7.S: Aqueous Solutions (Summary)

- Covalent bonds formed between atoms of differing electronegativity are *polarized*, resulting in a bond that is electron-rich on one end and electron-poor on the other. Covalent bonds that are polarized are said to have a **dipole**, where the term **dipole moment** refers to the direction and magnitude of the charge separation.
- If a molecule is *asymmetric* (such as a molecule with a bend structure) local dipoles along covalent bonds can combine, generating a **molecular dipole**, in which the entire molecule has an imbalance with regard to electron distribution. This can be shown with an dipole arrow (with a positive end) indicating the direction of the charge separation in the molecule.
- If a molecule is symmetrical (such as BH_3, which is trigonal planar), the individual dipoles associated with the covalent bonds cancel, leaving a molecule with no molecular dipole.
- Water has a significant molecular dipole, allowing it to strongly interact with other polar molecules and with individual ions from ionic compounds. Because of this, water is able to break the electrostatic attraction between ions in compounds and to move the ions into solution. In solution, cations will be surrounded by a **solvation shell** where the water molecules are oriented so that the *negative* end of the water molecule interacts with the *cation*. Likewise, the *cationic* end of water will surround and solvate *anions*.
- **Molarity** is simply defined as the *number of moles of a solute* dissolved in *one liter of solvent*, or (**moles/L**). The abbreviation for molarity is the **uppercase M**.
- You should remember that **concentration** multiplied by **volume** gives the number of **moles** of solute; (**moles/L**)×**L=moles.**
- When you are given the amount of solute in grams, remember, **mass** divided by **molar mass** gives **moles**. Dividing this by **volume** (in liters) gives **molarity**;

$$\frac{\left(\frac{grams}{grams/mole}\right)}{L} = molarity$$

- In a **standard solution**, we simply know the *molarity* of the solute(s). Because *concentration* (the molarity) multiplied by *volume* gives us *moles*, we can calculate the number of moles in given volume and use this value in standard stoichiometric calculations.
- A sample of a solution of known volume is called an aliquot. When an aliquot of a solution is diluted into a larger volume, the final concentration can be calculated as:

$$\left(\frac{volume\ of\ the\ aliquot}{final\ volume}\right) = \left(\frac{final\ concentration}{stock\ concentration}\right)$$

  or

$$\left(\frac{V}{V_f}\right) = \left(\frac{C_f}{C_i}\right)$$

  where C_i and C_f are the *stock* and *final* concentrations, respectively, V is the volume of the aliquot and V_f is the final volume of the solution. This relationship is also often stated as $V_1C_1 = V_2C_2$, where the subscripts refer to the initial and final concentrations and volumes.