

3.S: Chemical Bonding and Nomenclature (Summary)

- In **covalent bonding**, electrons are *shared* between atoms. In **ionic bonding**, electrons are transferred from one atom to another. Compounds that are formed using only covalent bonds are termed **molecular compounds**.
- The *outermost* electron level in any atom is referred to as the **valence shell**. The electron configuration of the valence shell of an atom can be shown graphically using a **Lewis diagram** (or *electron-dot structure*). The arrangement of “dots” around the chemical symbol for the element are shown singly up through four electrons, and then paired until eight electrons are present.
- To form ions from individual elements, electrons are added or subtracted from the valence shell in order to completely fill the shell with eight electrons (the **octet rule**). The charge on the ion reflects the electrons added or removed. Representative elements through Group IIIA will lose electrons to form cations, while those in groups IVA – VIIA will gain electrons and form anions.
- A **covalent bond** is constructed in a Lewis diagram by pairing a set of unpaired electrons from two different atoms. For the purposes of the “octet rule”, a pair of shared electrons is counted as two electrons for each atom. Multiple covalent bonds (**double bonds** and **triple bonds**) are used, if necessary, to give each bonded atom a full octet (except, of course, for helium and hydrogen). When two or more atoms are bonded together utilizing covalent bonds, the compound is referred to as a **molecule**.
- As a rule of thumb, **ionic** bonds will be formed whenever the compound contains a **metal**. **Covalent** bonding will be observed in compounds containing only **semimetals** or **nonmetals**.
- Groups of covalently bonded semimetals or nonmetals which are charged are called **polyatomic ions**. Common examples include sulfate dianion, nitrate anion, phosphate trianion, etc. These polyatomic ions are commonly paired with metals forming ionic compounds.
- Many (but not all) polyatomic ions can be drawn in two or more equivalent Lewis representations. These are called **resonance forms** of the ion. The actual electronic structure of the ion is a combination of these Lewis structures and is called the **resonance hybrid**.
- The **electronegativity** of an element is a measure of the tendency of that element to attract electrons towards itself. Electronegativities range from 0.6 to 4.0, with fluorine as the most electronegative element (a value of 4.0). The general trend in the periodic table is for electronegativity to increase from the lower left-hand corner (Fr) to the upper right-hand corner (F).
- Covalent bonds formed between atoms with different electronegativity will be **polarized** with the greatest electron density localized around the most electronegative atom. The effect of electronegativity on electron distribution within a molecule can be shown using a computer-calculated **electrostatic potential map** where colors are used to represent electron density.
- Elements in periods 3 – 7 can accommodate more than eight electrons in their valence shells. This phenomena is called **valence shell expansion** and molecules involving these elements may have 10 – 14 valence electrons in properly drawn Lewis diagrams. Exceptions to the “octet rule” also exist where the valence shell contains less than eight electrons, or contains unpaired electrons.
- When naming simple, binary **ionic** compounds, the cation is named first using the name of the element, followed by the anion, where the suffix *ide* is added to the root name of the element. Multipliers are **not used**. For transition metals in which the metal can assume a variety of oxidation states (different positive charges), the charge of the metal ion is shown in the name using Roman numerals, in parenthesis, following the name of the element (i.e., iron (III) chloride).
- When naming simple binary **molecular** compounds (compounds containing only covalent bonds) the *least electronegative* element is (generally) named first, followed by the second element, where the suffix *ide* is again added to the root name of the element. In molecular compounds multipliers are used to indicate the number of each atom present (mono-, di-, tri-, tetra-, etc.) with the exception that *mono* is not used for the first element in the compound.

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