Task 2

Due: 11:59pm on Friday, April 27, 2018

To understand how points are awarded, read the Grading Policy for this assignment.

Additional Problem 8.14

What is the Lewis symbol for each of the following atoms or ions?

Part A

Ca

Draw the Lewis dot structure for Ca. Show the formal charges of all atoms in the correct structure. To change the symbol of an atom, double-click on the atom and enter the letter of the new atom.

ANSWER:

Correct

Part B

P

Draw the Lewis dot structure for P. Show the formal charges of all atoms in the correct structure. To change the symbol of an atom, double-click on the atom and enter the letter of the new atom.

ANSWER:

Correct

Part C

$\mathrm{Mg^{2+}}$

Draw the Lewis dot structure for $\mathrm{Mg^{2+}}$. Show the formal charges of all atoms in the correct structure. To change the symbol of an atom, double-click on the atom and enter the letter of the new atom.

ANSWER:

Correct

Part D

S^{2-}

Draw the Lewis dot structure for S^{2-}. Show the formal charges of all atoms in the correct structure. To change the symbol of an atom, double-click on the atom and enter the letter of the new atom.

ANSWER:

Correct

Bond Polarity

In a covalent bond between two atoms, the more electronegative atom will attract more electron density toward itself, causing a *polar bond*. The more electronegative element is the negative pole, whereas the less electronegative element is the positive pole. The direction of the dipole is always toward the more electronegative element. This is often indicated by an arrow, as shown in the figure.

The magnitude of the bond polarity is the difference in electronegativity values of the atoms. For example, in the molecule Cl_2, the Cl–Cl bond is nonpolar because there is no difference in electronegativity between two atoms of the same element. In the $ClBr$ molecule, however, the Cl–Br bond is polar because Cl and Br have different electronegativity values. A Cl–I bond would be more polar than a Cl–Br bond because there is a greater electronegativity difference between Cl and I than between Cl and Br.

Part A

In the molecule HCl, which atom is the negative pole?

Hint 1. How to approach the problem

The negative pole of a polar bond is always the more electronegative element. Based on their positions in the periodic table, you can determine the relative polarities of Cl and H without knowing the exact values.

ANSWER:

- Cl
- H

Correct

Part B

Of the molecules SiF_4 and $SiBr_4$, which has bonds that are more polar?

Hint 1. How to approach the problem

The magnitude of bond polarity depends on the difference in electronegativity between the atoms involved in the bond. Based on their positions in the periodic table, you can determine relative electronegativities of Si, F, and Br without actually knowing the exact values.

Hint 2. Determine the electronegativity differences

Based on their locations in the periodic table, which two atoms have the greatest difference in electronegativity?

Hint 1. Electronegativity trend

Electronegativity increases as one moves from left to right and bottom to top in the periodic table. Therefore, elements that are close together in the periodic table tend to have similar electronegativities. Elements that are far apart in the periodic table tend to have very different electronegativities.

ANSWER:

- Si and F
- Si and Br

ANSWER:

- SiF_4
- $SiBr_4$

Correct

Part C

Both of the sulfur-chlorine single bonds in SCl_2 are polar. In which direction should the polarity arrows point?

ANSWER:

- to the right
- to the left
- toward the central sulfur atom
- away from the central sulfur atom

Correct

Formal Charge of a Diatomic Molecule

The Lewis structure of HBr is shown here:

The formula for formal charge

The formal charge of an atom in a molecule can be calculated based on the following formula:

$$\text{formal charge} = \text{valence e}^- - \left(\text{nonbonding e}^- + \frac{\text{bonding e}^-}{2}\right)$$

Part A

What is the formal charge on the hydrogen atom in HBr? Use the periodic table as needed.

Express the formal charge numerically (e.g., +1).

Hint 1. Determine the number of valence electrons for hydrogen

How many valence electrons are there in a hydrogen atom? Use the periodic table as needed.

Express your answer numerically as an integer.

ANSWER:

1

Correct

Hint 2. Determine the number of nonbonding electrons on hydrogen

How many nonbonding electrons are shown around the hydrogen atom in the Lewis structure for HBr?

Express your answer numerically as an integer.

ANSWER:

0

Correct

Hint 3. Determine the number of bonding electrons on hydrogen

How many bonding electrons are shown around the hydrogen atom in the Lewis structure for HBr?

Express your answer numerically as an integer.

ANSWER:

2

Correct

ANSWER:

0

Correct

Part B

What is the formal charge on the bromine atom in HBr? Use the periodic table as needed.

Express the formal charge numerically (e.g., +1).

Hint 1. Determine the number of valence electrons for bromine

How many valence electrons are there in an bromine atom? Use the periodic table as needed.

Express your answer numerically as an integer.

ANSWER:

7

Correct

Hint 2. Determine the number of nonbonding electrons on bromine

How many nonbonding electrons are shown around the bromine atom in the Lewis structure for HBr?

Express your answer numerically as an integer.

ANSWER:

6

Correct

Hint 3. Determine the number of bonding electrons on bromine

How many bonding electrons are shown around the bromine atom in the Lewis structure for HBr?

Express your answer numerically as an integer.

ANSWER:

2

Correct

ANSWER:

0

Correct

Because the formal charges of the individual atoms add to zero, the hydrogen bromide molecule is neutral overall. Formal charges are useful in assessing the feasibility of a bonding arrangement when more than one valid Lewis structure can be written for a molecule.

Drawing and Predicting Lewis Structures

Learning Goal:

To learn to draw Lewis structures, assign formal charges, and predict likely structures based on formal charges.

Use the following steps to draw Lewis structures from molecular formulas:

1. Sum the valence electrons from all of the atoms.
2. Connect the appropriate atoms using a single bond (a line, which represents two electrons). The central atom is usually the least electronegative atom, but never hydrogen.
3. Add lone pairs to complete the octets of the outer atoms. Keep in mind that hydrogen fills its "octet" with only two electrons.
4. Place any leftover electrons on the central atom. Keep in mind that elements in row 3 of the periodic table and beyond can exceed the general octet.
5. If any atom lacks an octet, try using double or triple bonds. You may have to use one or more lone pairs to make the double or triple bonds so that the total number of valence electrons remains the same.

Part A

Draw the Lewis structure of $\mathrm{PH_3}$.

To add lone pairs, click the  button *before* clicking on the molecule.

Draw the molecule by placing atoms on the grid and connecting them with bonds. Include lone pairs of electrons and hydrogen atoms.

Hint 1. Determine the total number of valence electrons

What is the total number of valence electrons in $\mathrm{PH_3}$?

Express your answer as an integer.

Hint 1. Determine the number of valence electrons in a phosphorus atom

Phosphorus, P, is in group 5A of the periodic table. How many valence electrons are in an atom of phosphorus?

Express your answer as an integer.

ANSWER:

5 valence electron(s)

Hint 2. Determine the number of valence electrons in a hydrogen atom

Hydrogen, H, is in group 1A of the periodic table. How many valence electrons are in an atom of hydrogen?

Express your answer as an integer.

ANSWER:

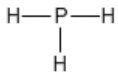
1 valence electron(s)

ANSWER:

8 valence electron(s)

Hint 2. Determine the number of electrons from bonds

Phosphorus will be the central atom since hydrogen can never be a central atom. So we begin by connecting three H atoms to one P atom, each with a single bond:



How many electrons are currently shown in this structure if each single bond counts as two electrons.

Express your answer as an integer.

ANSWER:

6 electrons

ANSWER:

Correct

Lewis structures of ions

A charge on a formula affects the total number of valence electrons that should appear in the structure. For anions, add the magnitude of the charge to the total. For cations, subtract the magnitude of the charge from the total.

Part B

Draw the Lewis structure of $\mathrm{NO_2}^-$.

To add lone pairs, click the button *before* clicking on the molecule.

Draw the molecule by placing atoms on the grid and connecting them with bonds. Include all lone pairs of electrons.

Hint 1. Determine the number of total valence electrons

What is the total number of valence electrons in $\mathrm{NO_2}^-$?

The total number of valence electrons in the NO_2^- ion is the sum of the valence electrons of the individual atoms and the charge on the ion.

Express your answer as an integer.

Hint 1. Determine the number of valence electrons in a nitrogen atom

Nitrogen, N, is in group 5A of the periodic table. How many valence electrons are in an atom of nitrogen?

Express your answer as an integer.

ANSWER:

5 valence electron(s)

Hint 2. Determine the number of valence electrons in an oxygen atom

Oxygen, O, is in group 6A of the periodic table. How many valence electrons are in an atom of oxygen?

Express your answer as an integer.

ANSWER:

6 valence electron(s)

Hint 3. Identify the effect of the charge on the total valence

What does a -1 charge indicate about the number of valence electrons in the ion?

ANSWER:

- There is one fewer electron than what the atoms contribute.
- (•) There is one extra electron beyond what the atoms contribute.
- The charge has no relation to the total number of valence electrons.

ANSWER:

18 electrons

Hint 2. Determine the number of lone pairs needed on each oxygen atom

The central atom in this case is nitrogen, just as we would have predicted based on electronegativities. So we begin by connecting two O atoms to one N atom, each with a single bond:

O—N—O

How many lone pairs should be added to each O atom to complete their octets?

ANSWER:

Add
- one lone pair (two electrons)
- two lone pairs (four electrons)
- (•) three lone pairs (six electrons)
- four lone pairs (eight electrons)

to each O atom.

Hint 3. Identify how to achieve the correct total valence electrons

At this point, our structure looks like this:

:Ö—N—Ö:

How many more electrons need to be added to the nitrogen atom to achieve a total of 18 electrons?

ANSWER:

- (•) two electrons (one lone pair)
- four electrons (two lone pairs)
- zero electrons (i.e., the structure currently shows the correct number)

ANSWER:

Correct

Assigning formal charges

To determine the formal charges of the atoms in a molecule, we must assign each electron to a particular atom. Then, we compare the number of electrons assigned to each atom to the expected valence of that element.

All lone pairs of electrons should be assigned to the atom on which they are found; bonding pairs of electrons should be divided equally between the atoms in the bond. The formal charge is calculated by subtracting the number of electrons assigned to the atom from the number of valence electrons in the free atom.

$$\text{Formal charge} = (\text{Number of valence } e^- \text{ in free atom}) - \frac{1}{2}(\text{Number of bonding } e^-) - (\text{Number of nonbonding } e^-)$$

For example, in $:\mathrm{C}{\equiv}\mathrm{O}:$, each atom is assigned five electrons (two nonbonding from the atom's lone pair and three bonding from half of the triple bond). Since the expected valence of carbon is four, its formal charge is $(4 \text{ valence}) - (5 \text{ assigned}) = -1$. Since the expected valence of oxygen is six, its formal charge is $(6 \text{ valence}) - (5 \text{ assigned}) = +1$.

For a neutral molecule such as CO, the sum of the formal charges is zero. For a polyatomic ion, the sum of the formal charges is equal to the overall charge on the ion.

Part C

Assign formal charges to each atom in the $\mathrm{O_3}$ molecule shown here. Be sure to click the + or - button *before* clicking on the molecule.

Click the + or - button until it turns yellow, then click on an atom, and then type the desired charge.

Hint 1. Count the electrons assigned to the left-most atom

How many electrons are assigned to this oxygen atom?

Express your answer as an integer.

ANSWER:

7 assigned electrons

Hint 2. Count the electrons assigned to the middle atom

How many electrons are assigned to this oxygen atom?

Express your answer as an integer.

ANSWER:

5 assigned electrons

Hint 3. Count the electrons assigned to the right-most atom

How many electrons are assigned to this oxygen atom?

Express your answer as an integer.

ANSWER:

6 assigned electrons

Hint 4. Determine the expected number of valence electrons for oxygen

Oxygen, O, is in group 6A of the periodic table. How many valence electrons are expected to be in an atom of oxygen?

Express your answer as an integer.

ANSWER:

6 valence electrons

ANSWER:

Correct

Using formal charges to predict Lewis structures

Sometimes, there will be more than one valid Lewis structure for a molecule. For example, consider these options for the $\mathrm{SO_2}$ molecule:

Both structures are valid because they each contain 18 electrons, all oxygen atoms have an octet, and both sulfur atoms have at least an octet. Formal charges however, are minimized in the second structure. Thus, based on formal charges alone, we would predict that the second structure is the preferred Lewis structure for $\mathrm{SO_2}$.

Part D

Based on formal charges, draw the *most preferred* Lewis structure for the chlorate ion, $\mathrm{ClO_3^-}$?

To add lone pairs, click the button *before* clicking on the molecule. To add formal charges, click the + or - button *before* clicking on the molecule.

Draw the molecule by placing atoms on the grid and connecting them with bonds. Include lone pairs of electrons and formal charges.

Hint 1. How to approach the problem

To draw the most preferred Lewis structure for ClO_3^- , the first step would be to identify all the possible Lewis structures for ClO_3^-. The central atom in this case is chlorine, just as we would have predicted based on electronegativities.

Then add the lone pairs of electrons to each atom in the structures. Each of these ClO_3^- structures would have the same number of total valence electrons. The second step would be to assign the formal charges to each atom of the structures. Now the most preferred structure would be the structure with the most minimized formal charges. If there is more than one structure with the most minimized formal charges, then choose the structure in which the central atom has a formal charge closest to zero.

Hint 2. Complete some possible structures

Some possible Lewis structures for ClO_3^- are shown here. Add lone pairs to each of the possible structures. Be sure to click the button *before* clicking on the molecule to add the lone pairs.

Add nonbonding electrons as necessary.

Hint 1. Determine the total number of valence electrons

What is the total number of valence electrons in ClO_3^-?

The total number of valence electrons in the ClO_3^- ion is the sum of the valence electrons of the individual atoms and the charge on the ion.

Express your answer as an integer.

ANSWER:

26 valence electrons

ANSWER:

Hint 3. Assign formal charges to the possible structures

Assign a formal charge to each atom of the structures shown here. Be sure to clickthe + or - button *before* clicking on the molecule.

Click the + or - button, then click on an atom, and then type the desired charge.

Hint 1. How to assign formal charges

To assign the formal charges of the atoms in a molecule, determine the number of electrons assigned to each atom in the given structures. Then compare the number of assigned electrons to the expected valence electrons of that element.
If an atom has been assigned more electrons than its expected valence, it will have a negative formal charge. If an atom has been assigned fewer electrons than its expected valence, it will have a positive formal charge.

All lone pairs of electrons should be assigned to the atom on which they are found; bonding pairs of electrons should be divided equally between the atoms in the bond. The formal charge is calculated by subtracting the number of electrons assigned to the atom from the number of valence electrons in the isolated atom.

Hint 2. Determine the expected number of valence electrons for chlorine

Chlorine, Cl, is in group 7A of the periodic table. How many valence electrons are expected to be in an atom of chlorine?

Express your answer as an integer.

ANSWER:

7 valence electrons

ANSWER:

ANSWER:

Correct

Exceptions to the Octet Rule

Hydrogen, beryllium, and boron are examples of elements that can have less than an octet of electrons in a covalent compound. Even without an octet, hydrogen achieves the helium configuration, and therefore noble-gas stability, by forming one covalent bond. Beryllium and boron simply do not have enough electrons to form the number of bonds needed to achieve an octet. Instead, Be uses its two valence electrons to form two bonds, and B uses its three valence electrons to form three bonds. None of these elements will have any lone pairs.

Elements in the third row of the periodic table and beyond can have more than an octet of electrons in a covalent compound. This can be explained by the existence of an empty *d* subshell available to these elements, which allows them to expand their valence to a number greater than eight.

Part A

How many lone pairs are on the central atom in $\mathrm{BCl_3}$?

Express your answer numerically as an integer.

Hint 1. What is the valence of boron?

How many valence electrons are in an atom of boron?

Express your answer numerically as an integer.

Hint 1. The electron configuration of boron

The ground-state electron configuration of neutral boron is $1\mathrm{s}^2\,2\mathrm{s}^2\,2\mathrm{p}^1$

ANSWER:

3

ANSWER:

0

Correct

In neutral covalent compounds, boron typically has three bonds and no lone pairs. Therefore it has only six electrons around it (not eight), which is why it is an exception to the octet rule.

Part B

How many lone pairs are on the central atom of $\mathrm{BrF_3}$?

Express your answer numerically as an integer.

Hint 1. Determine the valence of bromine

How many valence electrons are in an atom of bromine, Br?

Express your answer numerically as an integer.

Hint 1. The electron configuration of bromine

The ground-state electron configuration of neutral bromine is $[\mathrm{Ar}]3\mathrm{d}^{10}4\mathrm{s}^2 4\mathrm{p}^5$.

ANSWER:

7

ANSWER:

2

Correct

In this compound, Br has a total of ten electrons around it (not eight), which is why it is an exception to the octet rule.

Part C

The hydrogen atom is not actually electronegative enough to form bonds to xenon. Were the xenon-hydrogen bond to exist, what would be the structure of $\mathrm{XeH_4}$?

Double-click any atom and type Xe to change the label.

Draw the molecule by placing atoms on the grid and connecting them with bonds. Show all lone pairs of electrons.

Hint 1. Determine total valence

How many total valence electrons should be shown in this structure?

Express your answer numerically as an integer.

ANSWER:

12

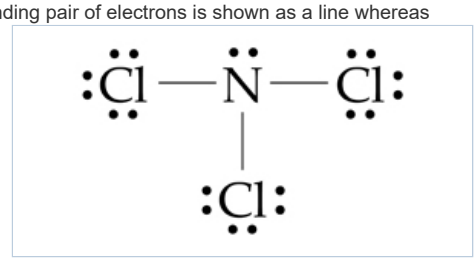
ANSWER:

Correct

In this compound, Xe has a total of twelve electrons around it (not eight), which is why it is an exception to the octet rule. Hydrogen is also an exception because is it stable with only two electrons rather than eight.

Lewis Structures and the Octet Rule

A Lewis structure shows the arrangement of atoms and valence electrons in a molecule. Most often, each bonding pair of electrons is shown as a line whereas nonbonding electrons are shown as dots as seen here. Notice that each atom is surrounded by a total of eight electrons, an octet.



Octet rule

For elements in neutral molecules, the number of bonds needed to achieve an octet follows a trend by group, as shown in the following table:

Group	4A	5A	6A	7A
Number of bonds	4	3	2	1
Number of lone pairs	0	1	2	3

Part A

Draw the Lewis structure for $\mathrm{SiCl_2Br_2}$.

To add the element Si, either double click on any atom and type the element symbol, or access a periodic table of elements from the More button.

Draw the molecule by placing atoms on the grid and connecting them with bonds. Include all lone pairs of electrons.

Hint 1. Steps in drawing a Lewis structure

1. Determine the total number of valence electrons.
2. Identify the central atom (usually the least electronegative element, but never H).
3. Give all outer atoms their normal number of bonds and lone pairs.
4. If there electrons left over, place them as lone pairs on the central atom.

Hint 2. Determine the number of valence electrons

How many total valence electrons should be used in the electron-dot structure for $\mathrm{SiCl_2Br_2}$?

Express your answer numerically as an integer.

Hint 1. How to determine the number of valence electrons

To calculate the total number of valence electrons in a formula, carry out the following steps:

1. Find the number of valence electrons for each atom in the formula.
2. Multiply by the subscript in the formula to account for the total from each type of atom. For example, in $\mathrm{N_2O}$, the total number of valence electrons from the nitrogen atoms would be $2 \times 5 = 10$.
3. If the species is negatively charged, add a number of electrons equal to the ionic charge. For example, for $\mathrm{ClO_4^-}$, one electron would be added.
4. If the species is positively charged, subtract a number of electrons equal to the ionic charge. For example, for $\mathrm{NH_4^+}$, one electron would be subtracted.

ANSWER:

32 valence electrons

Hint 3. Identify the central atom

In the compound $\mathrm{SiCl_2Br_2}$, which is the central atom?

Hint 1. How to identify the central atom

Here are some general guidelines to help identify the central atom in simple inorganic compounds and ions. The central atom

- is usually the least electronegative,
- is never hydrogen (H),
- usually forms the larger number of bonds, and
- frequently has a subscript of one.

ANSWER:

- Si
- Cl
- Br

Hint 4. Determine the number of bonds on the central atom

How many bonds does Si form in $\mathrm{SiCl_2Br_2}$?

ANSWER:

- 2
- 3
- 4

ANSWER:

Correct

Exceptions to the octet rule

Exceptions to the octet rule include hydrogen, which is stable with only two electrons (one bond), and boron, which is stable with only six electrons (three bonds). Additionally, elements in period 3 and beyond sometimes form more than the typical number of bonds, thus expanding their valence shells to more than eight electrons.

Part B

Consider the following four molecules. Which of these satisfy the octet rule and which do not?

Drag the appropriate items to their respective bins.

Hint 1. How to approach the problem

To determine whether the octet rule is followed, the number of electrons around each atom in the formulas must be identified. This can best be done by drawing the Lewis structure for each atom. Here is a set of guidelines to follow in drawing electron-dot structures.

1. Determine the total number of valence electrons using the number of valence electrons for each atom in the formula
2. If the species is a cation, subtract a number of electrons equal to the ionic charge. If the species is an anion, add a number of electrons equal to the ionic charge.
3. The central atom is usually the more electropositive element. This excludes H, which is never a central atom.
4. Bond all atoms in the appropriate skeletal arrangement so that all atoms (except H) have an octet (or their normal number of bonds). If there are not enough electrons to accomplish this, use double or triple bonds to achieve the octet.
5. If there are electrons left over, place the electrons as lone pairs on the central atom if the atom is in period 3 or higher.

Hint 2. Determine the Lewis structure for PF_5

Construct the Lewis structure for $\mathrm{PF_5}$.

Draw the molecule by placing atoms on the grid and connecting them with bonds. Include all lone pairs of electrons.

Hint 1. How to approach the problem

1. Determine the total number of valence electrons.
2. Identify the central atom (usually the least electronegative element, but never H).
3. Give all outer atoms their normal number of bonds and lone pairs.
4. If there are electrons left over, place them as lone pairs on the central atom.

ANSWER:

Hint 3. Determine the Lewis structure for CS_2

Construct the Lewis structure for $\mathrm{CS_2}$.

Draw the molecule by placing atoms on the grid and connecting them with bonds. Include all lone pairs of electrons.

Hint 1. How to approach the problem

1. Determine the total number of valence electrons.
2. Identify the central atom (usually the least electronegative element, but never H).
3. Give all outer atoms their normal number of bonds and lone pairs.
4. If there are electrons left over, place them as lone pairs on the central atom.

ANSWER:

Hint 4. Determine the Lewis structure for BBr_3

Construct the Lewis structure for $\mathrm{BBr_3}$.

Draw the molecule by placing atoms on the grid and connecting them with bonds. Include all lone pairs of electrons.

Hint 1. How to approach the problem

1. Determine the total number of valence electrons.
2. Identify the central atom (usually the least electronegative element, but never H).
3. Give all outer atoms their normal number of bonds and lone pairs.
4. If there are electrons left over, place them as lone pairs on the central atom.

ANSWER:

Hint 5. Determine the Lewis structure for CO_3^{2-}

Construct the Lewis structure for $\mathrm{CO_3}^{2-}$. You do *not* need to include brackets or charges.

Draw the molecule by placing atoms on the grid and connecting them with bonds. Include all lone pairs of electrons but do not include charges.

Hint 1. How to approach the problem

Drawing the structure of a polyatomic ion can be tricky. Use these steps as a guide.

1. Determine the total number of valence electrons. Don't forget to include the charge (a -1 charge means there is one extra electron).
2. Identify the central atom (usually the least electronegative element, but never H).
3. Start by giving all the outer atoms single bonds and then add lone pairs to the outer atoms until they all have an octet.
4. If there electrons left over, place them as lone pairs on the central atom.
5. If the central atom lacks an octet, turn one of the single bonds into a double bond and remove one lone pair from that outer atom.

ANSWER:

ANSWER:

Reset Help

Obeys octet rule

CO_3^{2-} CS_2

Violates octet rule

PF_5 BBr_3

Correct

Chemical Bonding

Learning Goal:

To understand bonding and be able to distinguish among covalent, ionic, and metallic substances.

A chemical formula such as H_2O indicates that two atoms of hydrogen and one atom of oxygen are bonded together to form a molecule of water. But what exactly is a bond? There are three main types of bonding:

1. *Covalent bonding* occurs between nonmetals.
2. *Metallic bonding* occurs between metals.
3. *Ionic bonding* occurs between a cation and an anion (often a metal and a nonmetal).

Nonmetals must gain electrons to achieve a noble-gas configuration. By sharing electrons with each other through covalent bonding, nonmetals can effectively gain the electrons they need to achieve stability.

Metals must lose electrons and form cations to achieve a noble-gas configuration. The electrons that are lost form what is often called a "sea of electrons" to which the positively charged metal ions are attracted. This is called *metallic bonding*.

The attraction of an anion to a cation results in ionic bonding. Since metals form cations and nonmetals form anions, ionic compounds are often composed of a metal and a nonmetal.

Part A

Carbon tetrafluoride (CF_4) is used as a low-temperature refrigerant in refrigerators. What type of bonding occurs between the atoms of a CF_4 molecule?

Hint 1. Classify carbon

Is carbon a metal or a nonmetal?

ANSWER:

- metal
- nonmetal

Hint 2. Classify fluorine

Is fluorine a metal or a nonmetal?

ANSWER:

- metal
- nonmetal

ANSWER:

- covalent
- metallic
- ionic

Correct

Part B

What type of bonding occurs in zinc(II) chloride, $ZnCl_2$?

Hint 1. Classify zinc

Is zinc a metal or a nonmetal?

ANSWER:

- metal
- nonmetal

Hint 2. Classify chlorine

Is chlorine a metal or a nonmetal?

ANSWER:

- metal
- nonmetal

ANSWER:

- covalent
- metallic
- ionic

Correct

Part C

What type of bonding occurs in a sample of pure sodium, Na? In other words, how is one sodium atom held to another sodium atom?

Hint 1. Classify sodium

Is sodium a metal or a nonmetal?

ANSWER:

- metal
- nonmetal

ANSWER:

- covalent
- metallic
- ionic

Correct

Pure sodium, Na, is a metal. It is gray and shiny just like many other metals that you may be more familiar with. NaCl is table salt, an ionic compound made up of sodium and chlorine. There is a huge difference between the properties of Na and NaCl because of the type of bonding that occurs.

Pure silver, Ag, is also a metal. But when silver bonds with chlorine, the result is AgCl, which looks very much like salt. Again, we see that the properties of a substance change considerably as the type of bonding changes.

Introduction to Lewis Structures

Learning Goal:

To understand how Lewis dot structures are used to show valence electrons and covalent bonding.

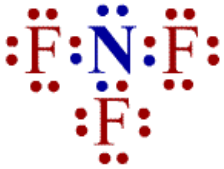
Covalent bonds involve the sharing of *valence electrons*. Valence electrons are those in the outermost subhsells of an atom. Nitrogen, for example, has an electron configuration of $1s^2 2s^2 2p^3$. Therefore, nitrogen has five valence electrons (two in the $2s$ and three in the $2p$).

Lewis dot structures show all the valence electrons in an atom or a molecule using dots. The Lewis dot structure of nitrogen could be drawn like this:

Five dots are drawn around N because nitrogen has five valence electrons. Also, since nitrogen must gain three electrons to achieve a noble-gas configuration, it will usually form three covalent bonds. That is why three of the dots are shown unpaired (ready for bonding) while the other two dots are shown paired (nonbonding).

Fluorine, with electron configuration $1s^2 2s^2 2p^5$, has seven valence electrons (two in the $2s$ subshell and five in the $2p$). Fluorine only needs to gain one electron to achieve a noble-gas configuration. Therefore, the Lewis dot structure of F should only show one unpaired electron, while the rest should be paired:

The Lewis dot structure of the molecule NF_3 is drawn as follows:



Each bonding electron from nitrogen is paired with the bonding electron from a fluorine atom. Notice that each atom is now surrounded by eight dots, called an *octet*. An octet is desirable because it means that the outermost *s* and *p* subshells are full; in other words, they have a noble-gas configuration.

Part A

Which of the following is the best Lewis dot structure for oxygen?

a b c d e f

Hint 1. How to approach the problem

The Lewis dot structure of an atom should show a dot for each valence electron. The number of unpaired electrons in the structure corresponds to the number of bonds that atom will make. Any leftover valence electrons are then drawn in pairs.

Hint 2. Count valence electrons

How many valence electrons does an oxygen atom have?

Express your answer as an integer.

Hint 1. How to determine valence electrons

Valence electrons are those in the outermost subshells. You will notice that the number of valence electrons in an atom corresponds to the "A" group number in which the element is found. For example, group 1A (group 1) elements have one valence electron. group 3A (group 13) elements have three valence electrons, and so on.

ANSWER:

6

Hint 3. Determine the number of bonds

How many bonds does oxygen usually form?

Express your answer as an integer.

Hint 1. How to approach the problem

Since a covalent bond gives an atom a share in one extra electron, the number of bonds a particular element usually forms corresponds to the number of electrons it needs to gain.

Hint 2. Count electrons

How many electrons does oxygen need to gain to achieve a noble-gas configuration?

Express your answer as an integer.

ANSWER:

2

ANSWER:

2

ANSWER:

- a
- b
- c
- d
- e
- f

Correct

Part B

How many electrons should be shown in the Lewis dot structure for hydrogen?

Express your answer as an integer.

Hint 1. Count valence electrons

How many valence electrons does a hydrogen atom have?

Express your answer as an integer.

Hint 1. How to determine valence electrons

Valence electrons are those in the outermost subshells. You will notice that the number of valence electrons in an atom corresponds to the "A" group number in which the element is found. For example, group 1A (group 1) elements have one valence electron. group 3A (group 13) elements have three valence electrons, and so on.

ANSWER:

1

ANSWER:

1

Correct

With only one valence electron, hydrogen can only form one bond. Hydrogen never forms an octet because it is stable once it achieves helium configuration (two total electrons).

Part C

Which of the following is the best Lewis dot structure for bromine?

a b c d e f

Hint 1. How to approach the problem

The Lewis dot structure of an atom should show a dot for each valence electron. The number of unpaired electrons in the structure corresponds to the number of bonds that atom will make. Any leftover valence electrons are then drawn in pairs.

Hint 2. Count valence electrons

How many valence electrons does a bromine atom have?

Express your answer as an integer.

Hint 1. How to determine valence electrons

Valence electrons are those in the outermost subshells. You will notice that the number of valence electrons in an atom corresponds to the "A" group number in which the element is found. For example, group 1A (group 1) elements have one valence electron, group 3A (group 13) elements have three valence electrons, and so on.

ANSWER:

7

Hint 3. Determine the number of bonds

How many bonds does bromine usually form?

Express your answer as an integer.

Hint 1. How to approach the problem

Since a covalent bond gives an atom a share in one extra electron, the number of bonds a particular element usually forms corresponds to the number of electrons it needs to gain.

Hint 2. Count electrons

How many electrons does bromine need to gain to achieve a noble-gas configuration?

Express your answer as an integer.

ANSWER:

1

ANSWER:

1

ANSWER:

- a
- b
- c
- d
- e
- f

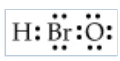
Correct

The Lewis dot structure for bromine looks just like the dot structure for fluorine because they have the same number of valence electrons. In general, all elements in a group behave similarly in chemical bonds.

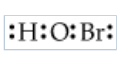
Part D

Based on your answers to Parts A, B, and C select the best dot structure for HOBr.

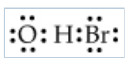
a



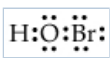
b



c



d



Hint 1. Determine the central atom

A general rule is that the least electronegative atom will be the central atom. Another strategy is to choose the atom that needs to make the greatest number of bonds. Earlier, you determined that O will form 2, whereas H and Br will each only form 1 bond.

Which element should be the central atom?

ANSWER:

- H
- O
- Br

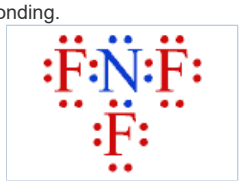
ANSWER:

- a
- b
- c
- d

Correct

Lewis Structures for Covalent Compounds

Lewis dot structures show all the valence electrons in an atom or molecule. Lewis dot structures are most often used to show covalent bonding. For example, the dot structure of NF_3 is shown in the figure, which shows a total of 26 valence electrons (6 bonding and 20 nonbonding).



Part A

Determine the number of bonding electrons and the number of nonbonding electrons in the structure of CO_2.

Enter the number of bonding electrons followed by the number of nonbonding electrons in the dot structure of this molecule separated by a comma (e.g., 1, 2).

Hint 1. How to approach the problem

You should use a pencil and paper to try to draw the structure of this molecule. Once you have drawn the structure, you just have to count the electrons. Each bond counts as two bonding electrons. Each lone pair counts as two nonbonding electrons.

Hint 2. Identify the central atom

Which element is the central atom in CO_2?

Hint 1. How to approach the problem

The central atom in a molecule is usually the one with the lowest electronegativity.

ANSWER:

- C
- O

Hint 3. What is the bonding arrangement?

Which bonding arrangement is most appropriate? Lone pairs have been omitted for simplicity.

ANSWER:

- O–C–O
- O=C–O
- O=C=O

Hint 4. Count the valence electrons

How many electrons should be shown in the structure of CO_2?

Express your answer numerically as an integer.

Hint 1. How to approach the problem

You should look at the periodic table to determine the valence for each element. Then, you should add those valence numbers together. Keep in mind that there are two oxygen atoms in a carbon dioxide molecule.

Hint 2. What is the valence of carbon?

How many valence electrons does one carbon atom have?

Express your answer numerically as an integer.

ANSWER:

4

Hint 3. What is the valence of oxygen?

How many valence electrons does one oxygen atom have?

Express your answer numerically as an integer.

ANSWER:

6

ANSWER:

16

ANSWER:

8,8

Correct

Part B

Draw the main Lewis structure of NOF.

Draw nonbonding electrons using the dot notation and bonding electrons as a bond.

Hint 1. Identify the central atom

Which atom is the central atom?

ANSWER:

- O
- F
- N

Hint 2. Determine the total number of valence electrons

How many valence electrons are in a molecule of NOF?

Express your answer numerically as an integer.

ANSWER:

18

ANSWER:

Correct

Part C

Determine the number of bonding electrons and the number of nonbonding electrons in the structure of $\mathrm{SeCl_2}$.

Enter the number of bonding electrons followed by the number of nonbonding electrons, separated by a comma, in the dot structure of this molecule (e.g., `1,2`).

Hint 1. How to approach the problem

You should use a pencil and paper to try to draw the structure of this molecule. Once you have drawn the structure, you just have to count the electrons. Each bond counts as two bonding electrons. Each lone pair counts as two nonbonding electrons.

Hint 2. Identify the central atom

Which element is the central atom in $\mathrm{SeCl_2}$?

Hint 1. How to approach the problem

The central atom in a molecule is usually the one with the lowest electronegativity.

ANSWER:

- Se
- Cl

Hint 3. Determine the bonding arrangement

Which bonding arrangement is most appropriate? Lone pairs have been omitted for simplicity.

ANSWER:

- $\mathrm{Cl-Se-Cl}$
- $\mathrm{Cl=Se-Cl}$
- $\mathrm{Cl=Se=Cl}$

Hint 4. Count the valence electrons

How many total electrons should be shown in the structure of $\mathrm{SeCl_2}$?

Express your answer numerically as an integer.

Hint 1. What is the valence of selenium?

How many valence electrons does selenium have?

Express your answer numerically as an integer.

ANSWER:

6

Hint 2. What is the valence of chlorine?

How many valence electrons does chlorine have?

Express your answer numerically as an integer.

ANSWER:

7

ANSWER:

20

ANSWER:

4,16

Correct

Periodic Trends in Electronegativity

Learning Goal:

To understand the properties of the elements that are associated with electronegativity, including oxidation-reduction potential and conductivity, and to identity an element as a metal or nonmetal.

Electronegativity is a value that is determined for each element according to its electron affinity and its ionization energy. It is useful in characterizing the chemistry of the elements. Electronegativities are often given as Pauling values (after the famous scientist Linus Pauling). Pauling arbitrarily set the value for the most electronegative element, fluorine, at 4.0. Note that electronegativity values tend to *increase* going across the rows of the periodic table and to *decrease* going down the columns of the periodic table.

The strongly electronegative elements, found in the upper right corner of the periodic table, have a tendency to gain electrons in chemical reactions and are therefore good oxidizers. Because they tend to hold electrons, these elements are good insulators (poor conductors) and are characterized as nonmetals. Nonmetals generally have Pauling electronegativity values of $\chi \geq 2.2$.

The weakly electronegative elements, making up much of the left side and middle of the table, have a tendency to lose electrons in chemical reactions and are therefore good reducing agents. Because they tend to lose electrons, these elements are good conductors of electricity and are characterized as metals. Metals generally have Pauling electronegativity values of $\chi \leq 1.8$.
There is a group of elements that have some properties of metals and some of nonmetals. These semimetals (also called metalloids) are characterized by intermediate Pauling electronegativity values of $\chi \approx 2$.

Part A

Classify each element as a metal, nonmetal, or semimetal.

Drag the appropriate items to their respective bins.

Hint 1. How to approach the problem

An element can be classified as a metal, nonmetal, or semimetal (metalloid) based on its electronegativity value. Metals have Pauling electronegativity values of $\chi \leq 1.8$ (generally low electronegativity values), nonmetals have Pauling electronegativity values of $\chi \geq 2.2$ (generally high electronegativity values), and semimetals have Pauling electronegativity values of $\chi \approx 2$.

Use trends in the electronegativity values on the periodic table to classify each element as a metal, nonmetal, or semimetal (metalloid).

Hint 2. Identify the periodic trends of electronegativity on the periodic table

Complete each statement about the general trends in electronegativity on the periodic table.

Drag the appropriate items to their respective bins.

ANSWER:

Reset Help

Electronegativity increases...

...as you move up a group.

...as you move to the right across a period.

Electronegativity decreases...

...as you move down a group.

...as you move to the left across a period.

ANSWER:

Reset Help

Metal

Zn Mg Tl

Nonmetal

Ar F

Semimetal

At

Correct

Part B

Arrange the following elements in order of decreasing electronegativity.

Rank from most to least electronegative. To rank items as equivalent, overlap them.

Hint 1. Identify the periodic trends of electronegativity on the periodic table

Complete each statement about the general trends in electronegativity on the periodic table.

Drag the appropriate items to their respective bins.

ANSWER:

Reset Help

Electronegativity increases...

...as you move up a group.

...as you move to the right across a period.

Electronegativity decreases...

...as you move down a group.

...as you move to the left across a period.

ANSWER:

Reset Help

Most electronegative Least electronegative

nitrogen sulfur manganese rubidium

The correct ranking cannot be determined.

Correct

Part C

An oxidant is a substance that can accept the electrons from another reagent. Use electronegativity values to determine which one of the following elements is a good oxidant.

Hint 1. How to approach the problem

The electronegativity of an element determines its properties; highly electronegative elements tend to be nonmetals that are poor conductors of electricity. Because nonmetals tend to gain electrons in chemical reaction, they are oxidants (which are reduced in a chemical reaction).

ANSWER:

- F
- Rb
- Fe

Correct

Part D

Use electronegativity values to determine which of the following elements you would expect to be the best insulator.

Hint 1. Electronegativity of insulators

Elements that are more electronegative tend to be better insulators. Although semimetals have some ability to act as insulators, the best insulators (and thus the worst conductors of electricity) are the nonmetals.

ANSWER:

- Mg
- As
- C

Correct

Periodic Trends of the Elements

Learning Goal:

To understand the periodic nature of the properties of the elements.

The periodic table is an organized listing of all the known elements. The table displays elements in order of their atomic number and groups them in vertical and horizontal rows that indicate similar properties.

Several properties of atoms are *periodic*, meaning that the trends repeat themselves from row to row. This gives the periodic table its predictive value. In fact, several elements were predicted to exist before they were actually discovered, because there seemed to be gaps when the known elements were first accurately grouped into a periodic table

Atomic Radius

Atomic radius is one periodic property of the elements. Atoms become smaller from left to right across a row, but they become larger going down a column. The largest atoms are thus found in the bottom left corner of the periodic table.

Part A

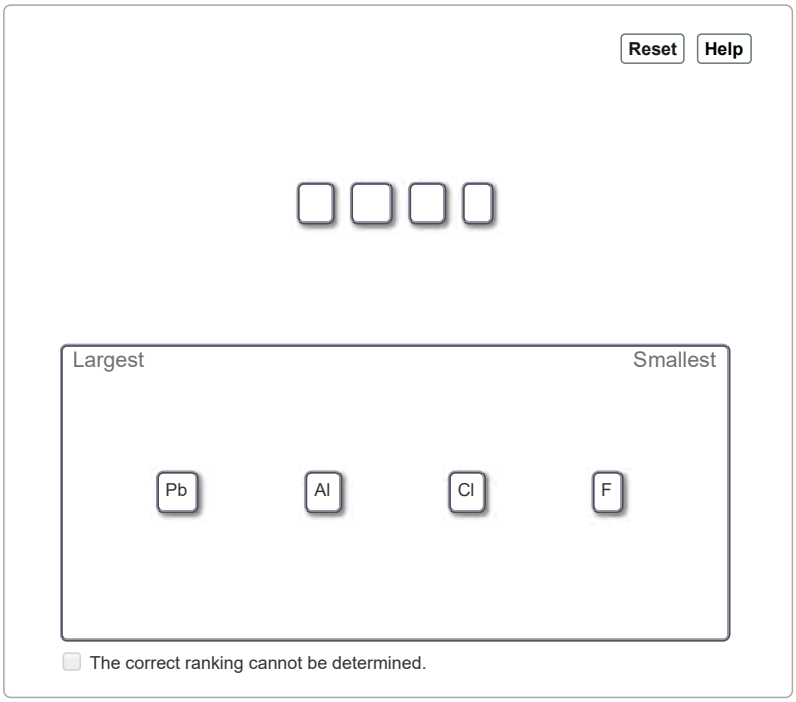
Place the following elements in order of decreasing atomic radius: chlorine (Cl), lead (Pb), aluminum (Al), and fluorine (F).

Rank the atomic radius from largest to smallest. To rank items as equivalent, overlap them.

Hint 1. Periodic trends in atomic size

Keep in mind that the size of atoms tends to increase as you go down a column, but it decreases as you go across a row. Therefore, the largest atoms should be found at the bottom left, whereas the smallest atoms should be found at the top right of the periodic table.

ANSWER:



Correct

Ionization energy

Ionization energy is another periodic property of the elements. The ionization energy of an atom is the energy required to remove an electron when the element is in a gaseous state. Ionization energies tend to increase from left to right across the row, but they decrease going down a column. The highest ionization energies are thus found in the upper right corner of the periodic table. Some sample ionization energies are shown in the table below.

Element	Group	Ionization energy (kJ/mol)
Li	1	520
Cs	1	382
He	18	2377
Rn	18	1042

Part B

A Venus probe detects an element with a first ionization energy of 2088 $\mathrm{kJ/mol}$. Given the table of known ionization energies provided in the introduction, what is the most likely identity of this element?

Hint 1. Determine the group from the ionization energy

Elements within the same group in the periodic table have similar properties. Based on the given ionization energy value, what group is this element in?

Express the group number as an integer from 1 to 18.

ANSWER:

18

ANSWER:

- Fe
- Rb
- Ne
- Hg

Correct

Electron affinity

The *electron affinity* of atoms is also periodic. The electron affinity of an atom is the energy that is lost or gained when an electron is added to a neutral atom of the element in a gaseous state. A more negative electron affinity is considered a higher electron affinity value. It tends to become more negative going from left to right across a row in the table and less negative going down a column. Electron affinities are not as smoothly periodic as some of the other properties, with some major exceptions opposing the general trends. For example, the group 2 and group 18 elements have positive electron affinities, and each of the row 2 elements B through F are less negative than the elements just below them in the column. In general, however, electron affinities tend to be more negative as you move toward the upper right-hand corner of the periodic table.

Part C

Which element has the highest (most negative) electron affinity?

Hint 1. How to approach the problem

A positive electron affinity means that energy is required to add an electron. In other words, the element does not readily gain electrons. A negative electron affinity means that energy is released when an electron is added. In other words, the element readily gains electrons. Therefore, you are looking for an element that forms a negatively charged ion.

ANSWER:

- ○ Ar
- ◉ S
- ○ Cs
- ○ Ca
- ○ Cr

Correct

Electronegativity

Electronegativity is another periodic property. An atom's electronegativity describes its ability to attract electrons to itself when it is part of a chemical compound. Electronegativity increases diagonally from the lower left to the upper right of the periodic table. Highly electronegative elements (with a *chi* value of $\chi \geq 2.2$) are insulating nonmetals, whereas elements with low electronegativity (with $\chi \leq 1.8$) are conducting metals.

Part D

A probe sent to the planet Mercury has measured the electronegativities of several elements it has detected on the planet.

Element	1	2	3	4	5
Electronegativity, χ	2.19	1.31	2.18	3.00	3.98

Which element would you expect to identify as a metal?

Hint 1. How to approach the problem

Metals have low electronegativities. More specifically, a Pauling electronegativity value of 1.9 or less identifies the element as a metal.

ANSWER:

- ○ element 1
- ◉ element 2
- ○ element 3
- ○ element 4
- ○ element 5

Correct

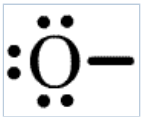
The element with an electronegativity value of 1.31 is magnesium (Mg), which is indeed a metal. Arsenic ($\chi = 2.18$) is a semimetal, and fluorine ($\chi = 3.98$), krypton ($\chi = 3.00$), and phosphorus ($\chi = 2.19$) are nonmetals.

Structures of Oxoanions

An *oxoanion* is a polyatomic ion that contains oxygen and has a net negative charge. Examples are $\mathrm{SO_4}^{2-}$, $\mathrm{NO_3}^{-}$, and $\mathrm{CO_3}^{2-}$.

In many oxoanions, the oxygen atoms form one single bond with three lone pairs, giving them each a formal charge of -1. However, if this bonding arrangement does not give the central atom an octet of electrons, multiple bonds form between the central atom and one or more of the outer oxygen atoms.

An oxoanion may combine with a cation to form an ionic compound. Examples are K_2SO_4, $NaNO_3$, and $CaCO_3$.



Part A

Draw the structure of NO_3^-. Include all lone pairs of electrons, but do *not* include the ionic charge.

Draw the ion by placing atoms on the grid and connecting them with bonds. Include all lone pairs of electrons.

Hint 1. How to approach the problem

Start by connecting each oxygen atom to the central nitrogen atom with a single bond. Second, put three lone pairs around each oxygen atom. Third, calculate the total valence needed in the molecule (be sure to include the charge), and add lone pairs to the central atom as needed. Fourth, check to see that the central atom has an octet. Finally, if the central atom does not have an octet, change one oxygen atom from having a single bond and three lone pairs to having a double bond and two lone pairs. Continue this process on each oxygen atom until the central atom has an octet.

Hint 2. Calculate the total number of valence electrons

How many electrons should be shown in the Lewis structure of NO_3^-?

Express your answer numerically as an integer.

Hint 1. How to find the total number of valence electrons

Sum the number of valence electrons in one nitrogen atom and three oxygen atoms. Then include one more electron to account for the -1 charge.

The valence of an atom includes the outermost s and p electrons. Nitrogen, which has an electron configuration of $1s^2 2s^2 2p^3$, has five valence electrons. Oxygen, which has an electron configuration of $1s^2 2s^2 2p^4$, has six valence electrons.

ANSWER:

24

Hint 3. Determine whether the central atom has an octet

With all single bonds (and three lone pairs on each outer oxygen atom), which of the following would be true of the central nitrogen atom, N, in NO_3^-?

ANSWER:

The central nitrogen atom

- has less than an octet of electrons.
- has more than an octet of electrons.
- has an octet of electrons.

ANSWER:

Correct

In this structure, the double bond between nitrogen and oxygen could be placed between nitrogen and any one of the three oxygen atoms. The NO_3^- ion can be represented by three equivalent Lewis structures and is, therefore, said to exhibit *resonance*.

Part B

Draw the structure of CO_3^{2-}. Include all lone pairs of electrons, but do *not* include the ionic charge.

Draw the ion by placing atoms on the grid and connecting them with bonds. Include all lone pairs of electrons.

Hint 1. How to approach the problem

Start by connecting each oxygen atom to the central atom with a single bond. Second, put three lone pairs around each oxygen atom. Third, calculate the total valence needed in the molecule (be sure to include the charge), and add lone pairs to the central atom as needed. Fourth, check to see that the central atom has an octet. Finally, if the central atom does not have an octet, change one oxygen atom from having a single bond and three lone pairs to having a double bond and two lone pairs. Continue this process on each oxygen atom until the central atom has an octet.

Hint 2. Calculate the total number of valence electrons

How many total electrons should be shown in this structure?

Express your answer numerically as an integer.

Hint 1. How to find the total number of valence electrons

Sum the number of valence electrons in one carbon atom and three oxygen atoms. Then include two more electrons to account for the -2 charge.

The valence of an atom includes the outermost s and p electrons. Carbon, which has an electron configuration of $1s^2 2s^2 2p^2$, has four valence electrons. Oxygen, which has an electron configuration of $1s^2 2s^2 2p^4$, has six valence electrons.

ANSWER:

24

Hint 3. Determine whether the central atom has an octet

With all single bonds (and three lone pairs on each outer oxygen atom), which of the following would be true of the central carbon atom, C, in CO_3^{2-}?

ANSWER:

The central carbon atom
- has less than an octet of electrons.
- has an octet of electrons.
- has more than an octet of electrons.

ANSWER:

Correct

In this structure, the double bond between carbon and oxygen could be placed between carbon and any one of the three oxygen atoms. The CO_3^{2-} ion can be represented by three equivalent Lewis structures and is, therefore, said to exhibit *resonance*.

Molecular Geometry

The *molecular geometry* of a molecule describes the three-dimensional shape of just the atoms. This is in contrast to the electronic geometry, which describes the shape of all *electron regions*.

To determine the molecular geometry of a molecule, one must first determine the electronic geometry by drawing the Lewis structure.

Part A

What is the molecular geometry of PF_3?

Enter the molecular geometry of the molecule.

Hint 1. Determine the electronic geometry

What is the electronic geometry of PF_3?

ANSWER:

- linear
- trigonal planar
- tetrahedral
- trigonal bipyramidal
- octahedral

ANSWER:

trigonal pyramidal

Correct

Part B

What is the molecular geometry of CO_2?

Enter the molecular geometry of the molecule.

Hint 1. Determine the electronic geometry

What is the electronic geometry of CO_2?

ANSWER:

- (•) linear
- ( ) trigonal planar
- ( ) tetrahedral
- ( ) trigonal bipyramidal
- ( ) octahedral

ANSWER:

linear

Correct

Part C

What is the molecular geometry of SCl_2?

Enter the molecular geometry of the molecule.

Hint 1. Determine the electronic geometry

What is the electronic geometry of SCl_2?

ANSWER:

- ( ) linear
- ( ) trigonal planar
- (•) tetrahedral
- ( ) trigonal bipyramidal
- ( ) octahedral

ANSWER:

bent

Correct

Molecular Geometry II

The molecular geometry of a molecule describes the three-dimensional shape of just the atoms. This is in contrast to the electronic geometry, which describes the shape of all regions of high electron density. The number of regions of high electron density, or *steric number*, determines the shape of the molecule.

Part A

What is the molecular geometry of a molecule with 4 outer atoms and 1 lone pair on the central atom?

Enter the molecular geometry of the molecule.

Hint 1. Determine the electronic geometry

What is the electronic geometry of this molecule?

ANSWER:

- linear
- trigonal planar
- tetrahedral
- trigonal bipyramidal
- octahedral

ANSWER:

see saw

Correct

Part B

What is the molecular geometry of a molecule with 4 outer atoms and 2 lone pairs on the central atom?

Enter the molecular geometry of the molecule.

Hint 1. Determine the electronic geometry

What is the electronic geometry of this molecule?

ANSWER:

- linear
- trigonal planar
- octahedral
- trigonal bipyramidal
- tetrahedral

ANSWER:

square planar

Correct

Part C

What is the molecular geometry of a molecule with 2 outer atoms and 3 lone pairs on the central atom?

Enter the molecular geometry of the molecule.

Hint 1. Determine the electronic geometry

What is the electronic geometry of this molecule?

ANSWER:

- linear
- trigonal planar
- tetrahedral
- octahedral
- trigonal bipyramidal (selected)

ANSWER:

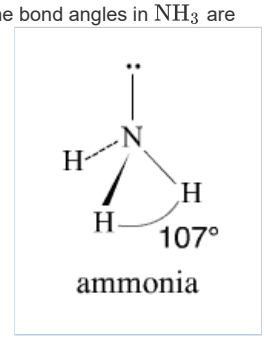
linear

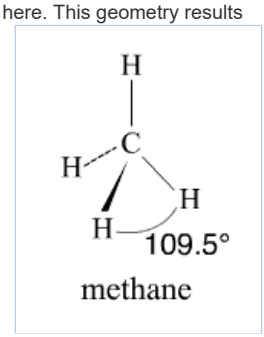
Correct

VSEPR

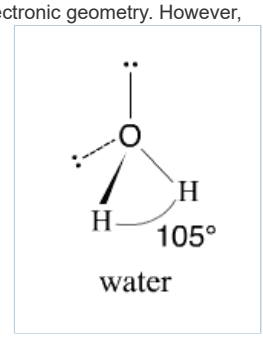
According to valence-shell electron-pair repulsion (VSEPR) theory, electron pairs around a central atom repel each other. This accounts for the geometries, or shapes, of molecules. For example, CH_4 is tetrahedral, with all four outer hydrogen atoms spaced 109.5° from each other, as shown here. This geometry results from the fact that the bonding electron pairs repel each other equally and therefore remain as far apart as possible.

NH_3, with three outer atoms and one lone pair, also has a tetrahedral electronic geometry. However, the bond angles in NH_3 are only 107°. This indicates that the magnitude of repulsion between a lone pair and a bonded pair of electrons must be different from the magnitude of repulsion between two bonded pairs of electrons.





H_2O, with two outer atoms and two lone pairs, also has a tetrahedral electronic geometry. However, the bond angles in H_2O are only 105°. This indicates that the magnitude of repulsion between two lone pairs must be be different from the magnitude of repulsion between a lone pair and a bonded pair of electrons.



If BP stands for a *bonded pair* of electrons and LP stands for a *lone pair* of electrons, then BP-BP represents the repulsion between two bonded pairs, LP-LP represents the repulsion between two lone pairs, and BP-LP represents the repulsion between a bonded pair and a lone pair of electrons.

Part A

Based on the bond angles in CH_4, NH_3, and H_2O, rank the magnitude of these repulsions.

Rank from strongest to weakest repulsion. To rank items as equivalent, overlap them.

Hint 1. Identify the trend in bond angle

ANSWER:

As we move from methane to ammonia to water, the addition of each lone pair appears to

- push the bonds together. (selected)
- pull the bonds apart.
- have no effect on bond angle.

ANSWER:

Reset Help

Strongest repulsion Weakest repulsion

LP-LP BP-LP BP-BP

The correct ranking cannot be determined.

Correct

Geometry, Bond Angles, and Polarity

Bromine pentaiodide, BrI_5, is used for research purposes in scientific laboratories. Clearly one would like to know quite a bit about the chemical nature of this compound. Therefore, to answer the questions in this problem, you will need to draw the Lewis structure of bromine pentaiodide.

Part A

What is the electron-domain (charge-cloud) geometry of BrI_5?

Enter the electron-domain geometry of the molecule.

Hint 1. Determine the steric number

The number of electron domains (also called charge clouds) around the central atom of a molecule is often called the *steric number*. What is the steric number of this molecule?

Express your answer as an integer.

Hint 1. How to approach the problem

The steric number is equal to the number of outer atoms plus the number of lone pairs on the central atom.

Hint 2. Determine the number of outer atoms

How many outer atoms are bonded to the central atom in this molecule?

Express your answer as an integer.

ANSWER:

5

Hint 3. Determine the number of lone pairs

How many lone pairs are on the central atom of this molecule?

Express your answer as an integer.

ANSWER:

1

ANSWER:

6

ANSWER:

octahedral

Correct

Part B

What is the molecular geometry of $\mathrm{BrI_5}$?

Enter the molecular geometry of the molecule.

Hint 1. How to approach the problem

The molecular geometry can be determined from the electronic geometry and the number of lone pairs on the central atom.

Hint 2. Determine the number of lone pairs

How many lone pairs are on the central atom of this molecule?

Express your answer as an integer.

ANSWER:

1

ANSWER:

square pyramid

Correct

Part C

Ignoring lone-pair effects, what is the smallest bond angle in $\mathrm{BrI_5}$?

Express your answer as an integer.

Hint 1. How to approach the problem

Knowing the electronic geometry of the molecule will allow you to determine the bond angles. Recall that the repulsion between the electron regions causes them to exist as far apart from each other as possible.

ANSWER:

90 degrees

Correct

Part D

Which choice best describes the polarity of $\mathrm{BrI_5}$?

Hint 1. Determine the electronegativity

Is there a difference in electronegativity between Br and I?

ANSWER:

- (•) Yes
- ( ) No

Hint 2. Determine the arrangement of the bonds

Based on the molecular geometry of this molecule, are the bonds arranged symmetrically about the central atom (such that bond polarities would cancel)?

ANSWER:

- Yes
- No

ANSWER:

- The molecule is polar and has polar bonds.
- The molecule is nonpolar and has polar bonds.
- The molecule is polar and has nonpolar bonds.
- The molecule is nonpolar and has nonpolar bonds.

Correct

Precipitation Reactions

Here are some general solubility rules:

1. Compounds of group I ions are soluble.
2. Nitrates and chlorates are soluble.
3. Except for the compounds covered by rule 1, carbonates, hydroxides, phosphates, and sulfides are insoluble.

Part A

Suppose that aqueous solutions of barium nitrate and potassium carbonate are mixed. What is the name of the compound or compounds that precipitate?

Enter the name of the precipitate.

Hint 1. How to approach the problem

A precipitate is an insoluble substance. Therefore, you need to find some combination of the given ions that is insoluble.

ANSWER:

barium carbonate

Correct

Part B

Insoluble sulfide compounds are generally black in color. Which of the following combinations could yield a black precipitate?

Check all that apply.

Hint 1. Identify insoluble sulfides

Which of the following sulfides are insoluble?

Select all that apply.

ANSWER:

- [x] tin (IV) sulfide
- [ ] potassium sulfide
- [ ] sodium sulfide
- [ ] lithium sulfide
- [x] lead (II) sulfide

ANSWER:

- [ ] $Na_2S(aq) + KCl(aq)$
- [x] $Li_2S(aq) + Pb(NO_3)_2(aq)$
- [ ] $Pb(ClO_3)_2(aq) + NaNO_3(aq)$
- [ ] $AgNO_3(aq) + KCl(aq)$
- [x] $K_2S(aq) + Sn(NO_3)_4(aq)$

Correct

Part C

Enter the net ionic equation for the reaction of aqueous sodium chloride with aqueous silver nitrate.

Express your answer as a chemical equation.

Hint 1. How to approach the problem

To write a net ionic equation, first write out the overall equation. Then, write the reaction as a complete ionic equation, showing which ions and molecules are hydrated and which are present in other forms and phases. Finally, cancel out spectator ions (species found on both sides of the equation), leaving the net ionic equation.

Hint 2. Determine the chemical formula of sodium chloride

Enter the chemical formula of aqueous sodium chloride.

Express your answer as a chemical formula.

ANSWER:

$NaCl(aq)$

Hint 3. Determine the chemical formula of silver nitrate

Enter the chemical formula of aqueous silver nitrate.

Express your answer as a chemical formula.

ANSWER:

$AgNO_3(aq)$

ANSWER:

$Cl^-(aq) + Ag^+(aq) \rightarrow AgCl(s)$

Correct

Part D

Imagine that A and B are cations and X, Y, and Z are anions, and that the following reactions occur:

$$AX(aq) + BY(aq) \rightarrow \text{no precipitate}$$

$$AX(aq) + BZ(aq) \rightarrow \text{precipitate}$$

Which of the following choices is insoluble?

Hint 1. Determine which products of the first reaction are soluble

Based on the fact that $AX + BY$ yields no precipitate, what can you determine about the solubilities of AY and BX?

ANSWER:

- AY is insoluble and BX is soluble.
- AY and BX are soluble.
- AY is soluble and BX is insoluble.
- AY and BX are insoluble.

Hint 2. Determine the common product of the reactions

Which product do both of these reactions have in common?

Enter two capital letters, for example, CZ.

ANSWER:

BX

ANSWER:

- AX
- AY
- AZ
- BX
- BY
- BZ

Correct

Acids, Bases, and Salts

An acid is a compound that can donate $\mathrm{H^+}$. You can recognize acid formulas as those that start with H or end with COOH.

A base is a compound that can accept $\mathrm{H^+}$. Base formulas often contain a metal cation with the hydroxide anion, $\mathrm{OH^-}$. Bases that do not contain OH usually contain N.
A salt is an ionic compound that is neither an acid nor a base. To identify a salt, look for a formula that contains a cation that is not $\mathrm{H^+}$ and an anion that is not $\mathrm{OH^-}$.

Some compounds that contain H are neither acid, nor base, nor salt. For example, hydrocarbons are *not* acids even though their formulas contain H. Hydrocarbons are easy to recognize since their formulas contain only C and H.

Part A

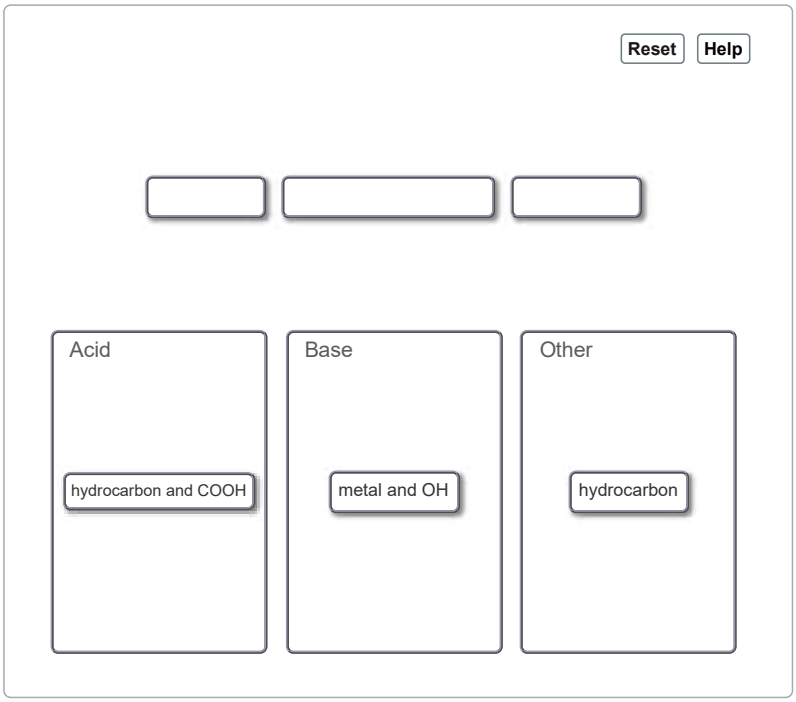
Classify these compounds as acid, base, salt, or other.

Drag the appropriate items to their respective bins.

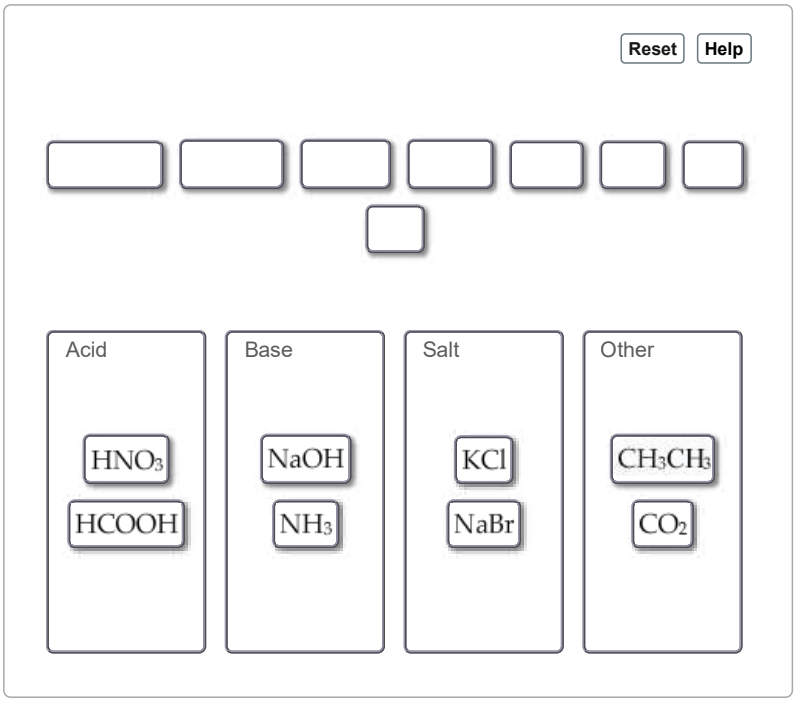
Hint 1. Distinguish formulas that contain H

A formula that contains H could be an acid, a base, or neither depending on the other elements in the formula. Sort these general formula descriptions by whether they indicate an acid, a base, or other compound.

ANSWER:



ANSWER:



Correct

In solution carbon dioxide, $\mathrm{CO_2}$, reacts with water to form a small amount of carbonic acid through the equilibrium

$$\mathrm{CO_2(aq) + H_2O(l) \rightleftharpoons H_2CO_3(aq)}$$

Carbonic acid is a weak acid, and thus solutions containing dissolved carbon dioxide are mildly acidic. Soda that has not lost its fizz will taste sour compared to soda that has gone "flat." The increase of dissolved carbon dioxide in the oceans is also leading to ocean acidification, and is an important environmental concern.

Solutions

Learning Goal:

To understand the concepts of solubility and dissociation.

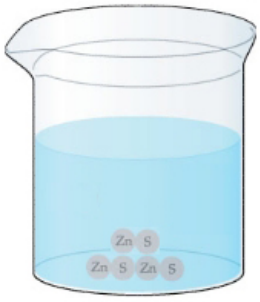
For an ionic solid to dissolve into a liquid, the ions must dissociate and become surrounded by solvent molecules. This is called *solvation*. When water is the solvent, it is called *hydration*. An example of a water-soluble ionic solid is table salt, NaCl. Salt-water can be written as $\mathrm{NaCl}(aq)$ indicating "aqueous." see the figure.

In contrast, an example of a water-insoluble ionic substance is zinc sulfide. Zinc sulfide does not dissociate in water, and therefore the ions cannot become hydrated. Its formula is written $\mathrm{ZnS}(s)$ where the *s* stands for "solid." See the figure.

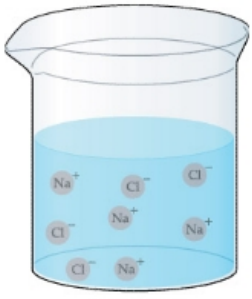
How can one tell whether an ionic compound is soluble? Chemistry texts usually contain a chart or list of guidelines for determining the solubility of almost any ionic compound.

Ionic compounds are not the only substances that can dissolve in liquids. Molecular compounds can also dissolve if the individual molecules can become solvated.

Based on this information, you can now understand the concept of electrolytes. An *electrolye* is any compound that can form ions when dissolved in water and therefore conduct electricity. Acids, bases, and soluble salts fit this categoy. NaCl is an electrolyte because it is a soluble salt and therefore dissociates. ZnS is not an electrolyte, because it is an insoluble salt and therefore does not dissociate. Sugar, even though it is soluble, is not an electrolyte because it does not dissociate into ions.



ZnS(*s*)
Insoluble



NaCl(*aq*)
Soluble

Part A

Substance X is a molecular compound and is not an acid nor a base. Based just on this information, what can you determine about substance X?

Check all that apply.

Hint 1. Solubility

This substance is not ionic and you do not know its identity. Therefore you cannot use a solubility chart to look up whether it is soluble or insoluble. Furthermore, you do not know anything about the intermolcular forces within this substance or between this substance and water. Therefore you cannot determine anything about the solubility of this substance.

Hint 2. What types of compounds are electrolytes?

Electrolytes are compounds that can form ions when dissolved in water. This allows the resulting solution to conduct electricity. Which of the following types of compounds dissociate to form ions when placed in water?

ANSWER:

- insoluble ionic compounds
- soluble molecular compounds
- insoluble molecular compounds
- soluble ionic compounds, acids, and bases

ANSWER:

- [ ] X is soluble in water.
- [ ] X is insoluble in water.
- [ ] X is an electrolyte.
- [x] X is not an electrolyte.

Correct

You do not know whether substance X is soluble in water. You do know that it cannot form ions. From this information you know that substance X cannot possibly be an electrolyte.

Part B

Which of the following situations is most favorable for solubility?

Hint 1. How to approach the problem

In order for a substance to dissolve, the molecules or ions must become solvated. Think about which situation would be most conducive to solvation.

ANSWER:

- strong attraction between solute particles
- strong attraction between solvent particles
- strong attraction of solute particles to solvent particles

Correct

Part C

Which of the following would you least expect to find in $\mathrm{CaCl_2}\,(aq)$?

Hint 1. The meaning of (aq)

The symbol *aq* following a chemical formula indicates an aqueous solution of that substance. In other words, $\mathrm{CaCl_2}\,(aq)$ means calcium chloride dissolved in water.

Hint 2. Is calcium chloride ionic or molecular?

Is $\mathrm{CaCl_2}$ ionic or molecular?

Hint 1. Definition of ionic and molecular compounds

Molecular compounds are composed exclusively of nonmetals. Ionic compounds consist of an anion and a cation, often a metal and a nonmetal.

ANSWER:

- ionic
- molecular

ANSWER:

- $\mathrm{H_2O}$ molecules
- $\mathrm{CaCl_2}$ molecules
- $\mathrm{Ca^{2+}}$ ions
- $\mathrm{Cl^-}$ ions

Correct

Oxidation States

Oxidation states are important for keeping track of electrons in oxidation-reduction reactions. Here are some general rules to remember:

- In most cases, oxygen has an oxidation state of -2.
- Group 1 and group 2 elements on the periodic table have +1 and +2 oxidation states, respectively.
- In most cases, hydrogen has an oxidation state of +1.
- Many elements can have more than one oxidation state. In such cases, use the other elements in the compound whose oxidation states are known to determine the oxidation state of the element that varies.

Neutral compounds

In a neutral compound, the *sum* of the oxidation states is zero. Note that the *sign* of the oxidation states and the *number* of atoms associated with each oxidation state must be considered. In $\mathrm{H_2O}$, for example, each hydrogen atom has an oxidation state of +1 and each oxygen atom has an oxidation state of -2 for a total of $2(+1)+(-2)=0$.

Part A

What is the oxidation state of an individual bromine atom in $\mathrm{NaBrO_3}$?

Express the oxidation state numerically (e.g., +1).

Hint 1. How to approach the problem

First, determine the oxidation state of individual Na and O atoms using the rules outlined in the introduction. Second, determine the total of the oxidation states for all of the Na and O atoms in the formula. Finally, consider that the oxidation state of the bromine atom must cancel out the sum of the other oxidation states.

If you prefer to work with equations, use x, y, and z to represent the oxidation states of Na, Br, and O, respectively. Then you can say that $x + y + 3z = 0$.

Hint 2. Determine the oxidation state of sodium

What is the oxidation state of the sodium atom in $\mathrm{NaBrO_3}$?

Express the oxidation state numerically (e.g., +1).

ANSWER:

+1

Hint 3. Determine the oxidation state of oxygen

What is the oxidation state of each *individual* oxygen atom in $\mathrm{NaBrO_3}$?

Express the oxidation state numerically (e.g., +1).

ANSWER:

-2

Hint 4. Determine the sum of the oxidation states of all the sodium and oxygen atoms

What is the sum of the oxidation states of all the Na and O atoms in the formula?

Express the oxidation state numerically (e.g., +1).

ANSWER:

-5

ANSWER:

+5

Correct

Part B

What is the oxidation state of an individual nitrogen atom in $\mathrm{KNO_3}$?

Express the oxidation state numerically (e.g., +1).

Hint 1. How to approach the problem

First, determine the oxidation state of individual O and N atoms using the rules outlined in the introduction. Second, determine the total of the oxidation states for all of the potassium and oxygen atoms in the formula. Finally, consider that the oxidation state of the nitrogen atom must cancel out the sum of the other oxidation states.

If you prefer to work with equations, use x, y, and z to represent the oxidation states of K,O, and N, respectively. Then you can say that $x + 1y + z + 2y = 0$.

Hint 2. Determine the oxidation state of potassium

What is the oxidation state of each *individual* potassium atom in $\mathrm{KNO_3}$?

Express the oxidation state numerically (e.g., +1).

ANSWER:

+1

Hint 3. Determine the oxidation state of oxygen

What is the oxidation state of the oxygen atom in $\mathrm{KNO_3}$?

Express the oxidation state numerically (e.g., +1).

ANSWER:

-2

Hint 4. Determine the sum of the oxidation states of all the potassium and oxygen atoms

What is the sum of the oxidation states of all the potassium and oxygen atoms in the formula?

Express the oxidation state numerically (e.g., +1).

ANSWER:

-5

ANSWER:

+5

Correct

Ions

In an ion, the *sum* of the oxidation states is equal to the overall ionic charge. Note that the *sign* of the oxidation states and the *number* of atoms associated with each oxidation state must be considered. In $\mathrm{OH^-}$, for example, the oxygen atom has an oxidation state of -2 and the hydrogen atom has an oxidation state of +1, for a total of $(-2)+(+1)=-1$.

Part C

What is the oxidation state of an individual sulfur atom in $\mathrm{SO_4}^{2-}$?

Express the oxidation state numerically (e.g., +1).

Hint 1. Derive the equation to solve for the oxidation state of sulfur

Since $\mathrm{SO_4}^{2-}$ has an overall charge of -2, the following equation applies:

$$x+4y=-2$$

where x is the oxidation state of an individual sulfur atom and y is the oxidation state of an individual oxygen atom. How can this equation be rearranged to solve for x?

Express your answer in terms of y.

ANSWER:

x = -9.83

Hint 2. Determine the oxidation state of oxygen

What is the oxidation state of each *individual* oxygen atom in the $\mathrm{SO_4}^{2-}$ ion?

Express the oxidation state numerically (e.g., +1).

ANSWER:

-2

ANSWER:

+6

Correct

Part D

What is the oxidation state of each *individual* carbon atom in $\mathrm{C_2O_4}^{2-}$?

Express the oxidation state numerically (e.g., +1).

Hint 1. Identify the equation to solve for the oxidation state of carbon

Since the $\mathrm{C_2O_4}^{2-}$ ion has an overall charge of -2, the following equation applies:

$$2x+4y=-2$$

where x is the oxidation state of an individual carbon atom and y is the oxidation state of an individual oxygen atom. How can this equation be rearranged to solve for x?

Express your answer in terms of y.

ANSWER:

$x = \frac{-2-4y}{2}$

Hint 2. Determine the oxidation state of oxygen

What is the oxidation state of each *individual* oxygen atom in the $C_2O_4^{2-}$ ion?

Express the oxidation state numerically (e.g., +1).

ANSWER:

-2

ANSWER:

+3

Correct

Classifying Reactions

Learning Goal:

To understand the different classes of reactions and be able to identify them.

Chemical reactions can be classified as one or more of the following types:

1. *Exchange* (*displacement*) reactions have the general formula $AB + C \rightarrow AC + B$ or $AB + CD \rightarrow AD + CB$.
2. *Condensation* (*combination*) reactions have the general formula $A + B \rightarrow AB$.
3. *Cleavage* (*decomposition*) reactions have the general formula $AB \rightarrow A + B$.
4. *Acid-base* (*neutralization*) reactions involve the transfer of H^+ ions.
5. *Oxidation-reduction* (*redox*) reactions involve changes in oxidation states.

Part A

Classify each reaction as an exchange reaction, a condensation reaction, or a cleavage reaction.

Drag the appropriate items to their respective bins.

Hint 1. How to identify exchange, condensation, and cleavage reactions

Exchange reactions involve the transfer of ions. In one scenario, two anion-cation pairs will "switch partners" such that each anion is combined with the opposite cation. In another scenario, only the anion is transferred from one element to another. To determine whether a reaction is classified as exchange, first determine which ions are present on both sides of the equation, then determine whether those ions have "switched partners." Condensation (combination) reactions involve the combining of reactants to produce one product, whereas cleavage (decomposition) reactions involve the decomposition (cleaving) of one reactant to produce products. Condensation reactions are the reverse of cleavage reactions, and cleavage reactions are the reverse of condensation reactions.

ANSWER:

Reset Help

Exchange

$Ca + 2HCl \longrightarrow CaCl_2 + H_2$

$2NH_4Cl + Ba(NO_3)_2 \longrightarrow 2NH_4NO_3 + BaCl_2$

$HCl + NaOH \longrightarrow NaCl + H_2O$

Condensation

$CaO + CO_2 \longrightarrow CaCO_3$

Cleavage

$2H_2O \longrightarrow 2H_2 + O_2$

Correct

Because acid-base reactions involve the transfer of H^+ ions, they are considered a type of exchange reaction.

Part B

Classify each reaction as a condensation reaction, a redox reaction, or both a condensation and redox reaction.

Drag the appropriate items to their respective bins.

Hint 1. Identify condensation reactions

Which reactions are condensation reactions?

Drag the appropriate items to their respective bins.

ANSWER:

Reset Help

Condensation

$CaO + CO_2 \longrightarrow CaCO_3$ $2Fe + 3Cl_2 \longrightarrow 2FeCl_3$

Not condensation

$2H_2O \longrightarrow 2H_2 + O_2$ $Ca + 2HCl \longrightarrow CaCl_2 + H_2$

Hint 2. Identify redox reactions

Which reactions are redox reactions?

Drag the appropriate items to their respective bins.

ANSWER:

Reset Help

Redox

$2H_2O \longrightarrow 2H_2 + O_2$

$Ca + 2HCl \longrightarrow CaCl_2 + H_2$

$2Fe + 3Cl_2 \longrightarrow 2FeCl_3$

Not redox

$CaO + CO_2 \longrightarrow CaCO_3$

ANSWER:

Reset Help

Condensation only

$CaO + CO_2 \longrightarrow CaCO_3$

Redox only

$2H_2O \longrightarrow 2H_2 + O_2$

$Ca + 2HCl \longrightarrow CaCl_2 + H_2$

Both condensation and redox

$2Fe + 3Cl_2 \longrightarrow 2FeCl_3$

Correct

Oxidation-Reduction Reactions

Oxidation-reduction reactions (often called "redox" for short) are reactions that involve the transfer of electrons from one species to another. Oxidation states, or oxidation numbers, allow chemists to keep track of these electron transfers. In general, one element will lose electrons (oxidation), with the result that it will increase in oxidation number, and another element will gain electrons (reduction), thereby decreasing in oxidation number.

An *oxidizing agent* is an element or compound in a redox reaction that oxidizes another species and itself gets reduced and is therefore the electron acceptor in the reaction.

A *reducing agent* is an element or compound in a redox reaction that reduces another species and itself gets oxidized and is therefore the electron donor in the reaction.

As a summary, keep in mind the following:

- Oxidation means an increase in oxidation state and a loss of electrons and involves a reducing agent.
- Reduction means a decrease in oxidation state and a gain of electrons and involves an oxidizing agent.

Part A

Which element is oxidized in this reaction?

$$Fe_2O_3 + 3CO \rightarrow 2Fe + 3CO_2$$

Enter the chemical symbol of the element.

Hint 1. How to approach the problem

Assign oxidation states to each element in the reaction to determine which elements gain or lose electrons. Also keep in mind that oxidation is the loss of electrons, which corresponds to an increase in oxidation state.

Hint 2. General rules for calculating oxidation states

Element classification	Oxidation state	Example
Free element	0	Al, Na, F, H_2, O_2, Cl_2, etc. have an oxidation state of zero.
Simple one atom	Equal to its charge	The oxidation state of Na^+ is +1, Be^{2+} is +2, and F^- is –1.
Group IA element	+1	The oxidation state of Li in $LiOH$ is +1. In Na_2SO_4, Na has a +1 oxidation state.
Group VIIA element (but see exception below)	–1	The oxidation state of Cl is –1 in HCl.
Group VIIA element combined with one having a higher electronegativity.	+1	The oxidation state of Cl is +1 in $HOCl$.
Group IIA element	+2	The oxidation state of Ca in $CaSO_4$ is +2.
Group IIIA element	+3	The oxidation state of Al in $AlCl_3$ is +3.
Hydrogen except hydrides	+1	The oxidation state of H in H_2SO_4 is +1.
Hydrides	–1	The oxidation state of H in LiH is –1.
Oxygen except peroxides	–2	The oxidation state of O in H_3PO_4 is –2.
Peroxides	–1	The oxidation state of O in H_2O_2 is –1, and in OF_2 it is +2.

For example, the oxidation state of sulfur in the sulfate ion (SO_4^{2-}) is calculated as follows.
Step 1: Allocate known charges:

$$x = \text{charge on sulfur}$$

$$\text{overall charge} = -2$$

$$\text{charge on each oxygen} = -2$$

Step 2: Write in the form of an equation:

$$\text{total charge on sulfur} + \text{total charge on oxygen} = \text{total charge on sulfate}$$

Substituting the values gives $x + (4 \times -2) = -2$ and so $x = +6$. The oxidation state of sulfur in SO_4^{2-} is therefore +6.

Hint 3. Find the oxidation states of iron

What are the oxidation states of iron in Fe_2O_3 and Fe, respectively?

Express your answer numerically separated by commas.

ANSWER:

Oxidation states of iron in Fe_2O_3, Fe = 3,0

Hint 4. Find the oxidation states of carbon

What are the oxidation states of carbon in CO and CO_2 respectively?

Express your answer numerically separated by commas.

ANSWER:

Oxidation states of C in CO, CO_2 = 2,4

ANSWER:

C is oxidized

Correct

Carbon gets oxidized in the process of reducing iron. Iron gets reduced in the process of oxidizing carbon.

Part B

Which substance is the *oxidizing agent* in this reaction?

$$Fe_2O_3 + 3CO \rightarrow 2Fe + 3CO_2$$

Express your answer as a chemical formula.

Hint 1. Define oxidizing agent

An oxidizing agent _____ another species in a chemical reaction, thereby _____ electrons in the reaction.

ANSWER:

- oxidizes/accepting
- reduces/donating
- reduces/accepting
- oxidizes/donating

ANSWER:

Fe_2O_3 is the oxidizing agent.

Correct

Notice that the oxidizing agent is itself reduced. The species oxidized is not the oxidizing agent, but the reducing agent. Think of the oxidizing agent as the species the acts upon another species causing oxidation.

Part C

Which element is reduced in this reaction?

$$2KMnO_4 + 3Na_2SO_3 + H_2O \rightarrow 2MnO_2 + 3Na_2SO_4 + 2KOH$$

Enter the chemical symbol of the element.

Hint 1. How to approach the problem

Assign oxidation states to each element in the reaction to determine which elements gain or lose electrons. Also keep in mind that reduction is the gain of electrons, which corresponds to a decrease in oxidation state.

Hint 2. General rules for calculating oxidation states

Element classification	Oxidation state	Example
Free element	0	Al, Na, F, H_2, O_2, Cl_2, etc. have an oxidation state of zero.
Simple one atom	Equal to its charge	The oxidation state of Na^+ is +1, Be^{2+} is +2, and F^- is –1.
Group IA element	+1	The oxidation state of Li in $LiOH$ is +1. In Na_2SO_4, Na has a +1 oxidation state.
Group VIIA element (but see exception below)	–1	The oxidation state of Cl is –1 in HCl.
Group VIIA element combined with one having a higher electronegativity.	+1	The oxidation state of Cl is +1 in $HOCl$.
Group IIA element	+2	The oxidation state of Ca in $CaSO_4$ is +2.
Group IIIA element	+3	The oxidation state of Al in $AlCl_3$ is +3.
Hydrogen except hydrides	+1	The oxidation state of H in H_2SO_4 is +1.
Hydrides	–1	The oxidation state of H in LiH is –1.
Oxygen except peroxides	–2	The oxidation state of O in H_3PO_4 is –2.
Peroxides	–1	The oxidation state of O in H_2O_2 is –1, and in OF_2 it is +2.

For example, the oxidation state of sulfur in the sulfate ion (SO_4^{2-}) is calculated as follows.
Step 1: Allocate known charges:

$$x = \text{charge on sulfur}$$

$$\text{overall charge} = -2$$

$$\text{charge on each oxygen} = -2$$

Step 2: Write in the form of an equation:

$$\text{total charge on sulfur} + \text{total charge on oxygen} = \text{total charge on sulfate}$$

Substituting the values gives $x + (4 \times -2) = -2$ and so $x = +6$. The oxidation state of sulfur in SO_4^{2-} is therefore +6.

Hint 3. Determine the elements that change oxidation state

Which two elements change oxidation state?

Check all that apply.

ANSWER:

- [x] Mn
- [ ] K
- [x] S
- [ ] Na

ANSWER:

Mn is reduced

Correct

Sulfur gets oxidized in the process of reducing manganese. Manganese gets reduced in the process of oxidizing sulfur.

Part D

Which substance is the *reducing agent* in this reaction?

$$2KMnO_4 + 3Na_2SO_3 + H_2O \rightarrow 2MnO_2 + 3Na_2SO_4 + 2KOH$$

Express your answer as a chemical formula.

Hint 1. Define reducing agent

A reducing agent _____ another species in a chemical reaction, thereby _____ electrons in the reaction.

ANSWER:

- ( ) oxidizes/donating
- (•) reduces/donating
- ( ) oxidizes/accepting
- ( ) reduces/accepting

ANSWER:

Na_2SO_3 is the reducing agent.

Correct

The reducing agent acts to reduce another species and therefore must itself be oxidized.

Score Summary:

Your score on this assignment is 117%.
You received 117.2 out of a possible total of 100 points.