CH}_3)_2\text{OH}^+$ is a stronger acid than $(\text{CH}_3)_3\text{NH}^+$.

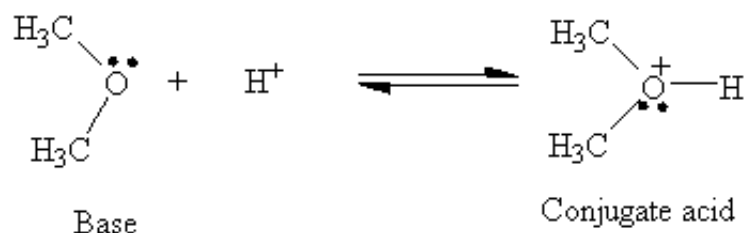
(b)



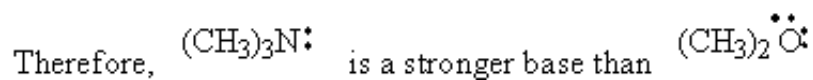
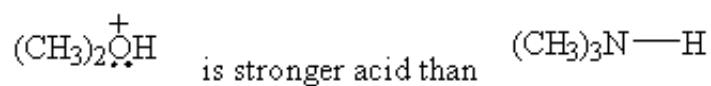
Or



Again,



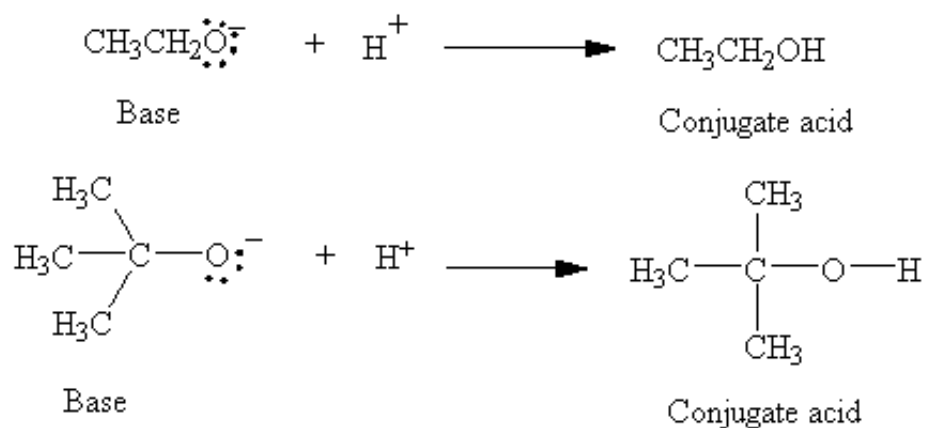
We know that the stronger the base, the weaker its conjugate acid, and vice-versa.



Problem 32P

Which is a stronger base, ethoxide $[\text{CH}_3\text{CH}_2\ddot{\text{O}}:]^-$ or tert-butoxide $[(\text{CH}_3)_3\text{C}\ddot{\text{O}}:]^-$?

Step-by-step solution



Acid	$\text{p}K_a$
$\text{CH}_3\text{CH}_2\text{OH}$	16
$(\text{CH}_3)_3\text{COH}$	18

So $\text{CH}_3\text{CH}_2\text{OH}$ is a stronger acid than $(\text{CH}_3)_3\text{C}-\text{OH}$

Thus, $\text{CH}_3\text{CH}_2\text{O}^-$ is a weaker base than $(\text{CH}_3)_3\text{CO}^-$

Therefore, $(\text{CH}_3)_3\text{CO}^-$ is a stronger base than $\text{CH}_3\text{CH}_2\text{O}^-$

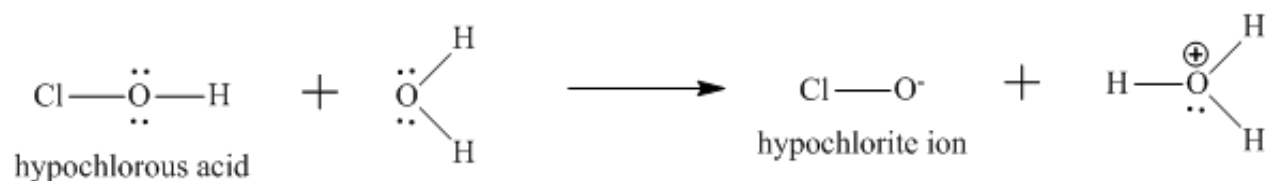
Problem 33P

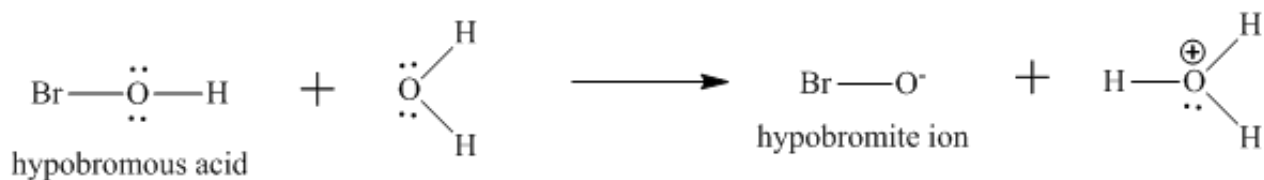
Hypochlorous and hypobromous acid (HOCl and HOBr) are weak acids. Write chemical equations for the ionization of each in water and predict which one is the stronger acid.

Step-by-step solution

step 1 of 2

The ionization reactions of Hypochlorous and hypobromous acid (HOCl and HOBr) in water are as follows:





step 2 of 2

The -I effect of chlorine is greater than -I effect of bromine because chlorine is more electronegative than bromine.

Due to stronger -I effect of chlorine, the conjugate base of HOCl , that is the hypochlorite ion ClO^- is stabilized more than that of BrO^- .

Thus, ClO^- is more stable than BrO^- .

Because of the greater stability of ClO^- , the equilibrium for ionization of HOCl lies further to the right than that of HOBr . Since, the stabilized conjugate base increases the acidity of the parent acid, HOCl is a stronger acid than HOBr .

Therefore, HOCl is a stronger acid than HOBr .

Problem 34P

What is the average formal charge on each oxygen in nitrate ion?

Step-by-step solution

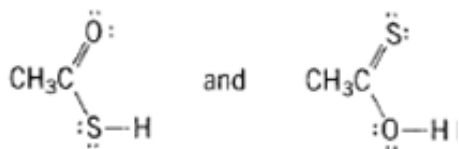
step 1 of 1

Both the negative charges are equally shared by three oxygens in the $-\text{NO}_3^-$ ion and all the canonical structures are equivalent.

Thus, the average formal charge on each oxygen in the $-\text{NO}_3^-$ ion $= -\frac{2}{3} = -0.67$

Problem 35P

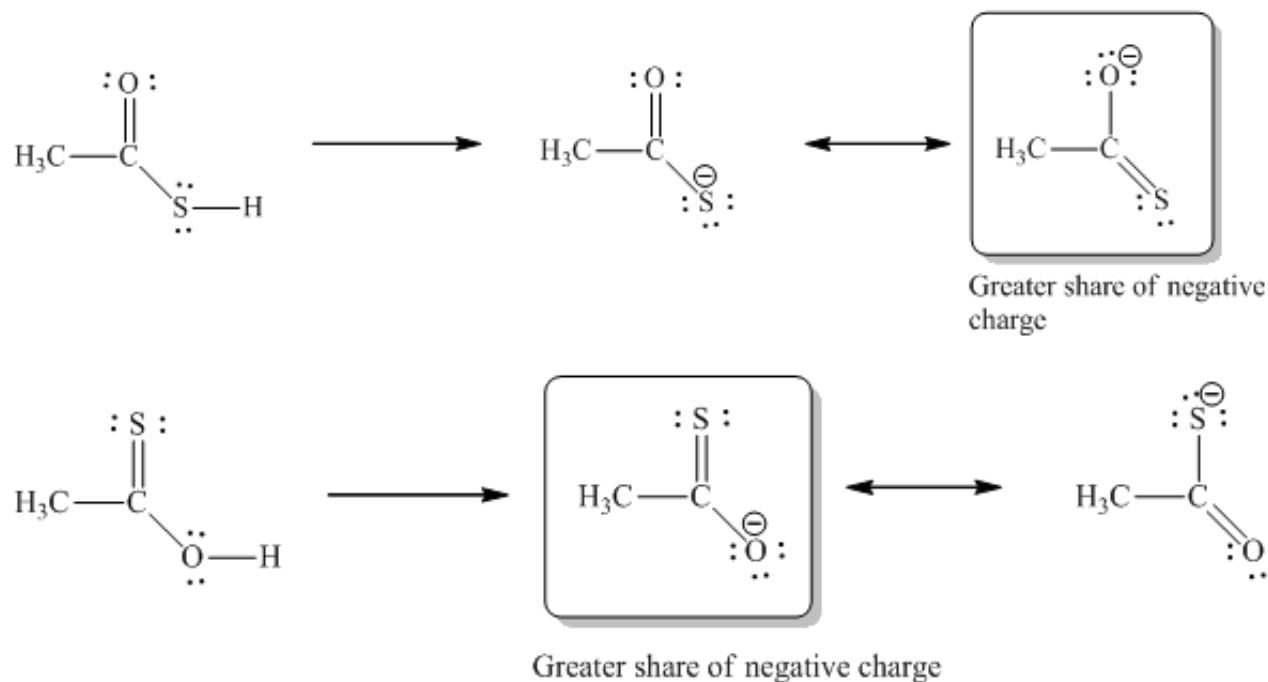
Show by writing appropriate resonance structures that the two compounds shown form the same conjugate base on ionization. Which atom in the conjugate base, O or S, bears the greater share of negative charge?



Step-by-step solution

step 1 of 2

The conjugate bases of the following acids on ionization are as follows:



step 2 of 2

For the conjugate base in which more electronegative element bears a negative charge, contributes to a great extent for the resonance structures.

The O atom, being more electronegative than the S atom, bears the greater share of the negative charge.

Therefore, the resonance structures in which the Oxygen atom bears a negative charge contributes a greater share of the negative charge.

Problem 36P

What is the equilibrium constant for the following acid-base reactions?

(a) ammonia and acetic acid

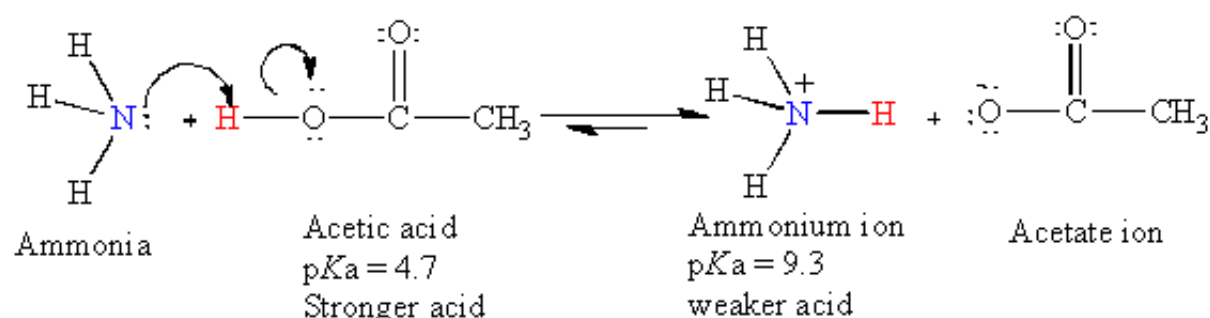
(b) fluoride ion and acetic acid

(c) ethanol and hydrobromic acid

Step-by-step solution

step 1 of 3

(a) Ammonia is a Bronsted base and accepts a proton from the -OH group of the acetic acid to form ammonium ion and acetate ion.



From the pK_a values, it is obvious that acetic acid is much stronger acid than the ammonium ion and hence the equilibrium lies to the right. The equilibrium for the process is

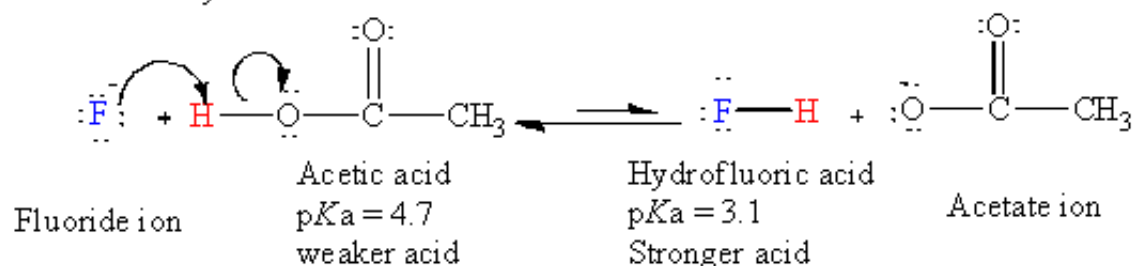
$$K_{eq} = \frac{10^{-pK_a} \text{ of acetic acid (reactant)}}{10^{-pK_a} \text{ of ammonium ion (product)}}$$

$$\text{or, } K_{eq} = \frac{10^{-4.7}}{10^{-9.3}}$$

$$\text{or, } \boxed{K_{eq} = 10^{4.6}}$$

step 2 of 3

(b) Fluoride ion is a Bronsted base and accepts a proton from the -OH group of the acetic acid to form hydrofluoric acid and acetate ion.



From the pK_a values, it is obvious that hydrofluoric acid is a stronger acid than the acetic acid and hence the equilibrium lies to the left. The equilibrium for the process is

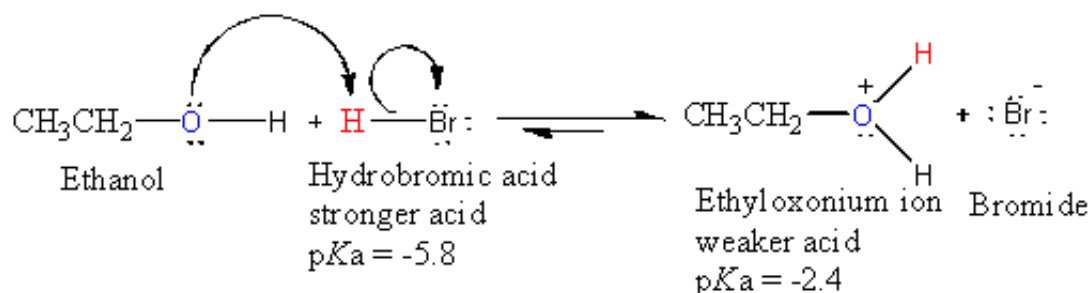
$$K_{eq} = \frac{10^{-pK_a} \text{ of acetic acid (reactant)}}{10^{-pK_a} \text{ of hydrofluoric acid (product)}}$$

$$\text{or, } K_{eq} = \frac{10^{-4.7}}{10^{-3.1}}$$

$$\text{or, } \boxed{K_{eq} = 10^{-1.6}}$$

step 3 of 3

(c) Ethanol is a Bronsted base and accepts a proton from HBr to form ethyloxonium ion and bromide ion.



From the pK_a values, it is obvious that hydrobromic acid is a stronger acid than the ethyloxonium ion and hence the equilibrium lies to the right. The equilibrium for the process is

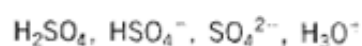
$$K_{eq} = \frac{10^{-pK_a \text{ of hydrobromic acid (reactant)}}}{10^{-pK_a \text{ of ethoxonium ion (product)}}$$

$$\text{or, } K_{eq} = \frac{10^{5.8}}{10^{2.4}}$$

$$\text{or, } \boxed{K_{eq} = 10^{3.4}}$$

Problem 37P

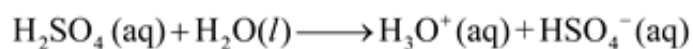
Rank the following in order of decreasing concentration in a solution prepared by dissolving 1.0 mol of sulfuric acid in enough water to give 1.0 L of solution. (It is not necessary to do any calculations.)



Step-by-step solution

step 1 of 3

Sulphuric acid is a strong acid, it dissociate almost completely to give hydronium ion and HSO_4^- .



step 2 of 3

HSO_4^- is a weak acid, it dissociate partially.



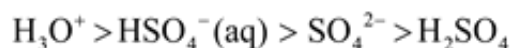
Thus, hydronium ion concentration is coming from both full and partial dissociation.

Therefore, among all hydronium ion concentration is more than others.

step 3 of 3

The concentration of $\text{SO}_4^{2-}(\text{aq})$ is lower than $\text{HSO}_4^-(\text{aq})$ as $\text{HSO}_4^-(\text{aq})$ incompletely dissociated.

Therefore, the order of decreasing concentration in the solution is as follows:



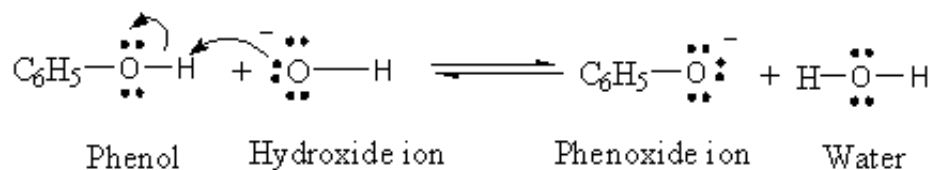
Problem 38P

Verify that the position of equilibrium for the reaction between phenol and hydroxide ion lies to the right by comparing the pK_a of the acid on the left to the acid on the right. Which acid is stronger? Do the same for the reaction of phenol with hydrogen carbonate ion.

Step-by-step solution

step 1 of 2

The equilibrium reaction of phenol with hydroxide ion:



Comparing the pK_a values of phenol and water,

Acid	pK_a
Phenol (Reactant)	10
Water (Product)	15.7

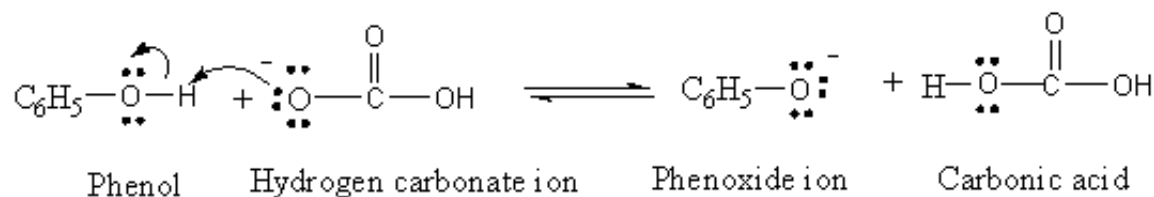
The greater the pK_a value, the weaker the acid. Phenol is a stronger acid than water, so the equilibrium lies to the right.

The equilibrium constant for the process is,

$$\begin{aligned} K_{\text{eq}} &= \frac{10^{-\text{pK}_a \text{ of phenol (reactant)}}}{10^{-\text{pK}_a \text{ of water (product)}}} \\ &= \frac{10^{-10}}{10^{-15.7}} \\ &= 5.0 \times 10^5 \end{aligned}$$

The equilibrium constant for proton transfer from phenol to hydroxide ion is greater than 1.

The equilibrium reaction of phenol with hydrogen carbonate ion:



Comparing the $\text{p}K_{\text{a}}$ values of phenol and carbonic acid,

Acid	$\text{p}K_{\text{a}}$
Phenol (Reactant)	10
Carbonic acid (Product)	6.4

The greater the $\text{p}K_{\text{a}}$ value, the weaker the acid. Carbonic acid is a stronger acid than phenol, so the equilibrium lies to the left.

The equilibrium constant for the process,

$$\begin{aligned}
 K_{\text{eq}} &= \frac{10^{-\text{p}K_{\text{a}} \text{ of phenol (reactant)}}}{10^{-\text{p}K_{\text{a}} \text{ of carbonic acid (product)}}} \\
 &= \frac{10^{-10}}{10^{-6.4}} \\
 &= 2.5 \times 10^{-4}
 \end{aligned}$$

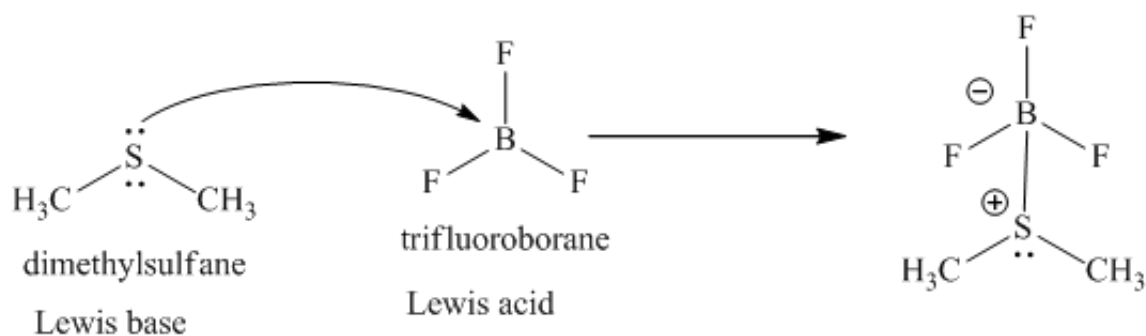
The equilibrium constant for proton transfer from phenol to hydrogen carbonate ion is less than 1.

Problem 39P

Write an equation for the Lewis acid/Lewis base reaction between boron trifluoride and dimethyl sulfide $[(\text{CH}_3)_2\text{S}]$. Use curved arrows to track the flow of electrons and show formal charges if present.

Step-by-step solution

The Lewis acid/Lewis base reaction between boron trifluoride and dimethyl sulfide is as follows:

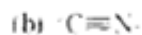
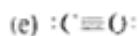
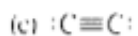


The formal charge on a sulfur atom and a boron atom in a Lewis acid/Lewis base complex are shown as follows.

Atom	Valence electrons of a neutral atom	Electron count	Formal charge
B	3	$\frac{1}{2}(8) = 4$	$(3 - 4) = -1$
S	6	$\frac{1}{2}(6) + 2 = 5$	$(6 - 5) = +1$

Problem 40P

Each of the following species will be encountered at some point in this text. They all have the same number of electrons binding the same number of atoms and the same arrangement of bonds: they are isoelectronic. Specify which atoms, if any, bear a formal charge in the Lewis structure given and the net charge for each species.



Step-by-step solution

step 1 of 6

$$\text{Formal charge} = \text{number of valence electrons} = 3 - \left[\left(\frac{1}{2} \times 6 \right) + 8 \right] = -1$$

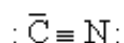
step 2 of 6

(a) Formal charge of each N atom = $5 - \left[\left(\frac{1}{2} \times 6 \right) + 2 \right] = 0$

step 3 of 6

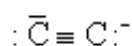
(b) Formal charge on C = $4 - \left[\left(\frac{1}{2} \times 6 \right) + 2 \right] = -1$

Formal charge on N = $5 - \left[\left(\frac{1}{2} \times 6 \right) + 2 \right] = 0$



step 4 of 6

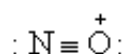
(c) Formal charge on each C atom = $4 - \left[\left(\frac{1}{2} \times 6 \right) + 2 \right] = -1$



step 5 of 6

(d) Formal charge on N = $5 - \left[\left(\frac{1}{2} \times 6 \right) + 2 \right] = 0$

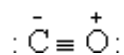
Formal charge on O = $6 - \left[\left(\frac{1}{2} \times 6 \right) + 2 \right] = +1$



step 6 of 6

(e) Formal charge on C = $4 - \left[\left(\frac{1}{2} \times 6 \right) + 2 \right] = -1$

Formal charge on O = $6 - \left[\left(\frac{1}{2} \times 6 \right) + 2 \right] = +1$



Problem 41P

The connectivity of carbon oxysulfide is OCS.

(a) Write a Lewis structure for carbon oxysulfide that satisfies the octet rule.

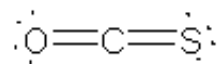
(b) What is the molecular geometry according to VSEPR?

(c) Does carbon oxysulfide have a dipole moment? If so, what is its direction?

Step-by-step solution

step 1 of 3

(a)

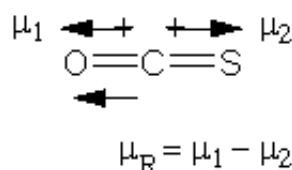


step 2 of 3

(b) The molecule is linear since there is no unshared electron on carbon that is sp hybridized.

step 3 of 3

(c)



Because both O and S are more electronegative than C, the two moments μ_1 and μ_2 act in opposite directions and virtually nullify each other.

Still, the compound shows a much smaller resultant dipole moment (μ_R) to the direction of oxygen, as it is more electronegative than sulfur.

Problem 42P

Write a Lewis structure for each of the following organic molecules:

(a) $\text{C}_2\text{H}_3\text{Cl}$ [vinyl chloride: starting material for the preparation of poly (vinyl chloride), or PVC, plastics]

(b) $\text{C}_2\text{HBrClF}_3$ (halothane: a nonflammable inhalation anesthetic: all three fluorines are bonded to the same carbon)

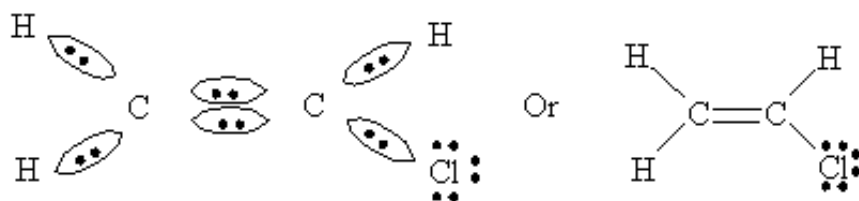
(c) $\text{C}_2\text{Cl}_2\text{F}_4$ (Freon 114: formerly used as a refrigerant and as an aerosol propellant; each carbon bears one chlorine)

Step-by-step solution

step 1 of 3

(a)

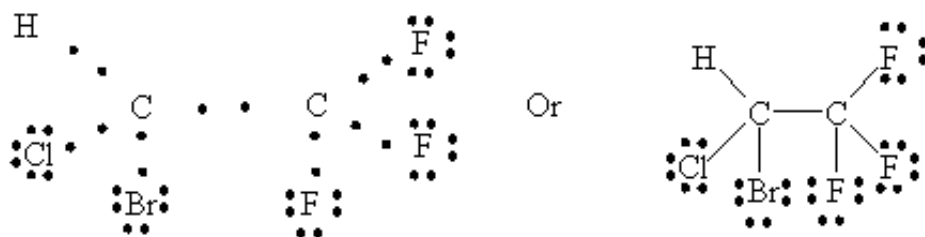
Lewis structure of vinyl chloride $\text{C}_2\text{H}_3\text{Cl}$



step 2 of 3

(b)

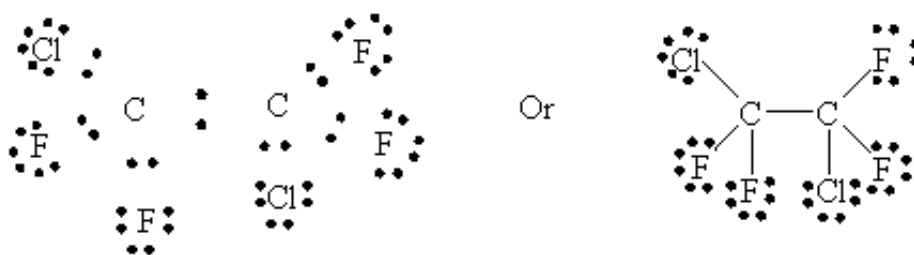
$\text{C}_2\text{HBrClF}_3$ (halothane)



step 3 of 3

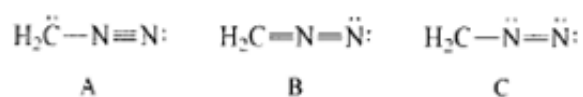
(c)

$\text{C}_2\text{Cl}_2\text{F}_4$ (freon 114)



Problem 43P

Consider Lewis formulas A, B, and C:



(a) Are A, B, and C constitutional isomers, or are they resonance contributors?

(b) Which have a negatively charged carbon?

(c) Which have a positively charged carbon?

(d) Which have a positively charged nitrogen?

(e) Which have a negatively charged nitrogen?

(f) What is the net charge on each?

(g) Which is a more stable structure. A or B? Why?

(h) Which is a more stable structure. B or C? Why?

(i) What is the CNN geometry in each according to VSEPR?

Step-by-step solution

step 1 of 9

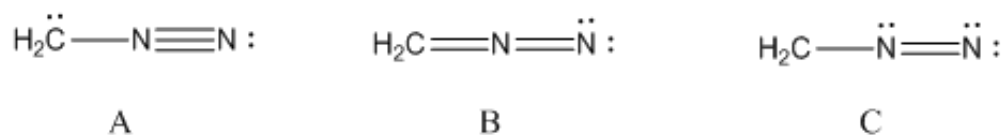
Constitutional isomers are those molecules that have the same molecular formula but differ in the connectivity of the atoms.

Resonance structures are the Lewis structures that describe the fractional bonds and fractional charges in a molecule.

step 2 of 9

(a)

The structures given are as follows:



These structures show the same connectivity of the atoms and are not constitutional isomers.

They are the resonance contributors.

step 3 of 9

The atoms must satisfy their valency. The carbon atom must contain four bonds and the

Nitrogen must have three bonds.

The carbon atom with less/more than four bonds carries a positive charge lone pair of electrons carries a negative charge.

The Nitrogen atom with less/more than three bonds carries a positive charge and with extra pair of electrons carries a negative charge.

step 4 of 9

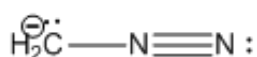
(b)

The Carbon atom in resonance structure A carries a negative charge because it has three bond pairs and one lone pair of electrons.

The Carbon atom of resonance structure B has four bond pairs. So, it carries no charge.

The Carbon atom of resonance structure C has only three bond pairs. So, it carries a positive charge.

Therefore, the Carbon atom of resonance structure A carries a negative charge.



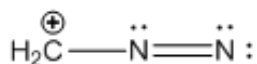
A

step 5 of 9

(c)

The Carbon atom of resonance structure C has only three bond pairs. So, it carries a positive charge.

Therefore, the Carbon atom of resonance structure A carries a positive charge.



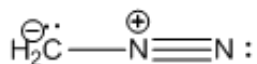
C

step 6 of 9

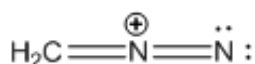
(d)

The middle Nitrogen atoms in resonance structures A and B have four bond pairs which exceed its valency. So, it carries a positive charge.

Therefore, the middle Nitrogen atoms of resonance structures A and B carry a positive charge.



A



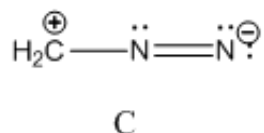
B

step 7 of 9

(e)

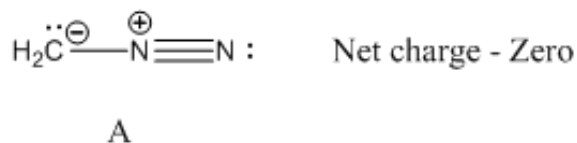
The terminal Nitrogen atom of resonance structure C has two bond pairs and two lone pair of electrons. So, it carries a negative charge.

Therefore, the terminal Nitrogen atom of resonance structure C carries a negative charge.

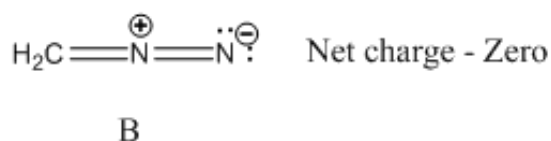


(f)

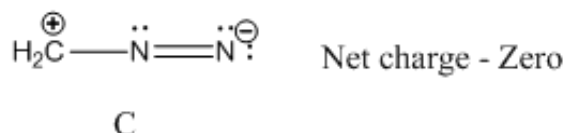
In resonance structure A, Carbon atom carries a positive charge and the middle Nitrogen atom carries a negative charge. Therefore, the net charge of the resonance structure A is zero.



In resonance structure B, the middle Nitrogen carries a positive charge and the terminal Nitrogen carries a negative charge. Therefore, the net charge of the resonance structure B is zero.



In resonance structure C, the Carbon atom carries a positive charge and the terminal Nitrogen carries a negative charge. Therefore, the net charge of the resonance structure C is zero.



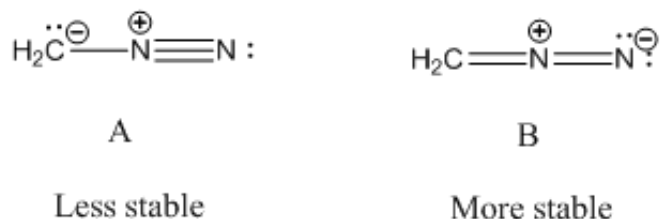
step 8 of 9

(g)

A resonance structure is more stable when there is no charge on any of the atoms (or) when there is less separation of the charges (or) when the negative charge is on the more electronegative element.

Of the resonance structures A and B, the resonance structure B is more stable than A because the negative charge is on the more electronegative atom, Nitrogen in B and there is also less separation of the charges.

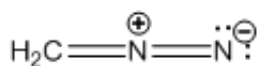
Therefore, resonance structure B is more stable than A.



(h)

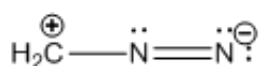
Of the resonance structures B and C, the resonance structure B is more stable than A because there is less separation of the charges.

Therefore, resonance structure B is more stable than C.



B

More stable



C

Less stable

step 9 of 9

(i)

According to VSEPR theory, the middle N atom in structure A is sp hybridized. Therefore, resonance structure A has a linear geometry.

According to VSEPR theory, the middle N atom in structure B is sp hybridized. Therefore, resonance structure B has a linear geometry.

According to VSEPR theory, the middle N atom in structure C is sp^3 hybridized. Therefore, resonance structure C has a tetrahedral geometry.

Problem 44P

In each of the following pairs, determine whether the two represent resonance contributors of a single species or depict different substances. If two structures are not resonance contributors, explain why.



Step-by-step solution

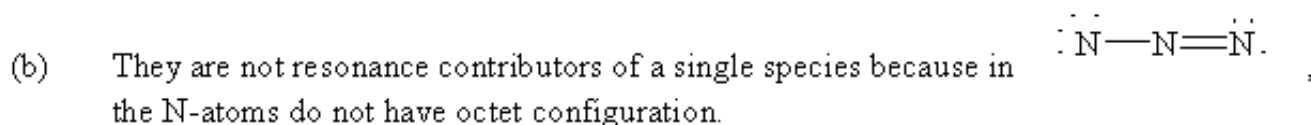
step 1 of 3

(a)



They are resonance contributors of a single species.

step 2 of 3

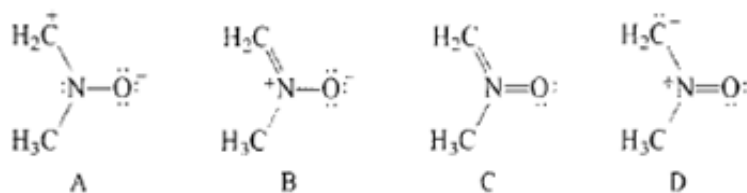


step 3 of 3

- (c) They are not resonance contributors because the numbers of lone pair electrons vary in the structures.

Problem 45P

- (a) Which one of the following is not a permissible contributing structure? Why?



- (b) Rank the three remaining structures in order of their contribution to the resonance hybrid. Explain your reasoning.

- (c) Using curved arrows, show the electron movement that connects the three resonance contributors.

Step-by-step solution

step 1 of 4

- (a)

Among the given structures, A, B, C, and D the structure C is not following octet rule.

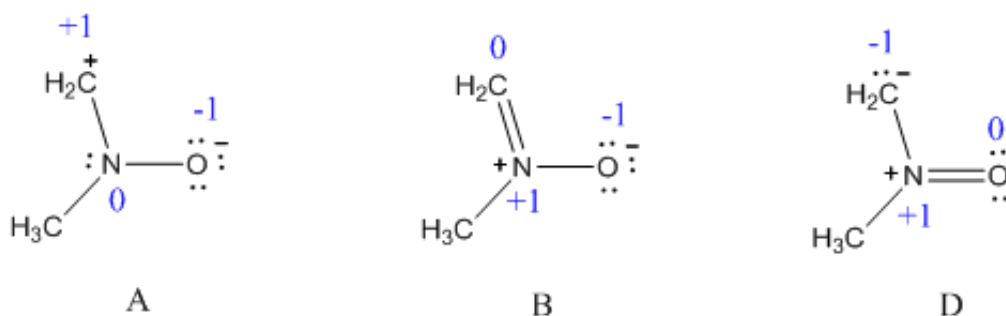
In structure-C, the nitrogen atom is not following the octet rule. The nitrogen atom has ten electrons are surrounded by it.

Thus, the structure C is not a permissible Lewis structure.

step 2 of 4

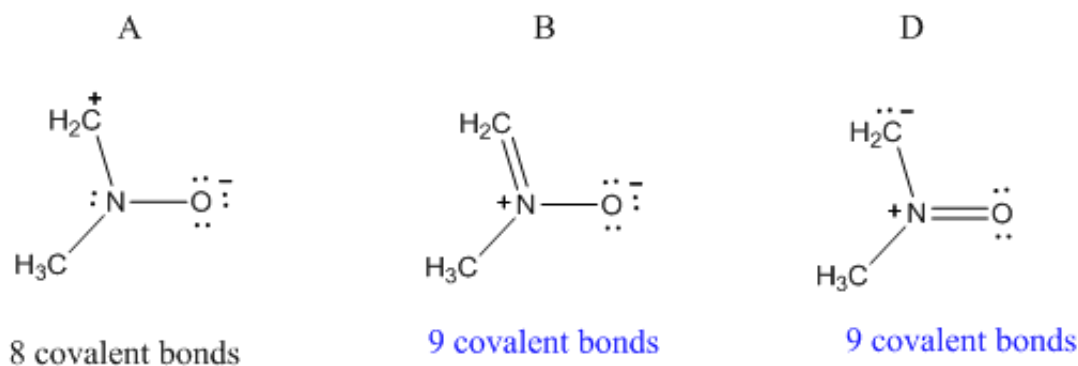
- (b)

The three remaining contributing structures with formal charges are as follows:



All the three resonance structures namely A, B and D are satisfying the octet rule and the net formal charge is zero.

If the contributing structure has more number of covalent bonds, then it contributes more to the resonance hybrid.

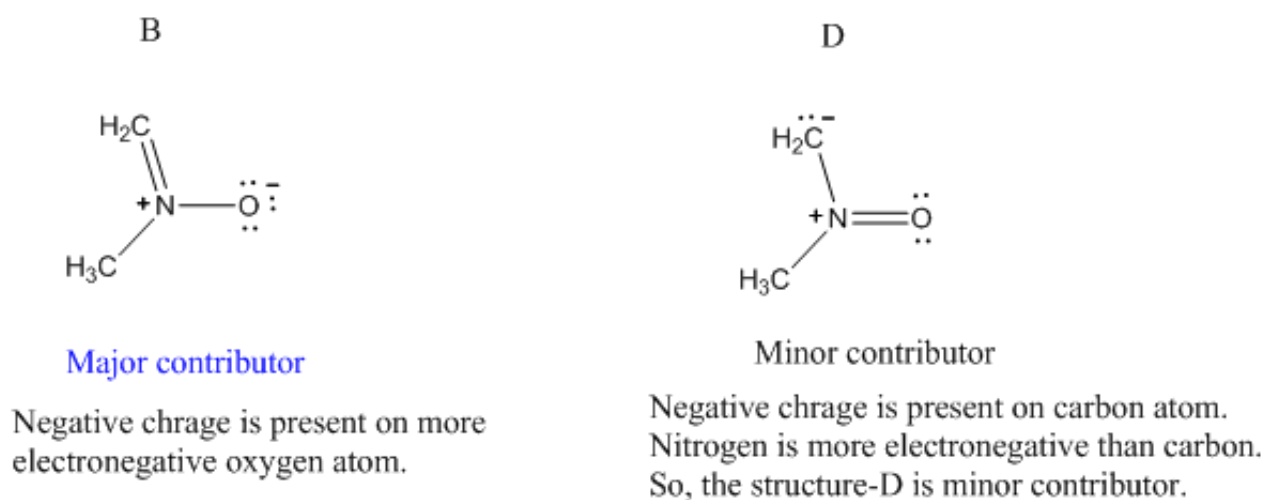


Thus, the structures D and B contribute more to the resonance hybrid than A.

In the contributing structures B and D, the major contributor one in which the negative charge is present on the more electronegative atom.

step 3 of 4

In structure B the negative charge is present on the more electronegative oxygen atom. And in the structure-C, the negative charge is present on less electronegative carbon atom.



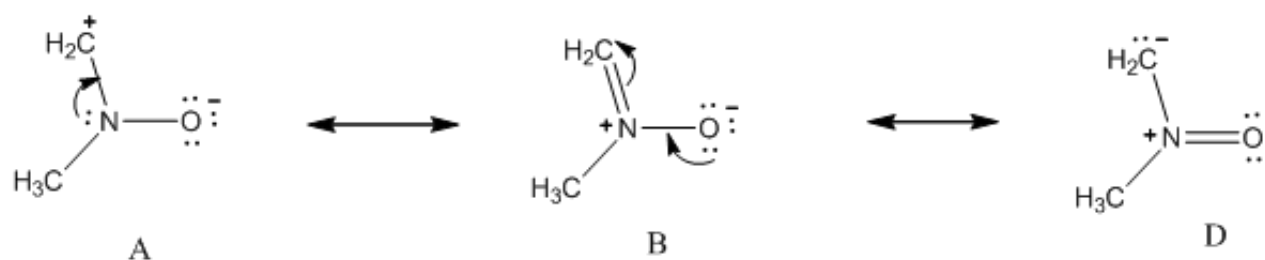
Thus, the structure-B is contributes more toward the resonance hybrid than structure D.

Thus, the order of contributors towards the resonance hybrid is B > D > A.

step 4 of 4

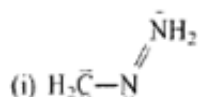
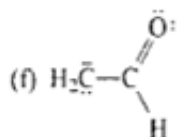
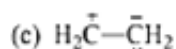
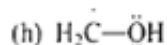
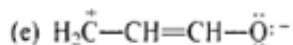
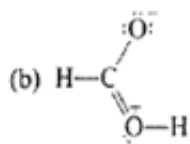
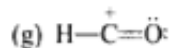
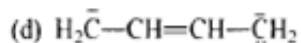
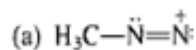
(c)

The curved arrow mechanism for remaining three resonance structures is as follows:



Problem 46P

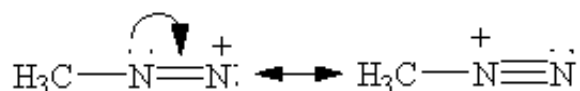
Write a more stable contributing structure for each of the following. Use curved arrows to show how to transform the original Lewis formula to the new one. Be sure to specify formal charges, if any.



Step-by-step solution

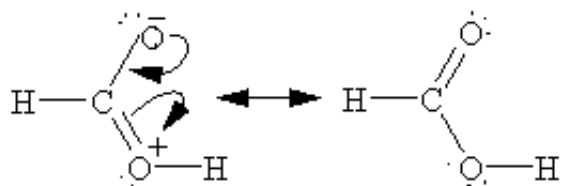
step 1 of 9

(a)



step 2 of 9

(b)



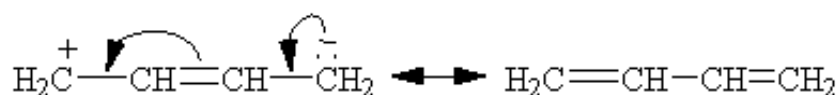
step 3 of 9

(c)



step 4 of 9

(d)



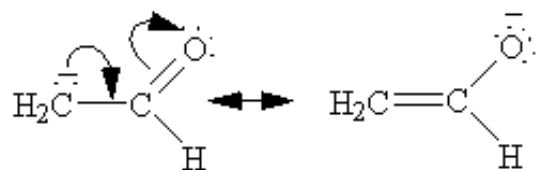
step 5 of 9

(e)



step 6 of 9

(f)



step 7 of 9

(g)



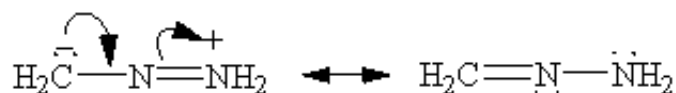
step 8 of 9

(h)



step 9 of 9

(i)



Problem 47P

Write structural formulas for all the constitutionally isomeric compounds having the given molecular formula.

(a) C₄H₁₀

(b) C₅H₁₂

(c) C₂H₄Cl₂

(d) C₄H₉Br

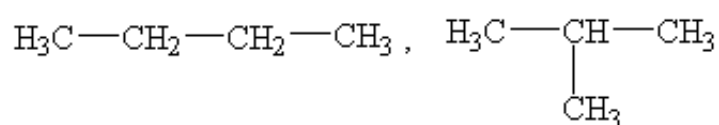
(e) C₃H₉N

Step-by-step solution

step 1 of 5

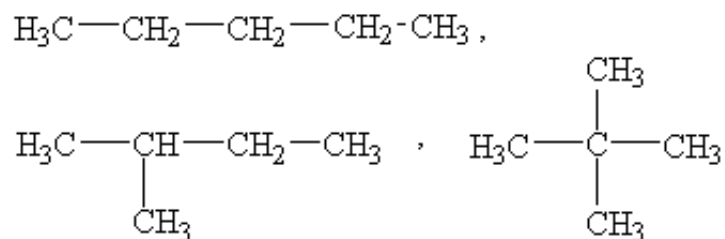
Constitutional isomers have the same molecular formula but differ in the connectivity of the atoms.

(a)



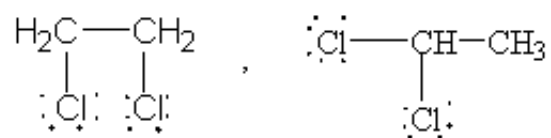
step 2 of 5

(b)



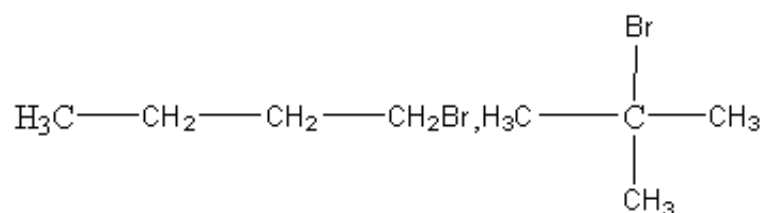
step 3 of 5

(c)



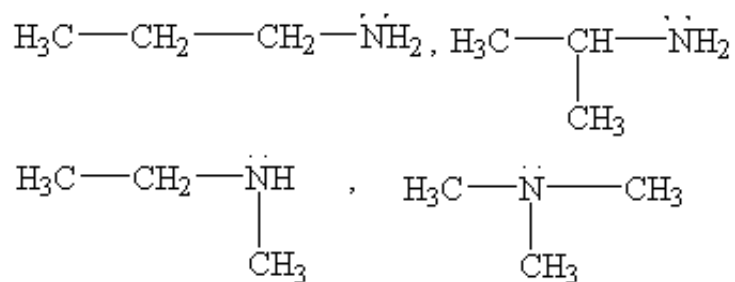
step 4 of 5

(d)



step 5 of 5

(e)



Problem 48P

Write structural formulas for all the constitutional isomers of

(a) C_3H_8

(b) C_3H_6

(c) C_3H_4

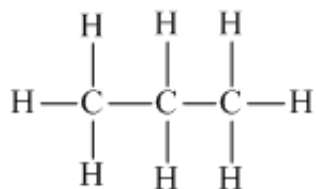
Step-by-step solution

step 1 of 3

Constitutional isomers:

Isomers which are having the same molecular formula but different structures are known as constitutional isomers.

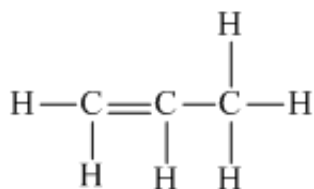
(a) The structural formula for the constitutional isomer of C_3H_8 is shown as follows:



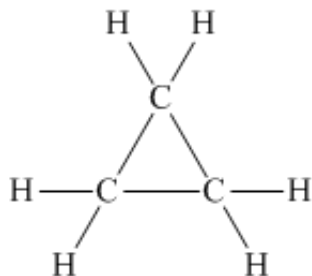
Chemical Formula: C_3H_8

step 2 of 3

(b) The structural formula for the constitutional isomers of C_3H_6 are shown as follows:



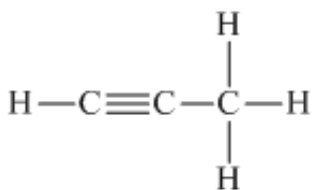
Chemical Formula: C_3H_6



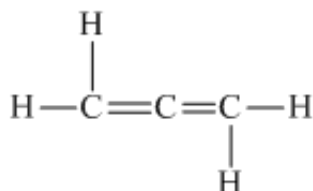
Chemical Formula: C_3H_6

step 3 of 3

(c) The structural formula for the constitutional isomers of C_3H_4 are shown as follows:



Chemical Formula: C_3H_4



Chemical Formula: C_3H_4

Problem 49P

Write structural formulas for all the constitutional isomers of molecular formula $\text{C}_3\text{H}_6\text{O}$ that contain

(a) Only single bonds

(b) One double bond

Step-by-step solution

step 1 of 3

Structural formula is the representation of a molecule in terms of a structure drawn by the

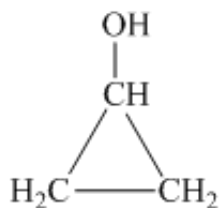
use of simple lines. The branching is also shown in this structure according to the nomenclature of the molecule.

Constitutional isomers are the isomers in which the connectivity of the groups differs in each molecule. The molecules **propanol** and propanone are constitutional isomers.

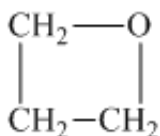
step 2 of 3

(a)

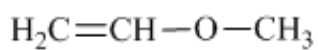
The constitutional isomers of molecule $\text{C}_3\text{H}_6\text{O}$ with only single bonds are as follows:



cyclopropanol



oxacyclobutane

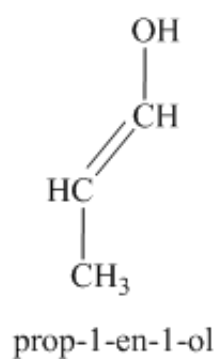
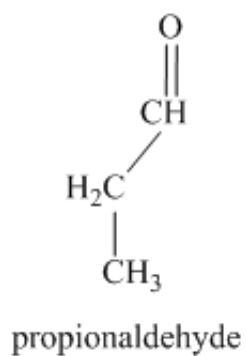
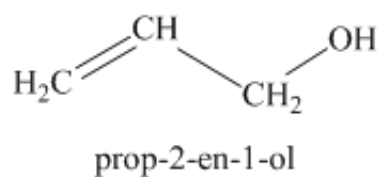
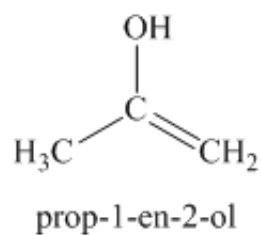
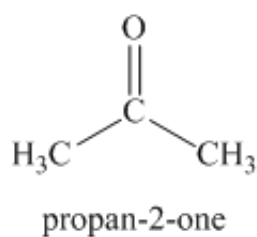


ethenylmethylether

step 3 of 3

(b)

The constitutional isomers of molecule $\text{C}_3\text{H}_6\text{O}$ with only one double bond are as follows:



Problem 50P

For each of the following molecules that contain polar covalent bonds, indicate the positive and negative ends of the dipole, using the symbol \leftrightarrow . Refer to Table 1.3 as needed.

(a) HCl

(c) H₂O

(b) HI

(d) HOCl

TABLE 1.3 Selected Values from the Pauling Electronegativity Scale

Period	Group number						
	1A	2A	3A	4A	5A	6A	7A
1	H 2.1						
2	Li 1.0	Be 1.5	B 2.0	C 2.5	N 3.0	O 3.5	F 4.0
3	Na 0.9	Mg 1.2	Al 1.5	Si 1.8	P 2.1	S 2.5	Cl 3.0
4	K 0.8	Ca 1.0					Br 2.8
5							I 2.5

Step-by-step solution

step 1 of 5

An electronic dipole exists when two dissimilar atoms are connected by a chemical bond. The unequal electronegativity values means that the atom which is more electronegative will attract more of the electron density than the other atom. This builds up a partial negative charge on the atom with higher electronegativity, and a partial positive charge on the atom with lower electronegativity.

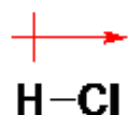
The existence and direction of an electronic dipole is designated by a special red arrow, with the head of the arrow pointing towards the atom possessing higher electronegativity.

For example, consider the bonds in the following chemicals. Indicate the positive and negative ends of each dipole along the polar covalent bonds using the special red arrow.

step 2 of 5

(a)

The first molecule is hydrochloric acid.

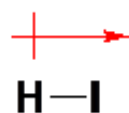


Chlorine has more electronegative than hydrogen, so the net dipole points towards The chlorine atom.

step 3 of 5

(b)

The second molecule is hydroiodic acid.

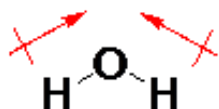


Iodine has more electronegative than hydrogen, thus the net dipole points towards the Iodine atom.

step 4 of 5

(c)

The third molecule is water.

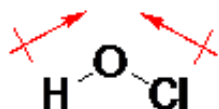


In water molecule, oxygen has more electronegative than hydrogen. So, the dipole of Each $\text{O}-\text{H}$ bond points towards the oxygen atom.

step 5 of 5

(d)

The fourth and final molecule is hypochlorous acid.



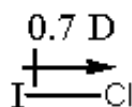
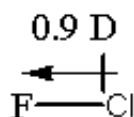
Since oxygen has more electronegative than either chlorine or hydrogen. Therefore, the Dipole of each bond points towards the oxygen atom.

Problem 51P

The compounds FCl and ICl have dipole moments μ that are similar in magnitude (0.9 and 0.7 D, respectively) but opposite in direction. In one compound, chlorine is the positive end of the dipole; in the other it is the negative end. Specify the direction of the dipole moment in each compound, and explain your reasoning.

Step-by-step solution

step 1 of 1



The direction of dipole moment depends on electronegativity. The dipole moment of a bond is directed towards the electronegative atom in the molecule.

In the molecule FCl, Fluorine is more electronegative than Cl, so the dipole moment is directed towards F. In ICl, Cl is more electronegative than I, so the dipole moment is directed towards Cl. Hence, the dipole moments in both the compounds are opposite in direction.

Problem 52P

Which compound in each of the following pairs would you expect to have the greater dipole moment μ ? Why?

- | | |
|---|--|
| (a) HF or HCl | (d) CHCl_3 or CCl_3F |
| (b) HF or BF_3 | (e) CH_3NH_2 or CH_3OH |
| (c) $(\text{CH}_3)_3\text{CH}$ or $(\text{CH}_3)_3\text{CCl}$ | (f) CH_3NH_2 or CH_3NO_2 |

Step-by-step solution

step 1 of 7

An electronic dipole exists whenever two dissimilar atoms are connected by a chemical bond. The unequal electronegativity values means that the atom which is more electronegative will attract more of the electron density than the other atom. This builds up a partial negative charge on the atom with higher electronegativity, and a partial positive charge on the atom with lower electronegativity.

The existence and direction of an electronic dipole is designated by a special red arrow, with the head of the arrow pointing towards the atom possessing higher electronegativity.

There is often more than one electronic dipole present in a molecule. To find the overall dipole (the net contribution of all the individual dipoles, a value called the dipole moment), you have to consider all of the individual dipoles as well as their relative magnitude. Which direction are they pulling, and how large is the difference in electronegativity for each? Consider the following examples.

step 2 of 7

(a)

Which compound would you expect to have the higher dipole moment, and why: HF or HCl?

Since fluorine is the most electronegative element (stronger than chlorine), HF would have a higher dipole moment than HCl as the bond in HF would be more polar.

step 3 of 7

(b)

Which compound would you expect to have the higher dipole moment, and why: HF or boron trifluoride?

Recall that boron trifluoride is planar; this is due to boron's position on the Periodic Table, which does not require it to possess any lone pairs in boron trifluoride in order to be stable. Each of the B-F bonds therefore points to a different intersection in the same triangle, meaning they all pull against each other and effectively cancel each other out. Hydrofluoric acid is very polar due to the electronegativity of fluorine. HF would have a higher dipole moment than boron trifluoride.

step 4 of 7

(c)

Which compound would you expect to have the higher dipole moment, and why: $(\text{CH}_3)_3\text{CH}$ or $(\text{CH}_3)_3\text{CCl}$?

A carbon-carbon bond is only weakly polar (via hyperconjugation effects). A carbon-chlorine bond, however, is highly polarized due to the electronegativity of chlorine. The $(\text{CH}_3)_3\text{CCl}$ molecule would therefore have a higher dipole moment than $(\text{CH}_3)_3\text{CH}$.

step 5 of 7

(d)

Which compound would you expect to have the higher dipole moment, and why: CHCl_3 or CCl_3F ?

A carbon-hydrogen bond reinforces the polarizing effect of the chlorines, as the C-H bond is polarized in favor of the carbon atom. However, the C-F bond fights the three C-Cl bonds in the tetrahedral compound, which reduces the overall effect of the three C-Cl dipoles.

Therefore, CHCl_3 would have a higher dipole moment than CCl_3F , even though the second compound contains highly electronegative fluorine.

step 6 of 7

(e)

Which compound would you expect to have the higher dipole moment, and why: CH_3NH_2 or CH_3OH ?

Nitrogen is less electronegative than oxygen, meaning that the methanol molecule will have a higher overall dipole moment than methylamine.

step 7 of 7

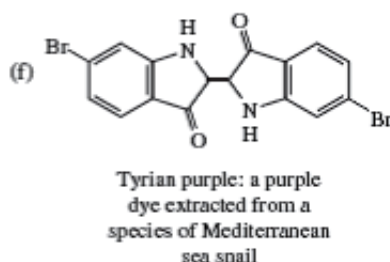
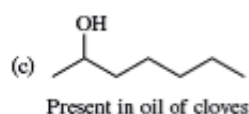
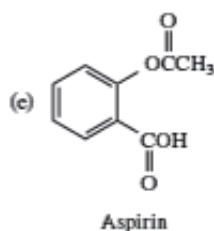
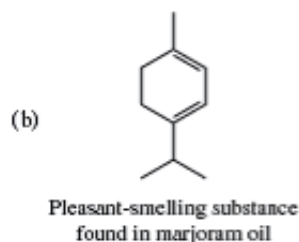
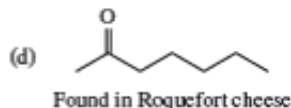
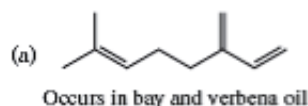
(f)

Which compound would you expect to have the higher dipole moment, and why: CH_3NH_2 or CH_3NO_2 ?

When you compare the two nitrogen atoms, it's easy to see that the nitromethane has a formal charge of +1 on its nitrogen (compared to the methylamine nitrogen, which is neutral). This positive charge makes the nitromethane nitrogen very electron-attracting, as it attempts to neutralize the positive charge. Nitromethane therefore has a higher overall dipole moment than methylamine.

Problem 53P

Expand the following structural representations so as to more clearly show all the atoms and any unshared electron pairs. What are their molecular formulas? Are any of them isomers?



Step-by-step solution

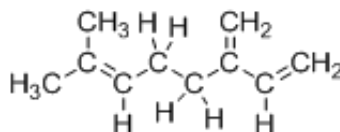
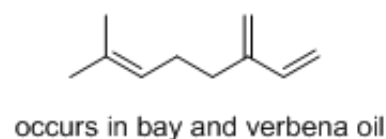
step 1 of 6

Structural representations can be condensed in order to make them faster to draw. However, sometimes it is helpful to see the full, expanded structure (showing all atoms and all lone pairs of electrons). Consider the following examples.

step 2 of 6

(a)

The first molecule is an alkene. We can expand the structure to show all the atoms and lone pairs of electrons.



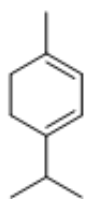
Therefore, the molecular formula of this compound is $\boxed{\text{C}_{10}\text{H}_{16}}$.

step 3 of 6

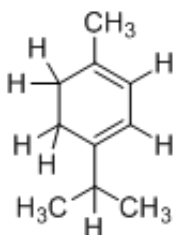
(b)

The second molecular is a cyclic alkene. We can expand the structure to show all the atoms

and lone pairs of electrons.



Pleasant-smelling substance
found in marjoram oil

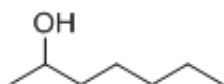


Therefore, the molecular formula of this compound is $\boxed{\text{C}_{10}\text{H}_{16}}$.

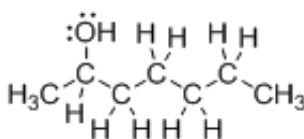
step 4 of 6

(c)

The third molecule is an alcohol. We can expand the structure to show all the atoms and lone pairs of electrons.



Present in oil of cloves

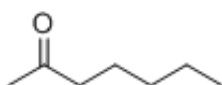


Therefore, the molecular formula of this compound is $\boxed{\text{C}_8\text{H}_{18}\text{O}}$.

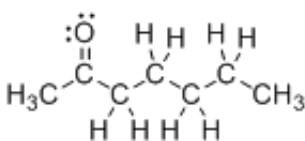
step 5 of 6

(d)

The fourth molecule is a ketone. We can expand the structure to show all the atoms and lone pairs of electrons.



Found in Roquefort cheese

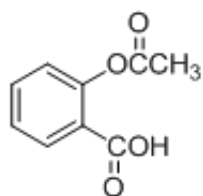


Therefore, the molecular formula of this compound is $\boxed{\text{C}_8\text{H}_{16}\text{O}}$.

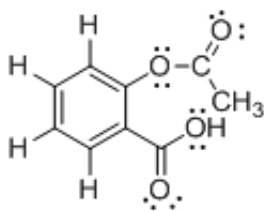
step 6 of 6

(e)

The fifth molecule is an aromatic compound. We can expand the structure to show all the atoms and lone pairs of electrons.



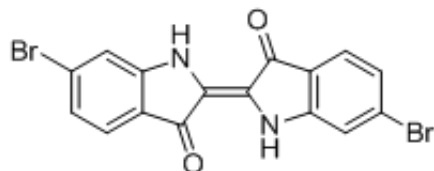
Aspirin



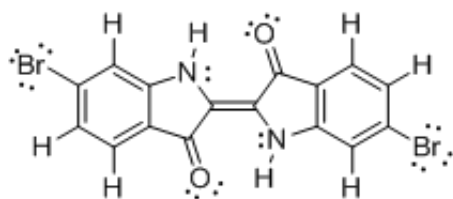
Therefore, the molecular formula of this compound is $\boxed{\text{C}_9\text{H}_8\text{O}_4}$.

(f)

The sixth molecule is a fused aromatic compound. We can expand the structure to show all the atoms and lone pairs of electrons.



Tyrian purple: a purple dye extracted from a species of Mediterranean sea snail

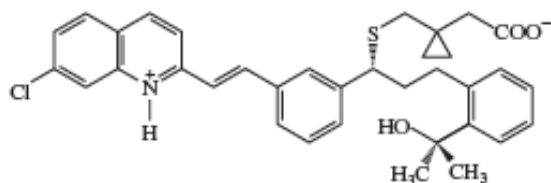


The molecular formula of this compound is $\text{C}_{16}\text{H}_8\text{Br}_2\text{N}_2\text{O}_2$.

Note that the first two examples in this set are constitutional isomers.

Problem 54P

The structure of montelukast, an antiasthma drug, is shown here.



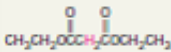
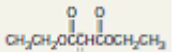












(a) Use Table 1.8 to identify the most acidic and most basic sites in the molecule. (Although you won't find an exact match in structure, make a prediction based on analogy with similar groups in simpler molecules.)

(b) Write the structure of the product formed by treating montelukast with one equivalent of sodium hydroxide.

(c) Write the structure of the product formed by treating

montelukast with one equivalent of HCl.

TABLE 1.8 Acidity Constants (pK_a) of Acids (Continued)				
Acid	pK_a	Formula	Conjugate base	Discussed in section
Piperidinium ion	11.2			21.4
Diethyl malonate	13			20.2
Methanol	15.2	CH_3OH	CH_3O^-	1.14; 14.4
2-Methylpropanal	15.5			20.2
Water*	15.7	H_2O	HO^-	1.14
Ethanol	16	CH_3CH_2OH	$CH_3CH_2O^-$	1.14
Cyclopentadiene	16			11.20
Isopropyl alcohol	17	$(CH_3)_2CHOH$	$(CH_3)_2CHO^-$	1.14
tert-Butyl alcohol	18	$(CH_3)_3COH$	$(CH_3)_3CO^-$	1.14
Acetone	19			20.2
Ethyl acetate	25.6			20.2
Acetylene	26	$HC\equiv CH$	$HC\equiv C^-$	9.5
Hydrogen	35	H_2	H^-	20.8
Ammonia	36	NH_3	H_2N^-	1.14; 1.15; 14.4
Diisopropylamine	36	$[(CH_3)_2CH]_2NH$	$[(CH_3)_2CH]_2N^-$	20.2
Benzene	43			14.4; 14.5
Ethylene	45	$H_2C=CH_2$	$H_2C=\dot{C}H$	9.4; 9.5
Methane	60	CH_4	$\dot{C}H_3$	1.14; 9.4; 9.5; 14.4
Ethane	62	CH_3CH_3	$CH_3\dot{C}H_2$	14.5

Step-by-step solution

step 1 of 5

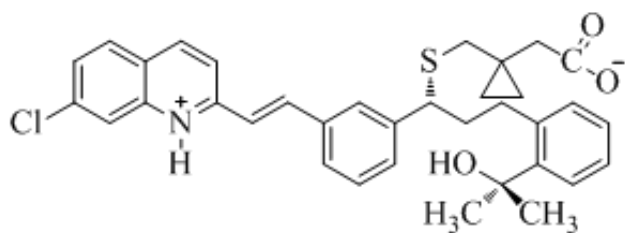
The conjugate base is the molecule formed after deprotonation of an acid. Stronger is the acid; weaker is its conjugate base. An acid reacts with a base to form its deprotonated form. Conjugate acid is the molecule formed after protonation of a base. Stronger is the base; weaker is its conjugate acid. A base reacts with an acid to form its protonated form.

Zwitter ion is the molecule in which both positively and negatively charged ions are in same molecule. The reactivity of Zwitter ion with acid and base depends on these two ionic sites. Drug like montelukast exist in Zwitter ion form. The quantity used to measure acidity

numerically is pK_a . Higher the value of pK_a , less is the acidity and lower values means higher acidity.

step 2 of 5

The structure of montelukast is as follows:



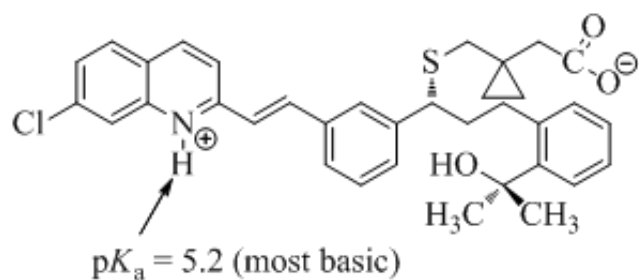
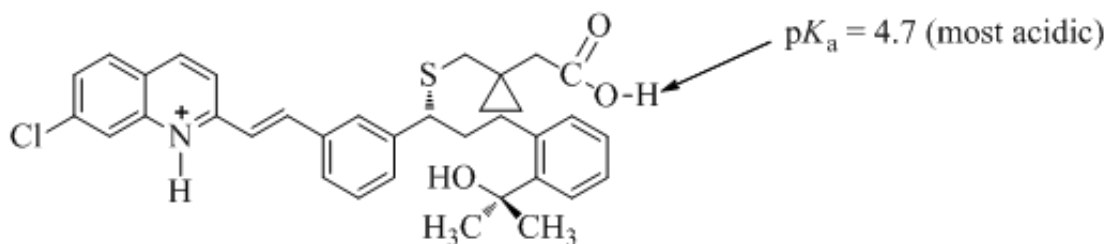
This molecule contains numerous acidic and basic sites.

step 3 of 5

Refer Table 1.8 to find conjugate base and pK_a of each species.

(a)

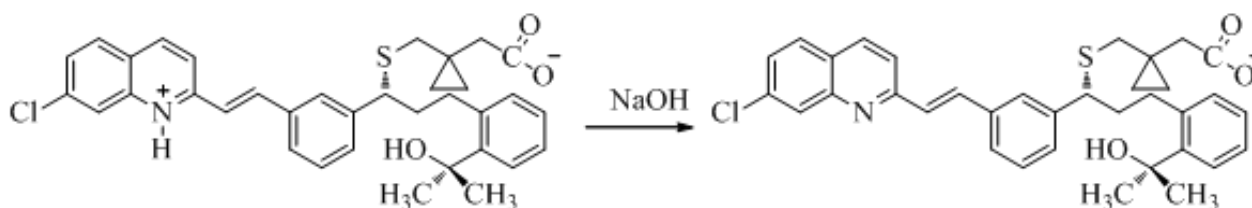
This carboxyl group present in the molecule has a similar structure to that of acetic acid, which means that its pK_a value should be similar to that of acetic acid. The most basic site in montelukast can be predicted by comparing the molecule to pyridinium ion. The positively-charged nitrogen should have a pK_a similar to that of pyridinium ion. The acidic and basic sites are represented as follows:



step 4 of 5

(b)

The reaction of molecule with one equivalent sodium hydroxide is as follows:



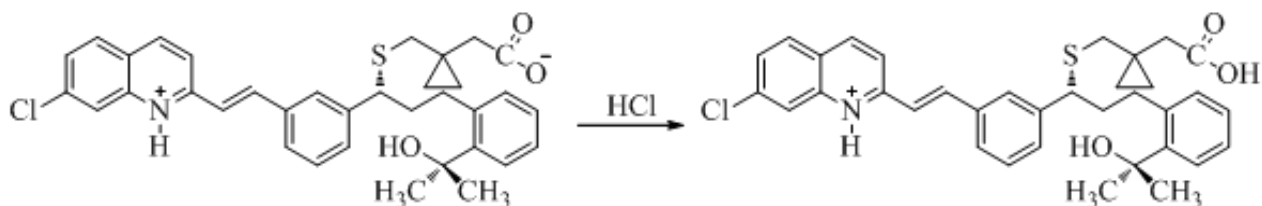
Sodium hydroxide is a strong base, and can easily deprotonate the positively-charged nitrogen atom. This leads to a molecule in which the carboxylic acid remains in the carboxylate form,

but the aromatic nitrogen is no longer protonated.

step 5 of 5

(c)

The reaction of molecule with one equivalent hydrochloric acid is as follows:



Hydrochloric acid is a strong acid, and can easily protonate the negatively-charged carboxylate oxygen. This leads to a molecule containing a neutral carboxylic acid along with a protonated (positively-charged) nitrogen atom.

Problem 55P

(a) One acid has a pKa of 2, the other has a pKa of 8. What is the ratio of their Ka's ?

(b) Two acids differ by 10,000 in their Ka's. If the pKa of the weaker acid is 5, what is the pKa of the stronger acid?

Step-by-step solution

step 1 of 3

(a) We know that, $K_a = 10^{-\text{p}K_a}$

Therefore, $(K_{a1}) = 10^{-(\text{p}K_{a1})}$

$$(\text{p}K_a)_1 = 2$$

$$(K_{a1}) = 10^{-2}$$

$$(K_{a2}) = 10^{-(\text{p}K_{a2})}$$

$$(\text{p}K_{a2}) = 8$$

$$(K_{a2}) = 10^{-8}$$

step 2 of 3

Now,

$$\frac{(K_{a1})}{(K_{a2})} = \frac{10^{-2}}{10^{-8}}$$

$$\boxed{\frac{(K_{a1})}{(K_{a2})} = 10^6}$$

step 3 of 3

(b) Given that –

$(K_{a1}) - (K_{a2}) = 10,000$ where K_{a1} and K_{a2} are the K_a values
of strong and weak acids respectively.

$$(pK_{a2}) = 5$$

$$\text{or, } (K_{a2}) = 10^{-(pK_{a2})}$$

$$\text{or, } (K_{a2}) = 10^{-5}$$

$$\text{Therefore, } (K_{a1}) = 10000 + (K_{a2})$$

$$\text{or, } (K_{a1}) = 10000 + 10^{-5}$$

$$\text{or, } (K_{a1}) = 10000.00001$$

$$\text{or, } \boxed{(pK_{a1}) = -4}$$

Problem 56P

Calculate K_a for each of the following acids, given its pK_a . Rank the compounds in order of decreasing acidity.

(a) Aspirin: $pK_a = 3.48$

(b) Vitamin C (ascorbic acid): $pK_a = 4.17$

(c) Formic acid (present in sting of ants): $pK_a = 3.75$

(d) Oxalic acid (poisonous substance found in certain berries): $pK_a = 1.19$

Step-by-step solution

step 1 of 5

(a)

The relationship between pK_a and K_a is $pK_a = -\log_{10} K_a$.

The $\text{p}K_a$ of aspirin is 3.48.

Calculate $\text{p}K_a$ of aspirin as follows:

$$\text{p}K_a = -\log_{10} K_a$$

$$3.48 = -\log_{10} K_a$$

$$K_a = 10^{-3.48}$$

$$= 3.31 \times 10^{-4}$$

Therefore, K_a of aspirin is $\boxed{3.31 \times 10^{-4}}$.

step 2 of 5

(b)

The $\text{p}K_a$ of ascorbic acid is 4.17.

Calculate $\text{p}K_a$ of ascorbic as follows:

$$\text{p}K_a = -\log_{10} K_a$$

$$4.17 = -\log_{10} K_a$$

$$K_a = 10^{-4.17}$$

$$= 6.76 \times 10^{-5}$$

Therefore, K_a of ascorbic acid is $\boxed{6.76 \times 10^{-5}}$.

(c)

step 3 of 5

The $\text{p}K_a$ of formic acid is 3.75

Calculate $\text{p}K_a$ of formic acid as follows:

$$\text{p}K_a = -\log_{10} K_a$$

$$3.75 = -\log_{10} K_a$$

$$K_a = 10^{-3.75}$$

$$= 1.78 \times 10^{-4}$$

Therefore, K_a of formic acid is $\boxed{1.78 \times 10^{-4}}$.

step 4 of 5

(d)

The $\text{p}K_a$ of oxalic acid is 1.19

Calculate $\text{p}K_a$ of oxalic acid as follows:

$$\text{p}K_a = -\log_{10} K_a$$

$$1.19 = -\log_{10} K_a$$

$$K_a = 10^{-1.19}$$

$$= 6.46 \times 10^{-2}$$

Therefore, K_a of oxalic acid is 6.46×10^{-2} .

step 5 of 5

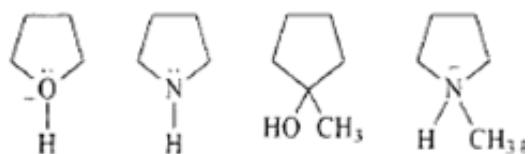
From determination of the above K_a values, that the greater the $\text{p}K_a$ values of the acid, the lower its K_a values. Higher $\text{p}K_a$ of substance, lower in acidic nature. Therefore, decreasing order of acidity of given compounds are as follows:

Oxalic acid > aspirin > formic acid > vitamin C

$$\text{p}K_a = 1.19 \quad \text{p}K_a = 3.48 \quad \text{p}K_a = 3.75 \quad \text{p}K_a = 4.17$$

Problem 57P

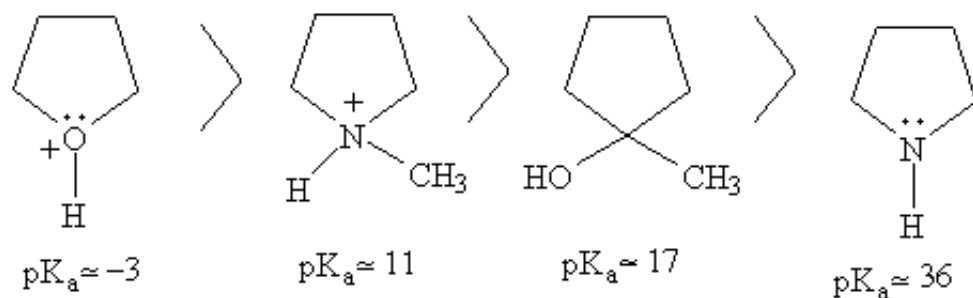
Rank the following in order of decreasing acidity. Although none of these specific structures appear in Table 1.8, you can use analogous structures in the table to guide your reasoning.



Step-by-step solution

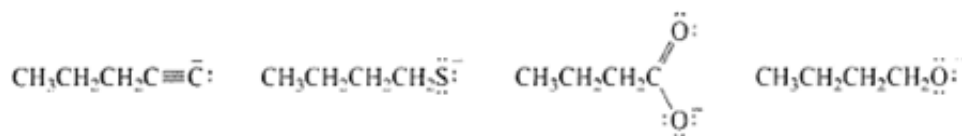
step 1 of 1

The acid strength depends upon the lower pK_a value.



Problem 58P

Rank the following in order of decreasing basicity. As in the preceding problem, Table 1.8 should prove helpful.



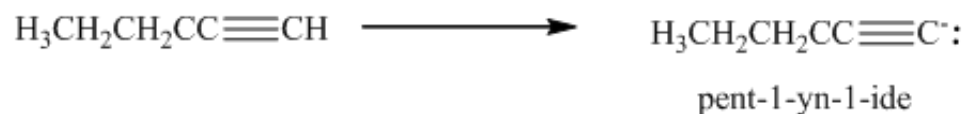
Step-by-step solution

step 1 of 4

The weaker the acid, the stronger is the conjugate base. So, a higher pK_a value of acids corresponds to a stronger basic character.

A negative charge on the more electronegative element in the conjugate base of an acid, the more stable it is and more acidic is the parent acid.

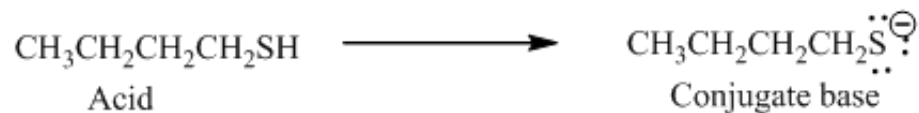
The pentyn-1-ide ion is the conjugate base for 1-pentyne.



The negative charge is on the less electronegative element, Carbon.

step 2 of 4

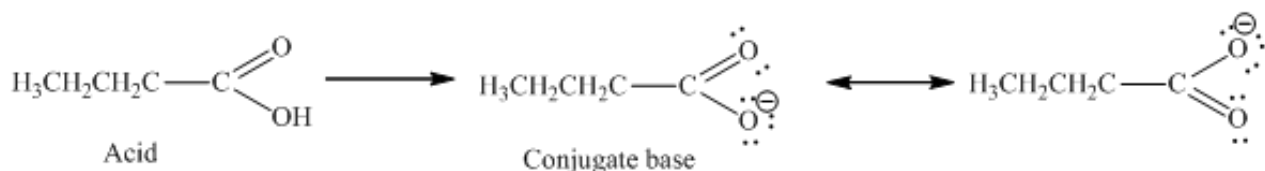
The conjugate base is formed by the loss of proton from 1-butane-thiol.



The negative charge is on the more electronegative element, Sulfur.

step 3 of 4

The conjugate base is formed from the carboxylic acid.

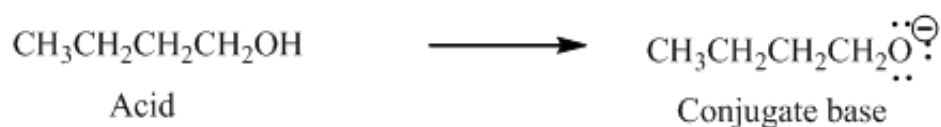


The negative charge in the conjugate base is on the more electronegative element, Oxygen atom.

In addition, the conjugate base exhibits resonance structures.

Therefore, the parent ion is more acidic.

The conjugate base is formed from weak acid, alcohols.



The negative charge in the conjugate base is on the more electronegative element, Oxygen atom.

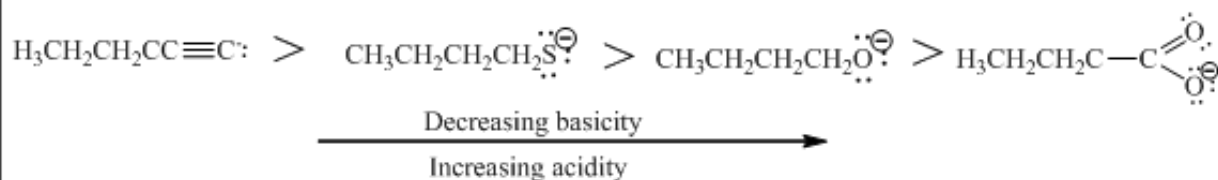
step 4 of 4

Comparing the conjugate bases of all the compounds, the conjugate base with the negative charge on the more electronegative element is more stable and the parent acid is more acidic.

In simple words, negative charge on more electronegative element with more resonance structures is more acidic and less basic.

Negative charge on less electronegative element is less acidic and more basic.

Therefore, the decreasing orders of the basicity of the conjugate bases are as follows:



Problem 59P

Consider 1.0 M aqueous solutions of each of the following. Which solution is more basic?

(a) Sodium cyanide (NaCN) or sodium fluoride (NaF)

(b) Sodium carbonate (Na₂CO₃) or sodium acetate $\begin{array}{c} \text{O} \\ \text{//} \\ (\text{CH}_3\text{C})\text{ONa} \end{array}$

(c) Sodium sulfate (Na₂SO₄) or sodium methanethiolate (NaSCH₃)

Step-by-step solution

step 1 of 3

(a) NaCN contains the ions Na^+ and CN^- . NaF contains the ions Na^+ and F^- .

CN^- is the conjugate base of HCN and F^- is the conjugate base of HF.

Base	CN^-	F^-
Conjugate acid	HCN	HF
pK_a of the conjugate acid	9.1	3.1

The conjugate acid of F^- is stronger than CN^- , so CN^- is a stronger base than F^- . Therefore, NaCN is more basic than NaF.

step 2 of 3

(b) Na_2CO_3 contains Na^+ and CO_3^{2-} ions. $\text{CH}_3\text{CO}_2\text{Na}$ contains Na^+ and CH_3CO_2^- ions.

CO_3^{2-} is the conjugate base of HCO_3^- and CH_3CO_2^- is the conjugate base of $\text{CH}_3\text{CO}_2\text{H}$.

Base	CO_3^{2-}	CH_3CO_2^-
Conjugate acid	HCO_3^-	$\text{CH}_3\text{CO}_2\text{H}$
pK_a of the conjugate acid	10.2	4.7

The conjugate acid of CH_3CO_2^- is stronger than that of HCO_3^- , so CO_3^{2-} is a stronger base than CH_3CO_2^- . Therefore, Na_2CO_3 is more basic than $\text{CH}_3\text{CO}_2\text{Na}$.

step 3 of 3

(c) Na_2SO_4 contains the ions Na^+ and SO_4^{2-} . NaSCH_3 contains the ions Na^+ and CH_3S^- .

SO_4^{2-} is the conjugate base of HSO_4^- and CH_3S^- is the conjugate base of CH_3SH .

Base	SO_4^{2-}	CH_3S^-
Conjugate acid	HSO_4^-	CH_3SH
pK_a of the conjugate acid	2.0	10.7

The conjugate acid of SO_4^{2-} is stronger than that of CH_3S^- , so CH_3S^- is a stronger base than SO_4^{2-} . Therefore, the $\boxed{\text{NaSCH}_3}$ solution is more basic than Na_2SO_4 .

Problem 60P

(a) Which is the stronger acid: $(\text{CH}_3)_3\text{NH}^+$ or $(\text{CH}_3)_3\text{PH}^+$?

(b) Which is the stronger base: $(\text{CH}_3)_3\text{N:}$ or $(\text{CH}_3)_3\text{P:}$?

Step-by-step solution

step 1 of 2

(a) Between $(\text{CH}_3)_3\text{NH}^+$ and $(\text{CH}_3)_3\text{PH}^+$, $\boxed{(\text{CH}_3)_3\text{PH}^+}$ is a stronger acid. This is because the ionizable proton is bonded to P in $(\text{CH}_3)_3\text{PH}^+$ and to N in $(\text{CH}_3)_3\text{NH}^+$. N and P are in the same group. The bond strength decreases down the group due to an increase in the atomic radius. Hence, the hydrogen in $(\text{CH}_3)_3\text{PH}^+$ is more easily ionizable than in $(\text{CH}_3)_3\text{NH}^+$ because of the lower bond strength of P-H.

step 2 of 2

(b) $\boxed{(\text{CH}_3)_3\ddot{\text{N}}}$ is a stronger base than $(\text{CH}_3)_3\ddot{\text{P}}$ because of the smaller size of N compared to P. Nitrogen is in the second period and the presence of three methyl groups (+I groups) would make the donation of the lone pair of electrons easily. On the other hand, P is in the third period. It has a d-orbital, so the lone pair of electrons can be engaged in increasing the valence, making it less available for donation.

Problem 61P

Write an equation for the Brønsted-Lowry acid-base reaction

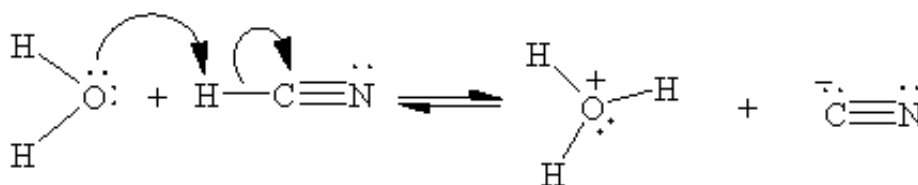
that occurs when each of the following acids reacts with water. Show all unshared electron pairs and formal charges, and use curved arrows to track electron movement.



Step-by-step solution

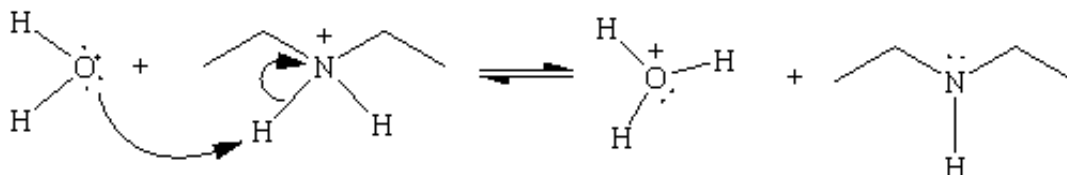
step 1 of 3

(a)



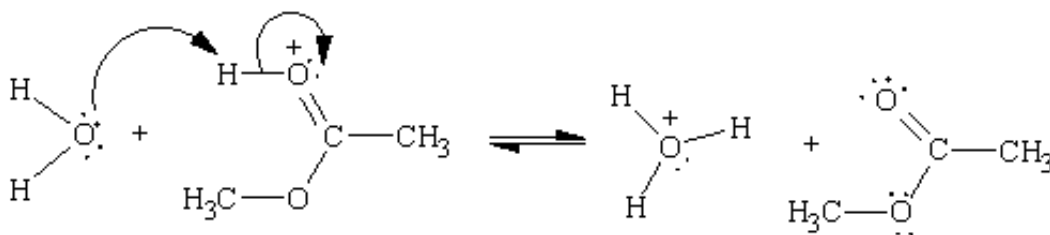
step 2 of 3

(b)



step 3 of 3

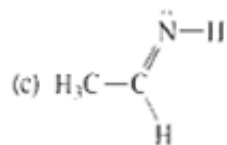
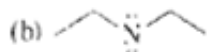
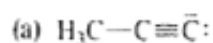
(c)



Problem 62P

Write an equation for the Brønsted-Lowry acid-base reaction that occurs when each of the following bases reacts with water. Show all unshared electron pairs and formal charges, and use curved arrows to track electron

movement.



Step-by-step solution

step 1 of 3

(a)

The anion is negatively charged and hence acts as a base. This upon reaction with water abstracts a proton from water and forms the following compound.

The reaction is as follows:



Formal charge for **C** = -1

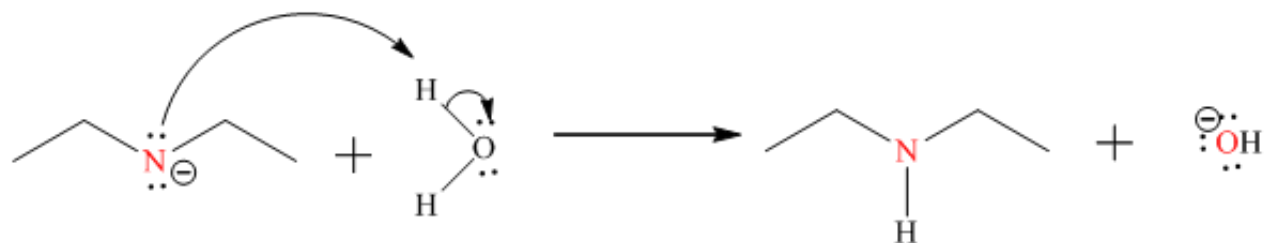
Formal charge for **C** = 0
Formal charge for **O** = -1

step 2 of 3

(b)

The anion is negatively charged and hence acts as a base. This upon reaction with water abstracts a proton from water and forms the following compound.

The reaction is as follows:



Formal charge for **N** = -1

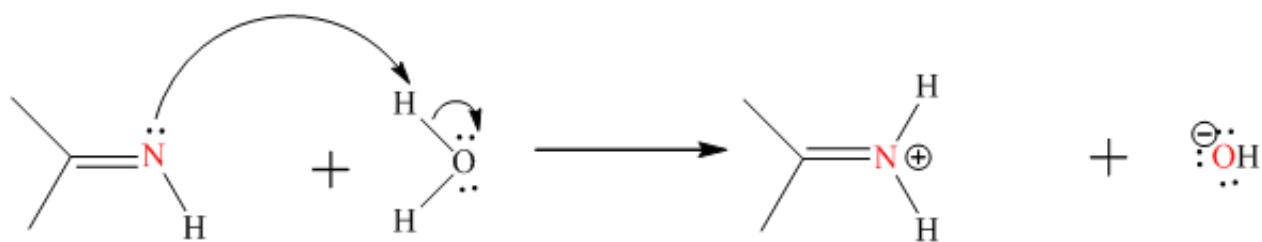
Formal charge for **N** = 0
Formal charge for **O** = -1

step 3 of 3

(c)

The anion is negatively charged and hence acts as a base. This upon reaction with water abstracts a proton from water and forms the following compound.

The reaction is as follows:



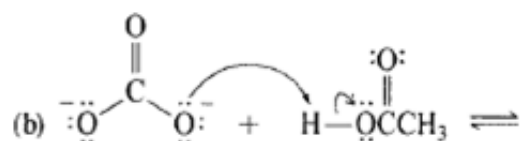
Formal charge for **N** = 0

Formal charge for **N** = +1

Formal charge for **O** = -1

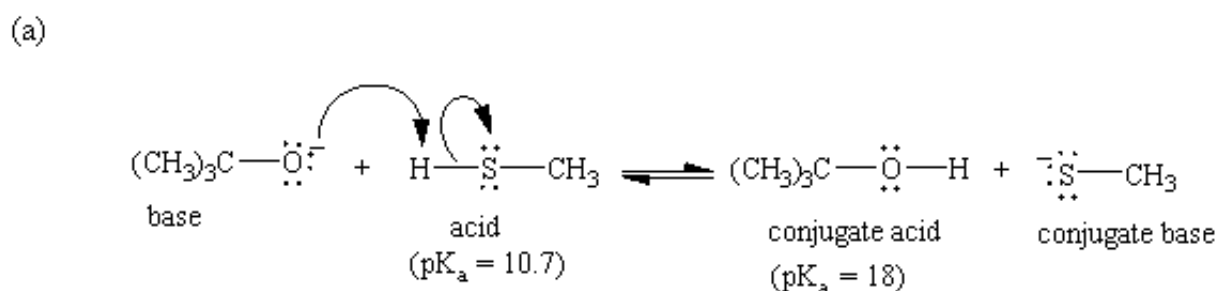
Problem 63P

All of the substances shown in the following acid-base reactions are found in Table 1.8, and the equilibrium lies to the right in each case. Following the curved arrows, complete each equation to show the products formed. Identify the acid, base, conjugate acid, and conjugate base. Calculate the equilibrium constant for each reaction.



Step-by-step solution

step 1 of 3



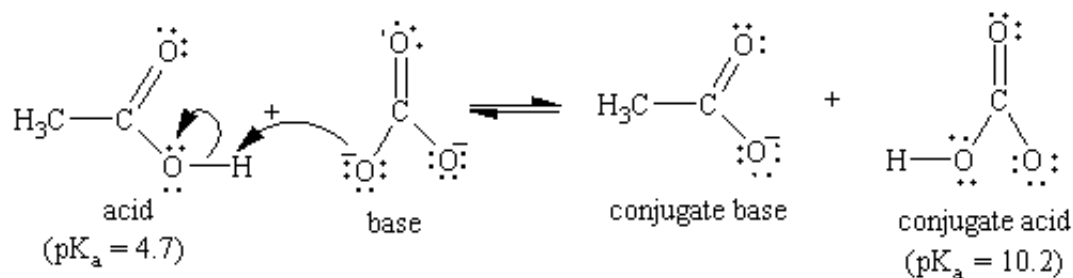
$$K_{\text{eq}} = \frac{10^{-\text{pK}_a \text{ of reactant acid}}}{10^{-\text{pK}_a \text{ of product acid}}}$$

$$K_{\text{eq}} = \frac{10^{-10.7}}{10^{-18}}$$

$$K_{\text{eq}} = 10^{7.3}$$

step 2 of 3

(b)

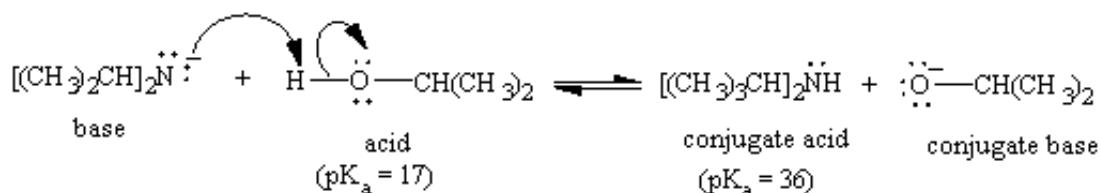


$$K_{eq} = \frac{10^{-4.7}}{10^{-10.2}}$$

$$K_{eq} = 10^{5.5}$$

step 3 of 3

(c)

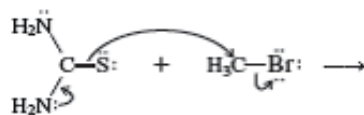


$$K_{eq} = \frac{10^{-17}}{10^{-36}}$$

$$\text{Or } K_{eq} = 10^{19}$$

Problem 64P

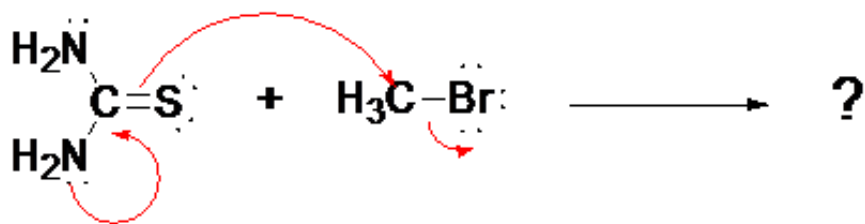
What are the products of the following reaction based on the electron flow represented by the curved arrows?



Step-by-step solution

step 1 of 2

Consider the following chemical reaction. Using the electron flow (as indicated), it is possible to predict the product(s) of the reaction.



Notice that a lone pair from a nitrogen atom forms a double bond with the carbon atom. Since this would lead to a carbon with five bonds (which is a violation of the octet rule), two electrons from the carbon-sulfur double bond simultaneously reach out and attack the methyl group of bromomethane. This methyl group is electrophilic due to the electronegative nature of bromine, which polarizes the methyl-bromine bond.

step 2 of 2

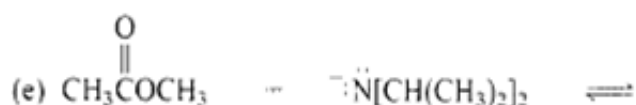
Since this would lead to a carbon with five bonds (which is a violation of the octet rule), the bromine atom leaves along with the two electrons from the carbon-bromine bond. This produces a negatively charged bromide anion and avoids the formation of a carbon with five substituents.



The end result is the attachment of a methyl group to the sulfur atom and the production of bromide anion. None of the atoms are violating the octet rule.

Problem 65P

Each of the following acid-base reactions involves substances found in Table 1.8. Use the pKa data in the table to help you predict the products of the reactions. Use curved arrows to show electron flow. Predict whether the equilibrium lies to the left or to the right and calculate the equilibrium constant for each reaction.

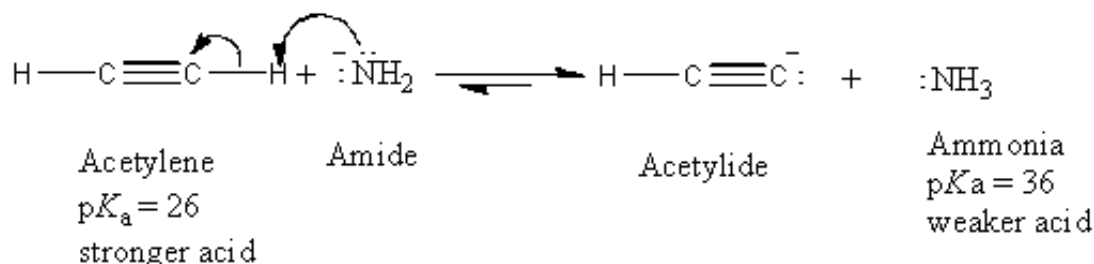


Step-by-step solution

step 1 of 15

(a) Amide is a Bronsted base and accepts a proton from the $-\text{C}\equiv\text{C}-\text{H}$ group of the acetylene to form ammonia and acetylide.

step 2 of 15



step 3 of 15

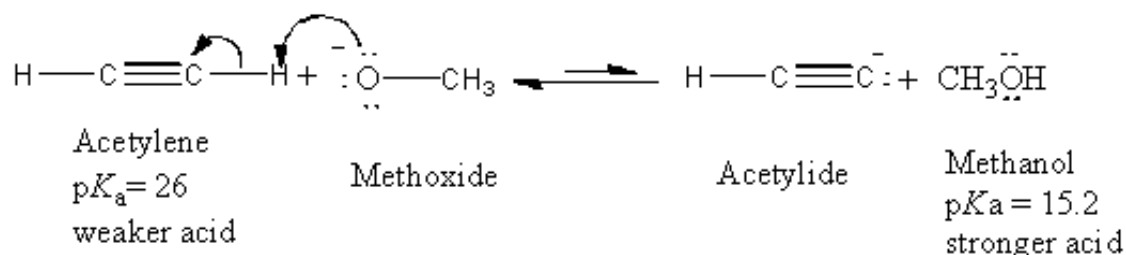
From the pK_a values, it is obvious that acetylene is much stronger acid than the ammonia and hence the equilibrium lies to the right. The equilibrium for the process is

$$K_{eq} = \frac{10^{-pK_a} \text{ of acetylene (reactant)}}{10^{-pK_a} \text{ of ammonia (product)}}$$
$$\text{or, } K_{eq} = \frac{10^{-26}}{10^{-36}}$$
$$\text{or, } \boxed{K_{eq} = 10^{10}}$$

step 4 of 15

(b) Methoxide is a Bronsted base and accepts a proton from the $-\text{C}\equiv\text{C}-\text{H}$ group of the acetylene to form methanol and acetylide.

step 5 of 15



step 6 of 15

From the pK_a values, we can say that methanol is a stronger acid than the acetylene and hence the equilibrium lies to the left. The equilibrium for the process is

$$K_{eq} = \frac{10^{-pK_a} \text{ of acetylene (reactant)}}{10^{-pK_a} \text{ of methanol (product)}}$$

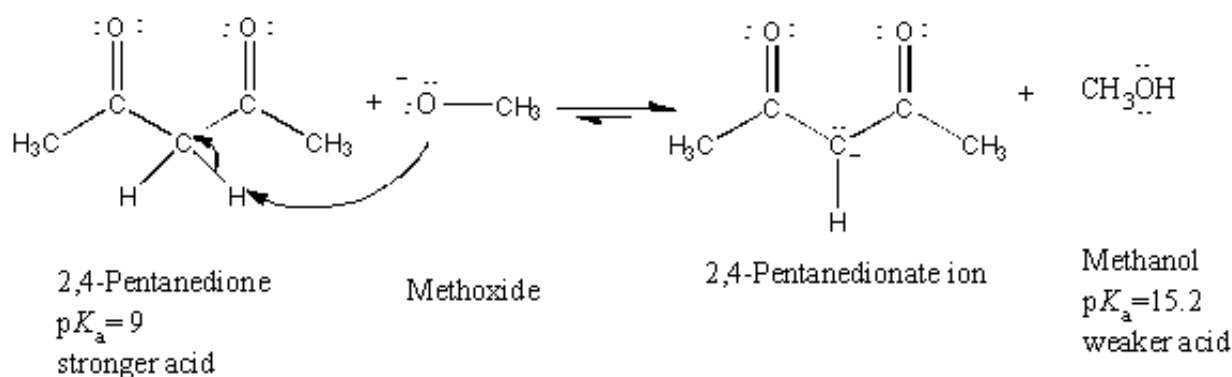
$$\text{or, } K_{eq} = \frac{10^{-26}}{10^{-15.2}}$$

$$\text{or, } \boxed{K_{eq} = 10^{-10.8}}$$

step 7 of 15

(c) Methoxide is a Bronsted base and accepts a proton from 2, 4-Pentanedione to form 2, 4-Pentanedionate ion and methanol.

step 8 of 15



step 9 of 15

From the pK_a values, it is obvious that 2, 4-Pentanedione is a stronger acid than the methanol and hence the equilibrium lies to the right. The equilibrium constant for the process is

$$K_{eq} = \frac{10^{-pK_a} \text{ of 2,4-pentanedione (reactant)}}{10^{-pK_a} \text{ of methanol (product)}}$$

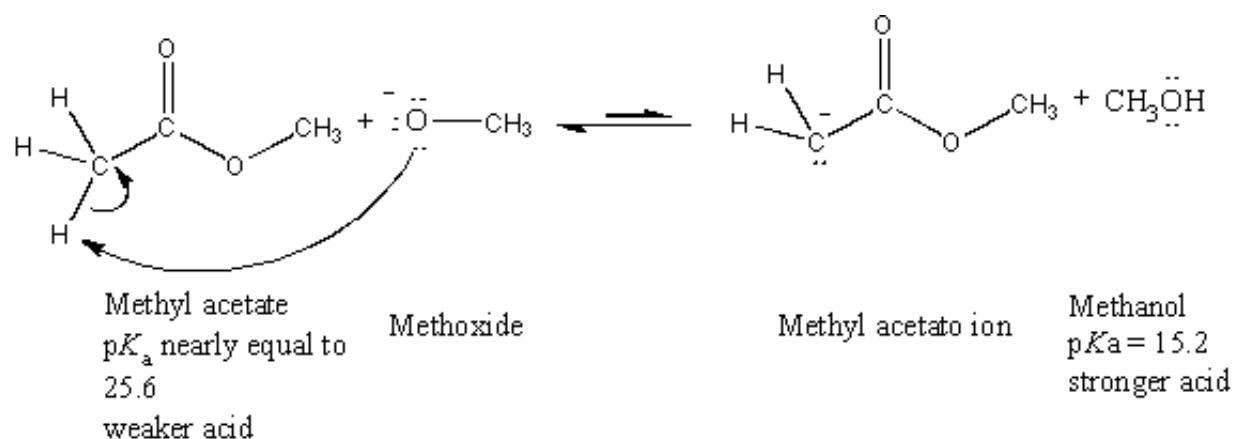
$$\text{or, } K_{eq} = \frac{10^{-9}}{10^{-15.2}}$$

$$\text{or, } \boxed{K_{eq} = 10^{6.2}}$$

step 10 of 15

(d) Methoxide is a Bronsted base and accepts a proton from the ketomethyl group of methyl acetate to form methanol and methylacetato ion.

step 11 of 15



step 12 of 15

From the p K_a values, we can say that methanol is a stronger acid than the methyl acetate and hence the equilibrium lies to the left. The equilibrium for the process is

$$K_{eq} = \frac{10^{-pK_a \text{ of methyl acetate (reactant)}}}{10^{-pK_a \text{ of methanol (product)}}$$

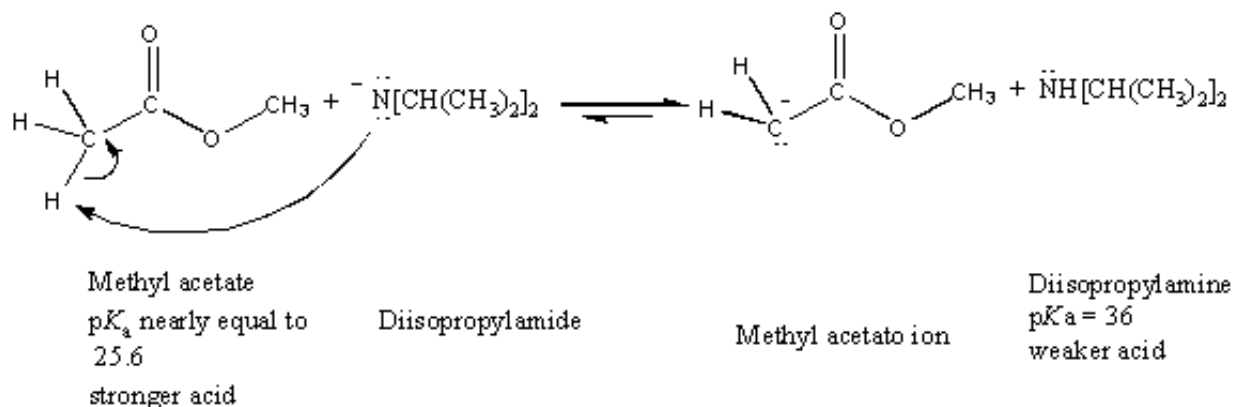
$$\text{or, } K_{eq} = \frac{10^{-25.6}}{10^{-15.2}}$$

$$\text{or, } \boxed{K_{eq} = 10^{-10.4}}$$

step 13 of 15

(e) Diisopropylamide is a Bronsted base and accepts a proton from the ketomethyl group of methyl acetate to form diisopropylamine and methyl acetate ion.

step 14 of 15



step 15 of 15

From the pK_a values, it is obvious that methyl acetate is a stronger acid than the diisopropylamine and hence the equilibrium lies to the right. The equilibrium constant for the process is

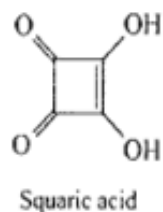
$$K_{eq} = \frac{10^{-pK_a \text{ of Methyl acetate (reactant)}}}{10^{-pK_a \text{ of Diisopropylamine (product)}}$$

$$\text{or, } K_{eq} = \frac{10^{-25.6}}{10^{-36}}$$

$$\text{or, } \boxed{K_{eq} = 10^{10.4}}$$

Problem 66P

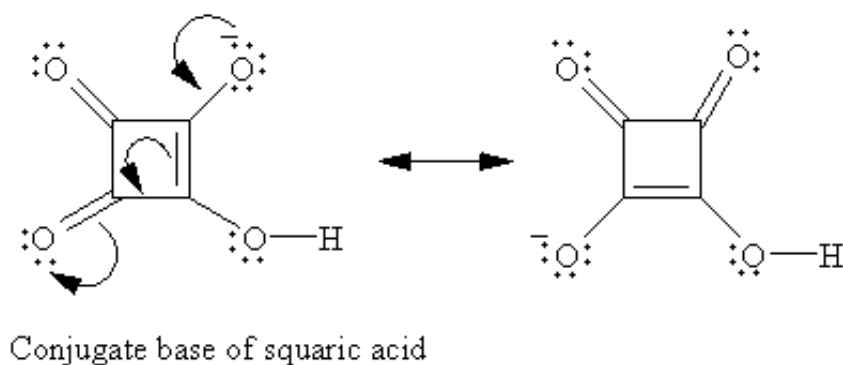
With a pK_a of 1.2, squaric acid is unusually acidic for a compound containing only C, H, and O.



Write a Lewis dot structure for the conjugate base of squaric acid and, using curved arrows, show how the negative charge is shared by two oxygens.

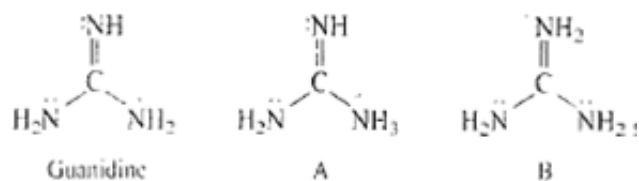
Step-by-step solution

step 1 of 1



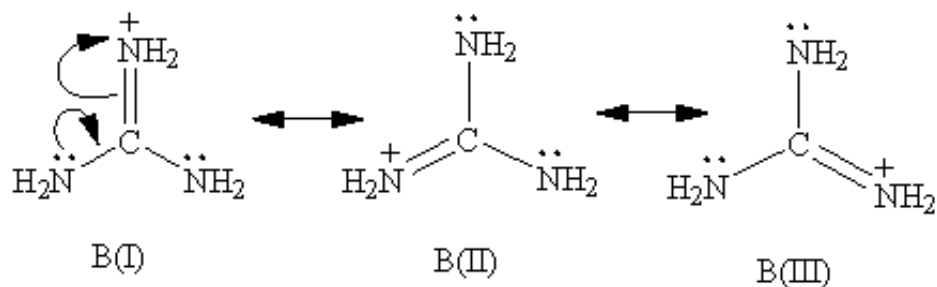
Problem 67P

Of two possible structures A and B for the conjugate acid of guanidine, the more stable is the one that is better stabilized by electron delocalization. Which one is it? Write resonance structures showing this electron delocalization.



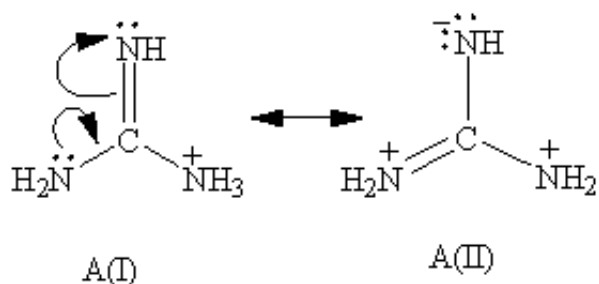
Step-by-step solution

step 1 of 2



All the canonical structures of B are equivalent and contribute equally to the resonance hybrid.

step 2 of 2



A(I) and A(II) are nonequivalent. Thus, their contribution is very minimal.
 So structure B is the more stable conjugate acid of guanidine.

Problem 68P

Formamide is a planar molecule. According to VSEPR, does the structural formula given for formamide satisfy this requirement?

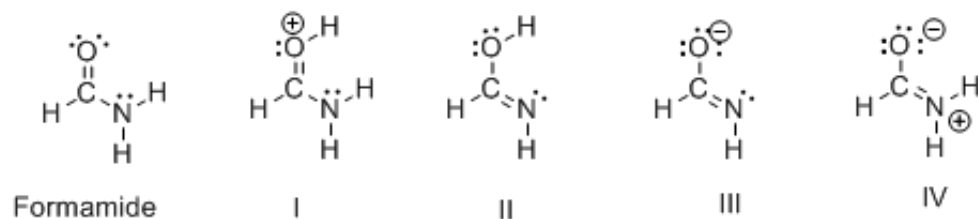
A. Yes

B. No

Step-by-step solution

step 1 of 2

Consider the molecule formamide, which is a planar molecule. It has a strong dipole moment and has multiple resonance structures. Consider the structures below, which show formamide along with a variety of Lewis structures that show species somewhat related to the Lewis structure of formamide.



According to VSEPR, does the formamide structural formula (above) explain the planarity of the formamide molecule?

- A. Yes
- B. No

step 2 of 2

VSEPR theory suggests that electrons in valence shells repel each other to their greatest extent. This avoids unnecessary overlap of the electron density. According to the structural formula shown above, the -NH_2 group should twist out of planarity to avoid overlap with the carbonyl.

Therefore, the correct option is (b).

VSEPR theory alone cannot account for the planarity of formamide. It is necessary to recognize the multiple-bond nature of the C-N bond (due to resonance), which hinders rotation of the C-N bond.

Problem 69P

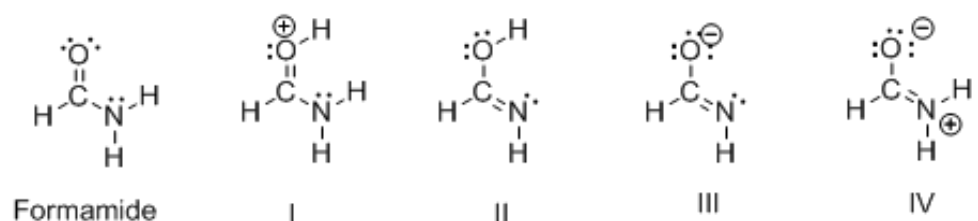
Which Lewis formula is both planar according to VSEPR and a resonance contributor of formamide?

- | | |
|-------|--------|
| A. I | C. III |
| B. II | D. IV |

Step-by-step solution

step 1 of 2

Consider the molecule formamide, which is a planar molecule. It has a strong dipole moment and has multiple resonance structures. Consider the structures below, which show formamide along with a variety of Lewis structures that show species somewhat related to the Lewis structure of formamide.



Which Lewis structure is both planar (according to VSEPR) and also a resonance contributor of

formamide?

- A. I
- B. II
- C. III
- D. IV

step 2 of 2

Compound I is not a resonance structure of formamide, as it has a different molecular formula; note the protonated carbonyl. The same is true for compound III. Compound II is not a resonance contributor of formamide; resonance structures can have their electrons moved around, but not the actual atoms.

Therefore, the correct option is **(d)**

Compound IV is both planar (according to VSEPR theory) as well as a resonance contributor of formamide. The lone pair of electrons on the formamide nitrogen can form a C-N double bond, leaving the oxygen with a negative charge. Due to the charge separation it may not be a particularly strong contributor to the overall hybrid structure but it's still a valid resonance structure.

Problem 70P

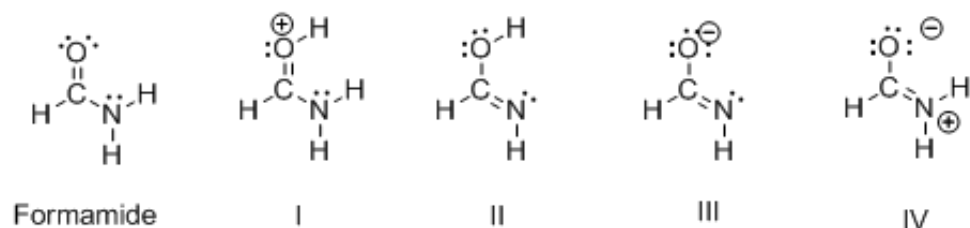
According to VSEPR, which Lewis formula has a pyramidal arrangement of bonds to nitrogen?

- | | |
|-------|--------|
| A. I | C. III |
| B. II | D. IV |

Step-by-step solution

step 1 of 2

Consider the molecule formamide, which is a planar molecule. It has a strong dipole moment and has multiple resonance structures. Consider the structures below, which show formamide along with a variety of Lewis structures that show species somewhat related to the Lewis structure of formamide.



According to VSEPR, which Lewis structure has a pyramidal-shaped arrangement of substituents around the nitrogen?

- A. I
- B. II
- C. III
- D. IV

step 2 of 2

A pyramidal-shaped arrangement of substituents around the nitrogen requires four separate substituents, so that the four substituents point to the four corners of the pyramid. Compounds II, III, and IV do not contain four different substituents.

Therefore, the correct option is (a).

Compound I has four substituents (the two hydrogens, the carbonyl carbon, and the lone pair of electrons) and so VSEPR theory would predict a pyramidal-shaped arrangement for these substituents.

Problem 71P

Which Lewis formula is a constitutional isomer of formamide?

A. I

B. II

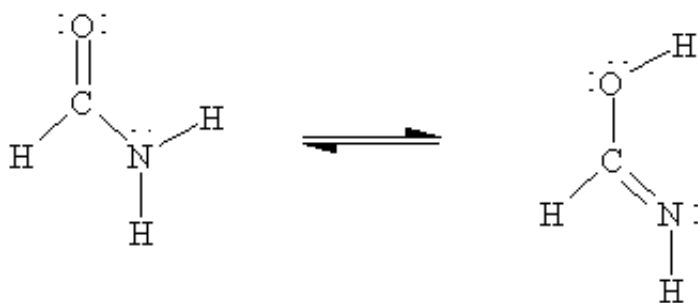
C. III

D. IV

Step-by-step solution

step 1 of 1

Constitutional isomers are isomers that differ in respect to the order in which the atoms are connected.



Formamide.
(Molecular formula CH_3NO)
(keto form)

Molecular formula CH_3NO
Constitutional isomers of formamide.
(II) (enol form.)

Thus, they are constitutional isomers of each other.

The correct answer is B (II).

Problem 72P

Which Lewis formula corresponds to the conjugate acid of formamide?

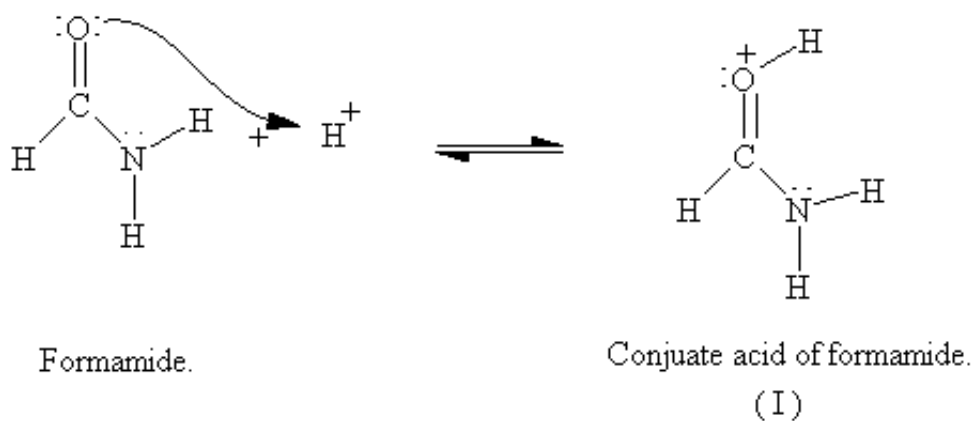
A. I

B. II

C. III

D. IV

Step-by-step solution



Thus, the answer is A I.

Problem 73P

Which Lewis formula corresponds to the conjugate base of formamide?

A. I

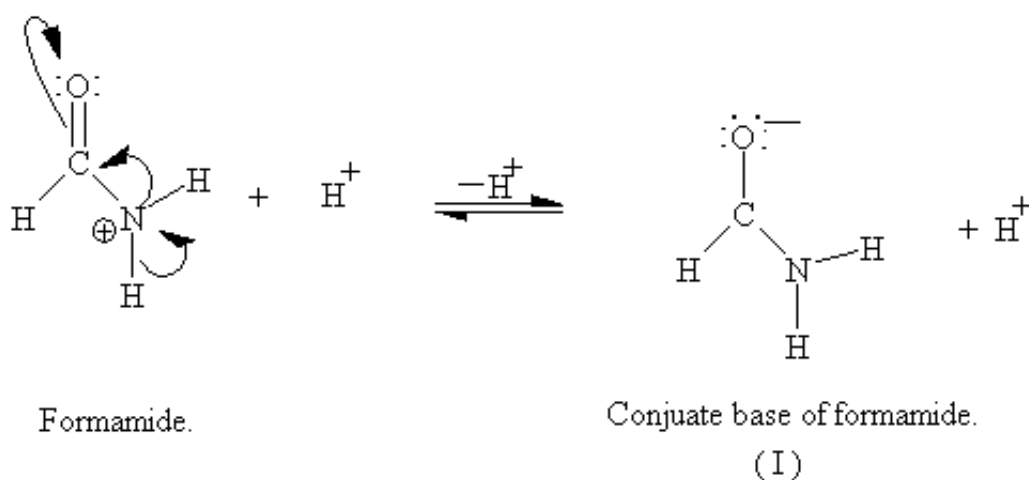
B. II

C. III

D. IV

Step-by-step solution

step 1 of 1



Thus, the answer is C III.

