Section 4.5: Molecular Polarity

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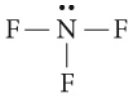
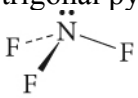
1. (a) Polarity of NF_3 molecule:

Step 1. Draw the simplified Lewis structure of nitrogen trifluoride.



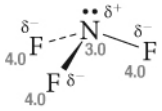
Step 2. Use VSEPR theory to predict the three-dimensional structure, then draw a diagram.

Since nitrogen is the central atom and it is surrounded by 4 electron pairs (3 bonding pairs and 1 lone pair), a molecule of nitrogen trifluoride will have a trigonal pyramidal shape.



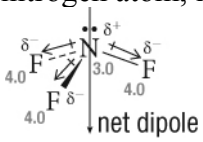
Step 3. Identify the electronegativity of each atom, and determine the partial charges in the molecule.

N = 3.0, δ+; F = 4.0, δ–



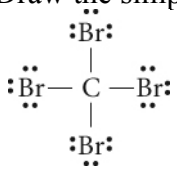
Step 4. Draw the bond dipoles, and determine whether the molecule has a net dipole.

The bond dipoles do not cancel each other out. Therefore, nitrogen trifluoride is a polar molecule with a net dipole pointing away from the lone pair on the nitrogen atom, in the direction of the fluorine atoms.



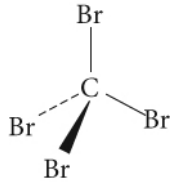
(b) Polarity of CBr_4 molecule:

Step 1. Draw the simplified Lewis structure of tetrabromomethane.



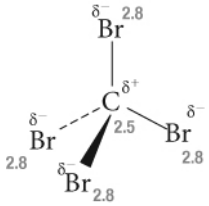
Step 2. Use VSEPR theory to predict the three-dimensional structure, then draw a diagram.

Since carbon is the central atom and it is bonded to 4 bromine atoms, a molecule of tetrabromomethane will have a tetrahedral shape.



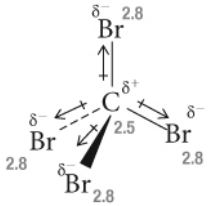
Step 3. Identify the electronegativity of each atom, and determine the partial charges in the molecule.

C = 2.5, δ+; Br = 2.8, δ–



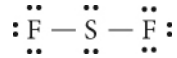
Step 4. Draw the bond dipoles, and determine whether the molecule has a net dipole.

In this tetrahedral shape, all of the atoms bonded to the central carbon atom are identical and all the directions are equivalent. Therefore, the 4 bond dipoles cancel each other. The overall net dipole is zero, so tetrabromomethane is a non-polar molecule.



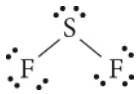
(c) Polarity of SF_2 molecule:

Step 1. Draw the simplified Lewis structure of sulfur difluoride.



Step 2. Use VSEPR theory to predict the three-dimensional structure, then draw a diagram.

The central sulfur atom is surrounded by 4 pairs of electrons (2 bonding pairs and 2 lone pairs), so SF_2 has a bent structure.

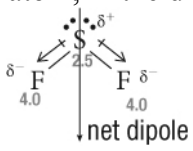


Step 3. Identify the electronegativity of each atom, and determine the partial charges in the molecule.

S = 2.5, δ+; F = 4.0, δ–

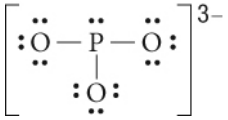
Step 4. Draw the bond dipoles, and determine whether the molecule has a net dipole.

In this bent shape, the bond dipoles do not cancel each other out. Therefore, sulfur difluoride is a polar molecule with a net dipole pointing away from the lone pairs on the nitrogen atom, in the direction of the fluorine atoms.



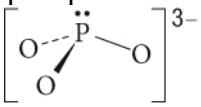
2. (a) Polarity of PO_3^{3-} ion:

Step 1. Draw the simplified Lewis structure of the phosphite ion.



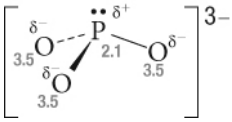
Step 2. Use VSEPR theory to predict the three-dimensional structure, then draw a diagram.

Because the central phosphorus atom has 3 bonding pairs of electrons and 1 lone pair, the PO_3^{3-} ion has a trigonal pyramidal arrangement of atoms around the phosphorus atom.



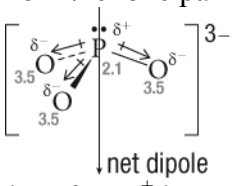
Step 3. Identify the electronegativity of each atom, and determine the partial charges in the molecule.

P = 2.1, δ+; O = 3.5, δ–



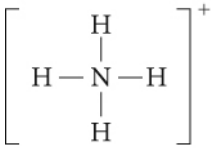
Step 4. Draw the bond dipoles, and determine whether the molecule has a net dipole.

In this trigonal pyramidal shape, the bond dipoles do not cancel each other out. Therefore, sulfur difluoride is a polar molecule with a net dipole pointing away from the lone pair on the phosphorus atom, toward the oxygen atoms.



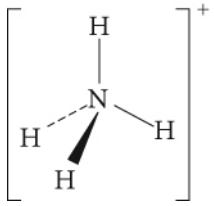
(b) Polarity of NH_4^+ ion:

Step 1. Draw the simplified Lewis structure of the ammonium ion.



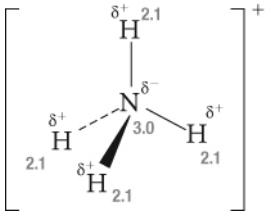
Step 2. Use VSEPR theory to predict the three-dimensional structure, then draw a diagram.

Since nitrogen is the central atom and it is bonded to 4 hydrogen atoms, the ammonium ion will have a tetrahedral shape.



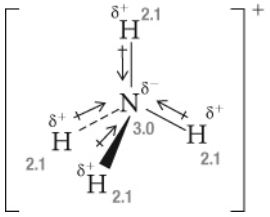
Step 3. Identify the electronegativity of each atom, and determine the partial charges in the molecule.

N = 3.0, δ–; H = 2.1, δ+



Step 4. Draw the bond dipoles, and determine whether the molecule has a net dipole.

In this tetrahedral shape, all of the atoms bonded to the central nitrogen atom are identical and all the directions are equivalent. Therefore, the 4 bond dipoles cancel each other. The overall net dipole is zero, so the ammonium ion is non-polar.

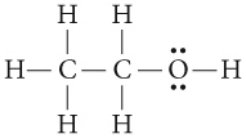


(c) Polarity of B_2H_6 molecule:

The electronegativities of boron and hydrogen are 2.0 and 2.1, respectively. Therefore, the bonds within this molecule must be non-polar covalent. Consequently, the molecule is non-polar regardless of its shape.

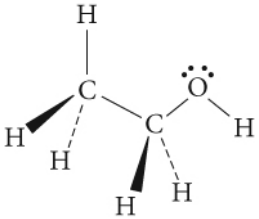
(d) Polarity of CH_3CH_2OH molecule:

Step 1. Draw the simplified Lewis structure of ethanol.



Step 2. Use VSEPR theory to predict the three-dimensional structure, then draw a diagram.

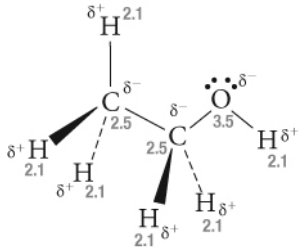
The ethanol molecule has 4 pairs of bonding electrons around each central carbon atom. There are 2 pairs of bonding electrons around the oxygen atom, which also has 2 lone pairs. The geometry around each carbon atom is tetrahedral. The geometry around the oxygen atom is a bent structure.



Step 3. Identify the electronegativity of each atom, and determine the partial charges in the molecule.

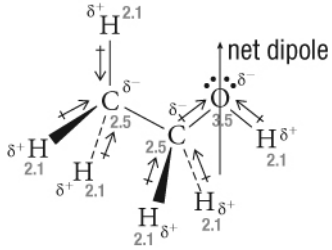
The carbon atoms are more electronegative than the surrounding hydrogen atoms, so they accumulate a partial negative charge from each bond with hydrogen: C = 2.5, δ–; H = 2.1, δ+.

Oxygen is more electronegative than either carbon or hydrogen: O = 3.5, δ–; C, H, δ+.



Step 4. Draw the bond dipoles, and determine whether the molecule has a net dipole.

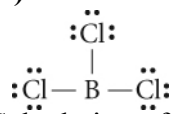
The bonds around the two carbon atoms are fairly balanced, but there is a strong collection of negative charges around the oxygen atom that will create a net dipole, so ethanol is a polar molecule.



3. Answers will vary. Sample answer:

(a) An example of a substance that has polar bonds, but the molecule is non-polar, is BCl_3. [Answer must be a symmetrical molecule.]

(b) Lewis structure for the BCl_3 molecule:

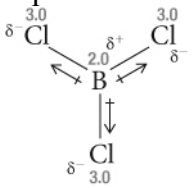


Calculation of the polarity of the B–Cl bonds:

$$\Delta EN_{Cl-B} = EN_{Cl} - EN_{B}$$
$$= 3.0 - 2.0$$
$$\Delta EN_{Cl-B} = 1.0$$

Since $0.5 < \Delta EN < 1.7$, the B–Cl bonds are polar covalent.

(c) Diagram of the structure of the BCl_3 molecule showing why the molecule is non-polar.



The three bond dipoles cancel each other out, resulting in a net dipole of 0, so the BCl_3 molecule is non-polar molecule.

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1. A molecule of diboron tetrafluoride, B_2F_4, is non-polar because the molecule has a trigonal planar shape around both B atoms, resulting in a symmetrical molecule. Since the B–F bonds will have the same difference in electronegativity but their dipoles will point in opposite directions, they will cancel each other out and the molecule will be non-polar.

2. A molecule of methanol, CH_3OH, is polar because it is tetrahedral in shape about the carbon atom and angular or bent around the oxygen atom. The C–O bond has a greater difference in electronegativity than the C–H bonds, and since the O has a negative partial charge, this causes the molecule to be polar with a net dipole pointing in the direction of the O.

3. PH_3 is non-polar, unlike PF_3, because P and H have identical electronegativities.

4. Using partial charges to represent the dipole in a bond:

(a) δ– F–B δ+

(b) There is no partial charge in the N–Cl bond, as $EN_N = EN_{Cl} = 3.0$.

(c) δ– C–H δ+

(d) There is no partial charge in the C–C bond because the atoms are identical.

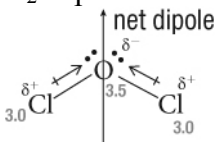
5. (a) Polarity of OCl_2:

$$\Delta EN_{O-C} = EN_O - EN_C$$
$$= 3.5 - 2.5$$
$$\Delta EN_{O-C} = 1.0$$

The molecule has more than 1 polar covalent bond.

The molecular structure is bent, so the molecule is not symmetrical.

OCl_2 is polar.



(b) Polarity of BeH_2:

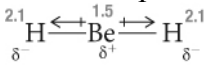
$$\Delta EN_{H-Be} = EN_H - EN_{Be}$$
$$= 2.1 - 1.5$$
$$\Delta EN_{H-Be} = 0.6$$

The molecule has more than 1 polar covalent bond.

The molecular structure is linear, so the molecule is symmetrical.

The atoms bonded to the central atom are the same.

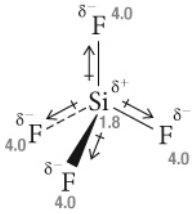
BeH_2 is non-polar.



(c) Polarity of SiF_4:

$$\Delta EN_{F-Si} = EN_F - EN_{Si}$$
$$= 4.0 - 1.8$$
$$\Delta EN_{F-Si} = 2.2$$

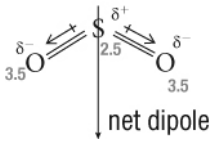
The difference in electronegativity suggests the bonding in this molecule should be ionic bonds. However, experimental evidence suggests that the bonds are still polar covalent in nature.
The molecular structure is tetrahedral, so the molecule is symmetrical.
The atoms bonded to the central atom are the same.
SiF_4 is non-polar.



(d) Polarity of SO_2:

$$\Delta EN_{O-S} = EN_O - EN_S$$
$$= 3.5 - 2.5$$
$$\Delta EN_{O-S} = 1.0$$

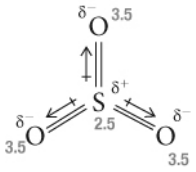
The molecule has more than 1 polar covalent bond.
The molecular structure is bent, so the molecule is not symmetrical.
SO_2 is polar.



(e) Polarity of SO_3:

$$\Delta EN_{O-S} = EN_O - EN_S$$
$$= 3.5 - 2.5$$
$$\Delta EN_{O-S} = 1.0$$

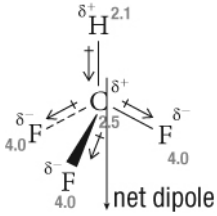
The molecule has more than 1 polar covalent bond.
The molecular structure is trigonal planar, so the molecule is symmetrical.
The atoms bonded to the central atom are the same.
SO_3 is non-polar.



(f) Polarity of CHF_3:

$$\Delta EN_{F-C} = EN_F - EN_C$$
$$= 4.0 - 2.5$$
$$\Delta EN_{F-C} = 1.5$$

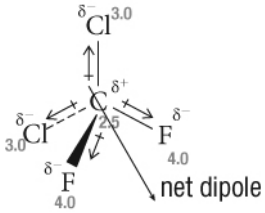
The molecule has more than 1 polar covalent bond.
The molecular structure is tetrahedral, but since the atoms attached to the central carbon atom are not identical, the molecule is asymmetrical.
Therefore, CHF_3 is polar.



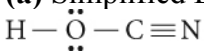
(g) Polarity of CCl_2F_2:

$$\Delta EN_{F-C} = EN_F - EN_C$$
$$= 4.0 - 2.5$$
$$\Delta EN_{F-C} = 1.5$$

The molecule has more than 1 polar covalent bond.
The molecular structure is tetrahedral, but since the atoms attached to the central carbon atom are not identical, the molecule is asymmetrical.
Therefore, CCl_2F_2 is polar.

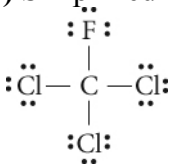


6. (a) Simplified Lewis structure for HOCN:



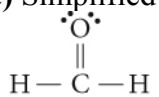
The O atom has 4 electron pairs (2 bonding and 2 lone), so the molecule is bent at the O atom. This is an asymmetrical structure, so the HOCN molecule is polar.

(b) Simplified Lewis structure for CF_3Cl:



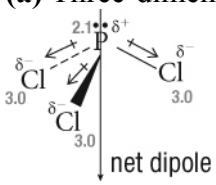
The central atom has 4 bonding electron pairs and no lone pairs, so the molecule is tetrahedral. The atoms bonded to the C are different, so the CF_3Cl molecule is polar.

(c) Simplified Lewis structure for H_2CO:

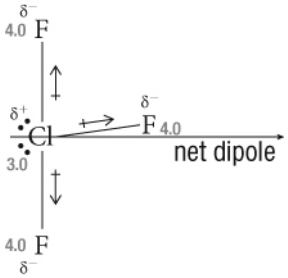


The central atom has 3 groups of electrons and no lone pairs, so the molecule is trigonal planar. The atoms bonded to the C are different, so the H_2CO molecule is polar.

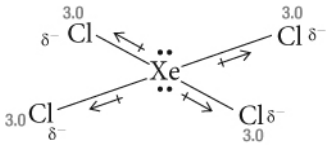
7. (a) Three-dimensional diagram indicating that PCl_3 is polar:



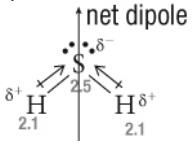
(b) Three-dimensional diagram indicating that ClF_3 is polar:



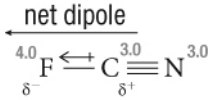
(c) Three-dimensional diagram indicating that $XeCl_4$ is non-polar:
Note that the electronegativity of the noble gases is not given on the periodic table because these elements rarely form compounds. However, since Xenon is near the bottom of the noble gas family, we can assume that it has a lower electronegativity than Cl.



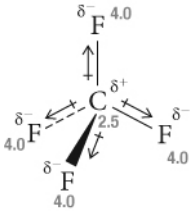
(d) Three-dimensional diagram indicating that H_2S is polar:



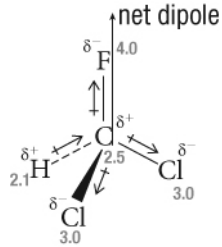
(e) Three-dimensional diagram indicating polarity for FCN:



8. (a) Three-dimensional diagram indicating that tetrafluoromethane, CF_4, is non-polar:



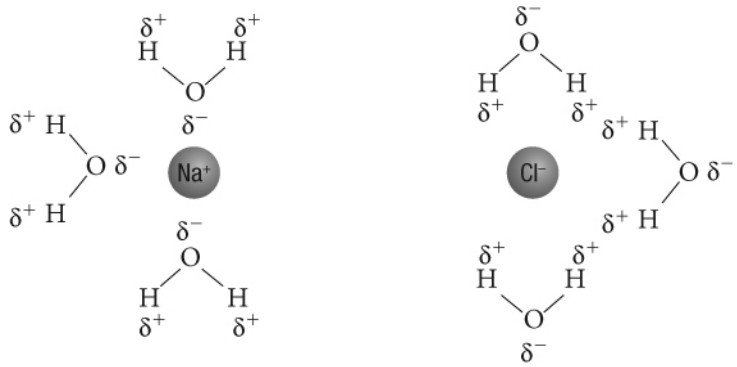
Three-dimensional diagram indicating that dichlorofluoromethane, $CHFCl_2$, is polar:



(b) Both dichlorofluoromethane and tetrafluoromethane are tetrahedral, so the difference between the molecules is in the distribution of charges in the bond dipoles. $CHFCl_2$ has multiple bond dipoles, which are not equivalent due to different electronegativities of the atoms bonded to the central atom. This results in a net dipole and thus the molecule is polar. CF_4 has an even distribution of charges in the bond dipoles, which cancel each other out, and is therefore non-polar.

(c) Answers may vary. Sample answer: Since dichlorofluoromethane is polar, there would be stronger intermolecular forces between its molecules than between those of CF_4, due to dipole–dipole interactions. Some physical properties that depend on intermolecular forces include melting point, boiling point, viscosity, solubility, binding affinity, miscibility, surface tension, adhesion, hydrophobicity, heat of vaporization, heat of fusion, elasticity, tensile strength, and capacitance. The boiling point and heat of vaporization are important considerations in the use of each compound as a refrigerant.

9.



10. (a) Carbon dioxide can absorb infrared light because the bonds between carbon and oxygen can undergo asymmetric stretching. When carbon dioxide absorbs infrared light, the oxygen atoms are pulled away from the carbon atom unequally. This changes the angle between the carbon–oxygen bonds, which creates temporary dipole moments.

(b) Answers may vary. Sample answer: When scientists discuss the greenhouse effect, carbon dioxide is held responsible because the concentration of carbon dioxide in the atmosphere is increasing due to human activities. The concentration of water vapour in the atmosphere has been constant for centuries and is not being altered significantly by humans. However, there are reasons to focus on other greenhouse gases, such as methane, that are also being released into the atmosphere in increasing amounts as a result of anthropogenic activities.