Chapter 8 Chemical bonding I: Basic concepts

1. Determine the number of valence electrons for any atom, and write its Lewis symbol
2. Recognize when the octet rule applies to the arrangement of electrons in the valence shell for an atom
3. Describe the origin of the energy terms that lead to stabilization of ionic lattices
4. Predict on the basis of the periodic table the probable formulas of ionic substances formed between common metals and nonmetals
5. Describe how the size or value of radii of ions relate to those of atoms
6. Explain the concept of an iso-electronic series and the origin of changes in ionic radius with in such a series
7. Describe in general differences in physical properties between substances with ionic bonds and those with covalent bonds
8. Describe the basis of Lewis theory, and predict the valence of common nonmetallic elements from their positions in the periodic table.
9. State the basic assumptions of the Lewis theory.
10. Relate the Lewis symbol for an element to its position in the periodic table.
11. Write Lewis structures for simple ionic compounds.
12. Explain the significance of electro negativity and in a general way relate the electro negativity of an element to its position in the periodic table
13. Write the Lewis structure for molecules and ions containing covalent bonds, using the periodic table.
14. Write the resonance forms for molecules or polyatomic ions that are not adequately described by a single Lewis structure
15. You should be able to write Lewis structure for molecules and ions containing covalent molecules that have ~~an odd number of electrons, a deficiency of electrons or~~ an expanded octet.
16. Use the basic rules of Lewis theory to propose a plausible skeleton structure for a molecule and assign valence electrons to this structure.
17. Compute the formal charge on each atom in a Lewis structure, and use formal charges to determine which of several Lewis structures is the most plausible.
18. Recognize situations when resonance occurs and draw plausible resonance structures
19. State which elements can have expanded octets and be able to draw Lewis structures with expanded octets.
20. Predict the relative polarities of bonds using either the periodic table or electronegative values
21. Relate bond enthalpies to bond strengths and use bond enthalpies to estimate ΔH for reactions
22. Define electro negativity, and use its value to assess the relative metallic/nonmetallic character of an element.
23. Use bond distances to help in writing Lewis structures, and bond energies to compute the enthalpy change of a reaction
24. Describe the relationship between electro negativity difference and the percent ionic character of a bond.

Covalent Bonds	
Chemical bond	Force that holds atoms together in combination
Chemical reaction	Process in which substances interact with one another to form new and different substances. Process in which chemical bonds are broken or formed.
Covalent bond	Two electrons shared between two atoms

Cause of formation of covalent bond	A covalent bond consists of two electrons shared by orbital overlap between two atoms. A certain amount of orbital overlap lowers the total energy of the atoms by balancing their attractive and repulsive interactions and makes them more stable.
Single covalent bond	Bond in which two and only two electrons (one pair) are shared between two atoms
Double covalent	Bond in which four (two pair) are shared between two atoms
Triple bonds	Bond in which six electrons (three pair) are shared between two atoms.
Octet rule	Two versions: 1) A hydrogen atom that forms a covalent bond will share 2 electrons, and every other atom that forms a covalent bond will share enough electrons to give it a total of 8 shared and unshared electrons in its valence shell. 2) An atom that forms a covalent bond will share enough electrons to give it the same number of valence electrons as a noble gas atom. 2 valence electrons for bonded hydrogen atom and 8 valence electrons for any other atom
Exceptions to the octet rule	All non-metal atoms except COFN can be exceptions to the octet rule. These are non-metals below the second period. These non-metals are more likely to satisfy their group number than COFN. That is to say, the atom wants to have its group number of electrons around it either as shared or unshared electrons Example: O= S=O Counting the electrons around the sulfur has 6 electrons around it (4 bonds count for 4 electrons and 2 unshared for 2 more) Most metals are also exceptions to the octet rule.
Molecule	Group of two or more atoms bonded to one another
Formula	Collection of symbols representing a molecule or ionic compound
Non polar covalent bond	Bond in which electrons are shared equally between two atoms with the same electro negativity
Polar covalent bond	Bond in which electrons are shared unequally between two atoms with different electro negativity
$\delta+$ and δ^-	Symbols using lower case Greek delta which show partial positive or partial negative charge on each end of a polar covalent bond
Electro negativity	Ability of an atom to attract shared electrons toward itself in a bond
Variation of electro negativity among elements in the periodic table	In general, electro negativity increase upward in a group and to the right in a period
Order of increasing electro negativity among Cl, F, O, H, N, C	H < C < Cl = N < O < F
Lewis formula	Combination of Lewis symbols that show the bonding in a molecule and represents each valence electron by a dot.
Line formula	Modified version of a Lewis formula in which each pair of electrons in a bond is represented by a line.
Molecular formula	Formula that show only the number of atoms of each element in a molecule. The subscript on the symbol shows the number of atoms (no subscript means 1 atom) for example CO_2
Steps to writing a line formula Lewis structure from a molecular formula or an ion	1. Count the number of valence electrons for each atom in the formula. This is the total valence electrons available for bonding and lone pairs (unshared electrons). If the ion is negative, add the charge as electrons.

	If the ion is positive, subtract the charge as ions. 2. Write the symbols for these elements in the following pattern: in the center put the least electronegative atom (this usually the first element in the formula or the element that has the smallest amount of symbols in the formula) Hydrogen and fluorine are NEVER in the center of a molecule. Around the central atom, write the other atoms as if they were on the circumference of a circle and are equidistant from each other. 3. Connect each atom to the central atom with a single bond. Subtract this number of total electrons from the total valence electrons 4. Put the remaining electrons on the outside atoms such that they satisfy the octet rule. If there are any left over atoms, they go on the central atom, even if doing so results in more than an octet 5. If the central atom doesn't have 8 electrons change an unshared pair on an outer electron to a shared pair with the central atom. Continue this process until either the central atom has 8 electrons or its group number is satisfied. 6. Check your structure with the formula charge.
Formal charge	Valence electrons minus electrons used for bonding and unshared electrons VE — (sticks (bonds) + dots (unshared electrons))
Lewis symbol	Symbol for the atom of a representative element in which valence electrons are shown as dots
Method for finding the Lewis symbol for the atom of a representative element	identify the group of the element using the periodic table. The characteristic Lewis symbol for the group show the number and arrangement of electrons in the Lewis symbol for the element.
Only element whose Lewis structure does not match the characteristic Lewis symbol for its group	Helium. The characteristic Lewis symbol shows eight valence electrons, but the first period has no p sub-level so helium is He: .

Of Interest:

- Lattice energy is always endothermic. The magnitude of the lattice energy depends primarily on the amount of charge on the ionic species.
- Transition metals tend to lose valence-shell s electrons first, then as many electrons as required to reach the charge of the ion.
- Polyatomic ions have a net charge but are internally bound with covalent bonds.
- A triple bond is not three times as strong as a single bond. The second and third bonds in a triple bond are invariable weaker that the first bond. The first bond is stronger than a normal single bond.
- In general the distance between a bonded atoms decreases as the number of shared electron pairs increases.
- Equal electron sharing in a covalent bond and complete electron separation in an ionic bond represent the two extremes cases. Most bonds show some charge separation and some electron sharing. These are called polar bonds.
- In general, if the difference in the electronegativities of two atoms in a bond is less than 0.5, the bond is considered non polar. If the difference is 0.5 <difference < 2.0, it is consider polar. If the difference is 2.0 it is considered ionic.
- Lewis structures are much easier to write and the results are more consistently correct when you follow the method outlined in class and in this handout. Reproducible logic is the key.
- The double-headed arrow (↔) is specifically reserved for indicating that two or more structures are resonance forms. This symbol must not be used in other contexts.

- Bond enthalpy is the energy required to break a bond.
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- Transition metals tend to lose valence-shell s electrons first, then as many electrons as required to reach the charge of the ion.