Nature of Aqueous Solutions

- Solute – substance being dissolved.
- Solvent – liquid water.
- Electrolyte – substance that when dissolved in water produces a solution that can conduct electricity.

Electrolytes

- Strong Electrolytes – conduct current very efficiently (bulb shines brightly). Completely ionized in water.
- Weak Electrolytes – conduct only a small current (bulb glows dimly). A small degree of ionization in water.
- Nonelectrolytes – no current flows (bulb remains unlit). Dissolves but does not produce any ions.

Chemical Reactions of Solutions

- We must know:
  - The nature of the reaction.
  - The amounts of chemicals present in the solutions.

Molarity

- Molarity (*M*) = moles of solute per volume of solution in liters:

$$M = \text{Molarity} = \frac{\text{moles of solute}}{\text{liters of solution}}$$

$$3\ M\ \text{HCl} = \frac{\text{6 moles of HCl}}{\text{2 liters of solution}}$$

Concentration of Ions

- For a 0.25 *M* $CaCl_2$ solution:

  $CaCl_2 \rightarrow Ca^{2+} + 2Cl^-$

  - Ca^{2+}: $1 \times 0.25\ M = 0.25\ M\ Ca^{2+}$
  - Cl^-: $2 \times 0.25\ M = 0.50\ M\ Cl^-$.

Notice

- The solution with the greatest number of ions is not necessarily the one in which:
  - the volume of the solution is the largest.
  - the formula unit has the greatest number of ions.

Dilution

- The process of adding water to a concentrated or stock solution to achieve the molarity desired for a particular solution.
- Dilution with water does not alter the numbers of moles of solute present.
- Moles of solute before dilution = moles of solute after dilution

$$M_1V_1 = M_2V_2$$

- Precipitation Reactions
- Acid–Base Reactions
- Oxidation–Reduction Reactions

Precipitation Reaction

- A double displacement reaction in which a solid forms and separates from the solution.
  - When ionic compounds dissolve in water, the resulting solution contains the separated ions.
  - Precipitate – the solid that forms.

Precipitates

- *Soluble* – solid dissolves in solution; (*aq*) is used in reaction equation.
- *Insoluble* – solid does not dissolve in solution; (*s*) is used in reaction equation.
- *Insoluble* and *slightly soluble* are often used interchangeably.

Simple Rules for Solubility

1. Most nitrate (NO_3^-) salts are soluble.
2. Most alkali metal (group 1A) salts and NH_4^+ are soluble.
3. Most Cl^-, Br^-, and I^- salts are soluble (except Ag^+, Pb^{2+}, Hg_2^{2+}).
4. Most sulfate salts are soluble (except $BaSO_4$, $PbSO_4$, Hg_2SO_4, $CaSO_4$).
5. Most OH^- are only slightly soluble (NaOH, KOH are soluble, $Ba(OH)_2$, $Ca(OH)_2$ are marginally soluble).
6. Most S^{2-}, CO_3^{2-}, CrO_4^{2-}, PO_4^{3-} salts are only slightly soluble, except for those containing the cations in Rule 2.

Formula Equation (Molecular Equation)

- Gives the overall reaction stoichiometry but not necessarily the actual forms of the reactants and products in solution.
- Reactants and products generally shown as compounds.
- Use solubility rules to determine which compounds are aqueous and which compounds are solids.

$$AgNO_3(aq) + NaCl(aq) \longrightarrow AgCl(s) + NaNO_3(aq)$$

Net Ionic Equation

- Includes only those solution components undergoing a change.
  - Show only components that actually react.

$$Ag^{+}(aq) + Cl^{-}(aq) \longrightarrow AgCl(s)$$

- Spectator ions are not included (ions that do not participate directly in the reaction).
  - Na^{+} and NO_3^{-} are spectator ions.

Solving Stoichiometry Problems for Reactions in Solution

1. Identify the species present in the combined solution, and determine what reaction occurs.
2. Write the balanced net ionic equation for the reaction.
3. Calculate the moles of reactants.
4. Determine which reactant is limiting.
5. Calculate the moles of product(s), as required.
6. Convert to grams or other units, as required.

Let's Think About It

- Where are we going?
  - To find the mass of solid $Pb_3(PO_4)_2$ formed.
- How do we get there?
  - What are the ions present in the combined solution?
  - What is the balanced net ionic equation for the reaction?
  - What are the moles of reactants present in the solution?
  - Which reactant is limiting?
  - What moles of $Pb_3(PO_4)_2$ will be formed?
  - What mass of $Pb_3(PO_4)_2$ will be formed?

CONCEPT CHECK! (Part II)

10.0 mL of a 0.30 *M* sodium phosphate solution reacts with 20.0 mL of a 0.20 *M* lead(II) nitrate solution (assume no volume change).

- What is the concentration of nitrate ions left in solution after the reaction is complete?

0.27 *M*

Let's Think About It

- Where are we going?
  - To find the concentration of nitrate ions left in solution after the reaction is complete.
- How do we get there?
  - What are the moles of nitrate ions present in the combined solution?
  - What is the total volume of the combined solution?

Let's Think About It

- Where are we going?
  - To find the concentration of phosphate ions left in solution after the reaction is complete.
- How do we get there?
  - What are the moles of phosphate ions present in the solution at the start of the reaction?
  - How many moles of phosphate ions were used up in the reaction to make the solid $Pb_3(PO_4)_2$?
  - How many moles of phosphate ions are left over after the reaction is complete?
  - What is the total volume of the combined solution?

Acid–Base Reactions (Brønsted–Lowry)

- Acid—proton donor
- Base—proton acceptor
- For a strong acid and base reaction:

$$H^+(aq) + OH^-(aq) \longrightarrow H_2O(l)$$

Performing Calculations for Acid–Base Reactions

1. List the species present in the combined solution *before any reaction occurs*, and decide what reaction will occur.
2. Write the balanced net ionic equation for this reaction.
3. Calculate moles of reactants.
4. Determine the limiting reactant, where appropriate.
5. Calculate the moles of the required reactant or product.
6. Convert to grams or volume (of solution), as required.

Acid–Base Titrations

- *Titration* – delivery of a measured volume of a solution of known concentration (the titrant) into a solution containing the substance being analyzed (the analyte).
- *Equivalence point* – enough titrant added to react exactly with the analyte.
- *Endpoint* – the indicator changes color so you can tell the equivalence point has been reached.

Let's Think About It

- Where are we going?
  - To find the moles of NaOH required for the reaction.
- How do we get there?
  - What are the ions present in the combined solution? What is the reaction?
  - What is the balanced net ionic equation for the reaction?
  - What are the moles of H^+ present in the solution?
  - How much OH^- is required to react with all of the H^+ present?

Redox Reactions

- Reactions in which one or more electrons are transferred.

Rules for Assigning Oxidation States

1. Oxidation state of an atom in an element = 0
2. Oxidation state of monatomic ion = charge of the ion
3. Oxygen = –2 in covalent compounds (except in peroxides where it = –1)
4. Hydrogen = +1 in covalent compounds
5. Fluorine = –1 in compounds
6. Sum of oxidation states = 0 in compounds
7. Sum of oxidation states = charge of the ion in ions

Redox Characteristics

- Transfer of electrons
- Transfer may occur to form ions
- Oxidation – increase in oxidation state (loss of electrons); reducing agent
- Reduction – decrease in oxidation state (gain of electrons); oxidizing agent

Balancing Oxidation–Reduction Reactions by Oxidation States

1. Write the unbalanced equation.
2. Determine the oxidation states of all atoms in the reactants and products.
3. Show electrons gained and lost using “tie lines.”
4. Use coefficients to equalize the electrons gained and lost.
5. Balance the rest of the equation by inspection.
6. Add appropriate states.

- Balance the reaction between solid zinc and aqueous hydrochloric acid to produce aqueous zinc(II) chloride and hydrogen gas.