CHAPTER FOUR

TYPES OF CHEMICAL REACTIONS AND SOLUTION
STOICHIOMETRY

For Review

1. Soluble ionic compounds break apart into their separate ions when in solution. KBr(aq) really means K⁺(aq) + Br⁻(aq). The hydration process for ions has the partial negative end of the polar water molecules surrounding and stabilizing the cations in solution. Here, many water molecules would align themselves so the oxygen end of water aligns with the K⁺ ions. The negative ions are stabilized in water by having the partial positive end of the polar water molecules surround the anions in solution. Here, many water molecules would align themselves so the hydrogen end of water aligns with the Br⁻ ion. All this is assumed when (aq) is placed after an ionic compound.

C₂H₅OH is a covalent compound and does not break up into ions when dissolved in water. C₂H₅OH is a polar covalent compound which means it has a partial negative end and a partial positive end. The hydration process for polar covalent solutes in water is again to have the opposite charged parts of the solute and solvent align themselves. Here, the hydrogens of many water molecules align with the partial negative end of C₂H₅OH and the oxygens of many water molecules align with the partial positive end of C₂H₅OH. This is the hydration process for polar covalent compounds and is always assumed when (aq) is listed after a covalent compound. Note: at this point, you are not able to predict the partial negative and partial positive ends for polar covalent compounds.

2. The electrolyte designation refers to how well the dissolved solute breaks up into ions. Strong electrolytes fully break up into ions when in water, weak electrolytes only partially break up into ions (less than 5% usually), and nonelectrolytes do not break up into ions when they dissolve in water. The conductivity apparatus illustrated in Figure 14.4 is one way to experimentally determine the type of electrolyte. As illustrated, a bright light indicates many charge carriers (ions) are present and the solute is a strong electrolyte. A dim light indicates few ions are present so the solute is a weak electrolyte, and no light indicates no ions are present so the solute is a nonelectrolyte.

3. The electrolyte designation refers to what happens to a substance when it dissolves in water, i.e., does it produce a lot of ions or a few ions or no ions when the substance dissolves. A weak electrolyte is a substance that only partially dissociates in water to produce only a few ions. Solubility refers to how much substance can dissolve in a solvent. "Slightly soluble" refers to substances that dissolve only to a small extent, whether it is an electrolyte or a nonelectrolyte. A weak electrolyte may be very soluble in water, or it may be slightly soluble. Acetic acid is an example of a weak electrolyte that is very soluble in water.
4. Consider a 0.25 \textit{M} solution of NaCl. The two ways to write 0.25 \textit{M} as conversion factors are:

\[
\frac{0.25 \text{ mol NaCl}}{L} \quad \text{or} \quad \frac{1 \text{ L}}{0.25 \text{ mol NaCl}}
\]

Use the first conversion factor when converting from volume of NaCl solution (in liters) to mol NaCl and use the second conversion factor when converting from mol NaCl to volume of NaCl solution.

5. Dilution refers to a method used to prepare solutions. In a dilution, one starts with a certain amount of a more concentrated solution; water is then added to a specific new volume forming a solution which has a lower concentration (it is diluted). The quantity that is constant in a dilution is the moles of solute between the concentrated solution and the dilute solution. The difference between the two solutions is that we have the same number of solute particles occupying a larger volume of water; the new solution is less concentrated. Molarity (mol/L) × volume (L) gives mol of solute. \(M_1 \times V_1 = \text{mol of solute in the concentrated solution.}\) \(M_2 \times V_2 = \text{mol of solute in the diluted solution.}\) Since the mol of solute are constant between the two solutions, \(M_1 \times V_1 = M_2 \times V_2\) for dilution problems.

6. In the first set of beakers, \(\text{Pb}^{2+}\) reacts with \(\text{Br}^-\) to form \(\text{PbBr}_2(s)\) (from the solubility rules). The \(\text{Na}^+\) and \(\text{NO}_3^-\) ions are spectator ions. There are 6 \(\text{Na}^+\), 6 \(\text{Br}^-\), 3 \(\text{Pb}^{2+}\) and 6 \(\text{NO}_3^-\) ions present initially. \(\text{Pb}^{2+}(\text{aq}) + 2 \text{Br}^-(\text{aq}) \rightarrow \text{PbBr}_2(s)\). The 3 \(\text{Pb}^{2+}\) ions will react with 6 \(\text{Br}^-\) ions to form 3 formula units of the \(\text{PbBr}_2\) precipitate. The ions remaining in solution will be 6 \(\text{Na}^+\) ions and 6 \(\text{NO}_3^-\) ions floating about in solution with three formula units of \(\text{PbBr}_2\) settled on the bottom as the precipitate.

In the second set of beakers, \(\text{Al}^{3+}\) reacts with \(\text{OH}^-\) to form \(\text{Al(OH)}_3(s)\) (from the solubility rules). There are 3 \(\text{Al}^{3+}\) ions, 9 \(\text{Cl}^-\) ions, 6 \(\text{OH}^-\) ions and 6 \(\text{K}^+\) ions present initially. \(\text{Al}^{3+}(\text{aq}) + 3 \text{OH}^-(\text{aq}) \rightarrow \text{Al(OH)}_3(s)\). The 6 \(\text{OH}^-\) ions will react with two of the three \(\text{Al}^{3+}\) ions to form 2 formula units of the \(\text{Al(OH)}_3\) precipitate. One \(\text{Al}^{3+}\) ion is in excess. Also remaining in solution are the \(\text{K}^+\) and \(\text{Cl}^-\) spectator ions. Therefore, your drawing should show 1 \(\text{Al}^{3+}\) ion, 9 \(\text{Cl}^-\) ions, and 6 \(\text{K}^+\) ions in solution with 2 \(\text{Al(OH)}_3\) formula units shown as the precipitate.

7. The formula equation keeps all of the ions together in nice, neutral formulas. This is not how soluble ionic compounds are present in solution. Soluble ionic compounds (indicated with \(\text{aq}\)) exist as separate ions in solution; only the precipitate has the ions together. So in the complete ionic equation, the soluble ionic compounds are shown as separate ions and the precipitate is shown as staying together. In the net ionic equation, we get rid of the ions that did nothing but balance the charge. These ions are called spectator ions. In the net ionic equation, only the ions that come together to form the precipitate are shown. In the following balanced equations, the formula equation is written first, the complete ionic equation is second, and the net ionic equation is third.
2 \text{NaBr}(aq) + \text{Pb(NO}_3\text{)}_2(aq) \rightarrow \text{PbBr}_2(s) + 2 \text{NaNO}_3(aq) \\
2 \text{Na}^+(aq) + 2 \text{Br}^-(aq) + \text{Pb}^{2+}(aq) + 2\text{NO}_3^-(aq) \rightarrow \text{PbBr}_2(s) + 2 \text{Na}^+(aq) + 2 \text{NO}_3^-(aq) \\
\text{Pb}^{2+}(aq) + 2 \text{Br}^-(aq) \rightarrow \text{PbBr}_2(s) \\
\text{AlCl}_3(aq) + 3 \text{KOH}(aq) \rightarrow \text{Al(OH)}_3(s) + 3 \text{KCl}(aq) \\
\text{Al}^{3+}(aq) + 3 \text{Cl}^-(aq) + 3 \text{K}^+(aq) + 3 \text{OH}^-(aq) \rightarrow \text{Al(OH)}_3(s) + 3 \text{K}^+(aq) + 3 \text{Cl}^-(aq) \\
\text{Al}^{3+}(aq) + 3 \text{OH}^-(aq) \rightarrow \text{Al(OH)}_3(s) \\

8. An acid-base reaction involves the transfer of a H$^+$ ion from an acid to a base. The H$^+$ ion is just a proton; an electron is removed from neutral hydrogen to form H$^+$. Acid-base reactions are commonly called proton transfer reactions. The acid is the proton donor and the base is the proton acceptor.

The strong bases are (by the solubility rules) LiOH, NaOH, KOH, RbOH, CsOH, Ca(OH)$_2$, Sr(OH)$_2$ and Ba(OH)$_2$. When OH$^-$ from these strong bases react with H$^+$ (a proton), water is formed \([\text{H}^+(aq) + \text{OH}^-(aq) \rightarrow \text{H}_2\text{O(l)})\].

Titration: a technique in which one solution is used to analyze another.

Stoichiometric point: when exactly enough of one solution has been added to react completely with the other solution.

Neutralization: a term used for acid-base reactions referring to the added OH$^-$ reacting with (neutralizing) the protons from the acid. It can be reversed; the added protons neutralizing the OH$^-$ ions from the base. Either way, the neutralization reaction is: \(\text{H}^+(aq) + \text{OH}^-(aq) \rightarrow \text{H}_2\text{O(l)}\).

Standardization: the experimental procedure of running a controlled acid-base reaction in order to determine the concentration of a specific solution.

9. Oxidation: a loss of electrons
Reduction: a gain of electrons.

Oxidizing agent: a reactant that accepts electrons from another reactant.
Reducing agent: a reactant that donates electrons to another reactant

The best way to identify a redox reaction is to assign oxidation states to all elements in the reaction. If elements show a change in oxidation states when going from reactants to products, then the reaction is a redox reaction. No change in oxidation states indicates the reaction is not a redox reaction. Note that the element oxidized shows an increase in oxidation state and the element reduced shows a decrease in oxidation state.
10. Half-reactions: the two parts of an oxidation-reduction reaction, one representing oxidation, the other reduction.

Overall charge must be balanced in any chemical reaction. We balance the charge in the half-reactions by adding electrons to either the reactant side (reduction half-reaction) or the product side (oxidation half-reaction). In the overall balanced equation, the number of electrons lost by the oxidation half-reaction has to exactly equal the number of electrons gained in the reduction half-reaction. Since electrons lost = electrons gained, then electrons will not appear in the overall balanced equation.

See Section 4.10 for a flow chart summarizing the half-reaction method for balancing redox reactions in acidic or basic solution. In all cases, the redox reaction must be mass balanced as well as charge balanced. Mass balance means that we have the same number and types of atoms on both sides of the equation; charge balance means that the overall net charge on each side of the reaction is the same.