Supplemental Activities

**Module**: Bonding

**Section**: Lewis Dot Structures
Lewis Structures & Octet Rule

Chemists follow several guidelines when visually representing covalent compounds. The conventional method of drawing structures produces Lewis structures. The following activities will insure that you practice the skills necessary to draw and interpret Lewis structures.

Activity 1

1. In a conventional Lewis structure a covalent bond is represented as a _______. A lone pair is represented as a _______ of _______. A free radical is represented as a single _______.

2. An important skill in predicting the correct Lewis structure of a molecule is to be able to count up the total number of valence electrons on a molecule. State the total number of valence electrons on the following molecules: C₃H₈, SO₂, I₂, PH₃ and CH₃OH.

3. What is the octet rule and how does it help us draw Lewis structures?

4. State a rule that might help when beginning to draw Lewis structures (everyone learns differently so this rule may or may not be helpful to you).

Activity 2

The purpose of this activity is to use the rules of Lewis Dot Structures to predict the structures of some common molecules. It’s important to remember that drawing Lewis structures is a process that sometimes requires trial and error. In later activities more skills will be added so that you can check the viability of your structure.

1. Draw the correct Lewis structures for the following molecules: OF₂, NH₂I and CH₃Cl.
2. Draw the correct Lewis structures for the following molecules: \( \text{N}_2, \text{N}_2\text{H}_2 \) and \( \text{N}_2\text{H}_4 \).

3. How do structural formulae help us draw Lewis structures? Compare and contrast these structural formulae: \( \text{CH}_3\text{OCH}_3 \) and \( \text{C}_2\text{H}_5\text{OH} \).

Formal Charge

**Activity 1**

When constructing Lewis structures there is often more than one possible way to put atoms together to form a molecule. Calculating the formal charge on each atom is an important method used to check the viability of your structure.

1. How is formal charge calculated?

2. Determine the formal charge on all the atoms in \( \text{NH}_3 \).

3. Determine the formal charge on all the atoms in \( \text{CF}_3\text{CH}_2\text{OH} \).

**Activity 2**

1. What are the formal charge characteristics for the most stable structure of a compound?
2. Draw two different structures for hydroxylamine, which has the chemical formula NOH₃. Then determine the formal charge on each atom in the structure and based on that calculation explain which is the better structure.

3. Suggest a viable Lewis structure for CH₃CN. Label the formal charge on each atom and use formal charge arguments to justify your Lewis structure.

**Activity 3**

1. What some examples of common polyatomic ions?

2. Draw the Lewis structure for the hydronium cation. Calculate the formal charge on each atom in the ion. Confirm that the sum of the FC’s on all the atoms adds up to equal the charge on the polyatomic ion.

3. Draw the Lewis structure for hydroxide anion. Calculate the formal charge on each atom in the ion. Confirm that the sum of the formal charge of each atom in the ion equals the charge on the polyatomic ion.
Resonance Structures

Compounds that exhibit resonance have structures that are the averaged combination of multiple drawn structures. One way to recognize if a compound is capable of resonance is when you can draw equivalent Lewis structures with a double bond placed in different locations.

Activity 1

1. An incorrect way to describe resonance might be to say that a double bond is flipping back and forth between different atoms. Why is this an incorrect interpretation of resonance and what is a better description?

2. Benzene is a molecule that exhibits two resonance structures, however, the bond energies of all the carbon-carbon bonds within the benzene ring in are the same. Explain the relative strength of these bonds.

Activity 2

1. Draw the Lewis structure of the polyatomic anion, carbonate, CO$_3^{2-}$. Be sure to show all possible resonance structures. Comment on the relative bond strength of all the bonds in this ion. Discuss the formal charge on each atom in the averaged structure.
2. Draw the Lewis structure, including all resonance structures, for dinitrogen monoxide, N₂O.

3. Draw the Lewis structure of NO₂. What two features of this molecule are interesting?

4. Draw 3 resonant Lewis structures for OCN⁻. Determine which makes the greatest contribution by assigning formal charge.

Lewis Structures & Exceptions to Octet Rule

Activity 1

Lewis structure rules don’t always work. In some cases an atom can be stable with fewer than 8 electrons in the valence shell. This is called an incomplete octet.

1. Which elements have an incomplete octet?

2. Draw the correct Lewis structure for BeF₂.

3. Draw the correct Lewis structure for BH₂Cl.
Activity 2

Lewis structure rules don’t always work. In some cases an atom can be stable with more than 8 electrons in the valence shell. This is called an expanded octet. If you use the \( S = N - A \) rule you can often predict that the octet rule is going to break down and will be expanded when the rule predicts fewer bonds than are necessary for the number of atoms in the molecule to be linked together.

1. Explain why nitrogen cannot have an expanded octet but phosphorus can.

2. Draw the Lewis structure for the XeO\(_4\).

3. Draw the Lewis structure for Cl\(_3\)F.

4. Draw the Lewis structure for SO\(_2\)Cl\(_2\).

5. Draw the Lewis structure for SO\(_4^{2-}\).

Activity 3

Lewis structure rules don’t always work. In some cases the sum of all the valence electrons in the molecule adds up to be an odd number. In this case you are going to have a single electron on the molecule. This is called an unpaired electron.

1. Which is more reactive – an unpaired electron or pair of electrons? What is another name for an unpaired electron?

2. The hydroxyl radical contains one hydrogen atom and one oxygen atom. It has a single unpaired electron on the oxygen. Draw the hydroxyl radical structure.

3. Draw the Lewis structure for the hydroxide anion (again!). Notice how it is different from the hydroxyl radical.