Chapter 4. Aqueous Reactions and Solution Stoichiometry

Media Resources

Figures and Tables	**Section**
In Transparency Pack and on Matter:	
Table 4.1 Solubility Guidelines for Common Ionic Compounds in Water	4.2 Precipitation Reactions
Figure 4.7 Electrolyte Scheme	4.3 Acid-Base Reactions
Table 4.5 Activity Series of Metals in Aqueous Solution	4.4 Oxidation-Reduction Reactions
Figure 4.20 Stoichiometry Procedure	4.6 Solution Stoichiometry and Chemical Analysis
On Matter:	
Figure 4.3 Dissolution of Compounds	4.1 General Properties of Aqueous Solutions
Figure 4.6 Formation of NH_4OH	4.3 Acid-Base Reactions
Figure 4.6.01 Space-Filling Models of Acids	4.3 Acid-Base Reactions
Table 4.2 Common Strong Acids and Bases	4.3 Acid-Base Reactions
Table 4.3 Summary of the Electrolytic Behavior of Common Ionic and Molecular Compounds	4.3 Acid-Base Reactions
Table 4.4 Some Common Antacids	4.3 Acid-Base Reactions
Figure 4.22 Titrating an Acid	4.6 Solution Stoichiometry and Chemical Analysis

Movies on Matter	**Section**
Animations:	
Dissolution of NaCl in Water	4.1 General Properties of Aqueous Solutions
Dissolution of $KMnO_4$	4.1 General Properties of Aqueous Solutions
Introduction to Aqueous Acids	4.3 Acid-Base Reactions
Introduction to Aqueous Bases	4.3 Acid-Base Reactions
Dissolution of $Mg(OH)_2$ by Acid	4.3 Acid-Base Reactions
Oxidation-Reduction Reactions – Part I	4.4 Oxidation-Reduction Reactions
Oxidation-Reduction Reactions – Part II	4.4 Oxidation-Reduction Reactions
Solution Formation from a Solid	4.5 Concentrations of Solutions
Solution Formation by Dilution	4.5 Concentrations of Solutions
Video Lab Demonstrations:	
Strong and Weak Electrolytes	4.1 General Properties of Aqueous Solutions
Formation of Silver Crystals	4.4 Oxidation-Reduction Reactions
Natural Indicators	4.6 Solution Stoichiometry and
Chemical	
	Analysis

Further Readings:	**Section**
An Analogy for Solubility: Marbles and Magnets	4.2 Precipitation Reactions
Significance, Concentration Calculations, Weak and Strong Acids	4.3 Acid-Base Reactions
When Is a Strong Electrolyte Strong?	4.3 Acid-Base Reactions
Pictorial Analogies X: Solutions of Electrolytes	4.3 Acid-Base Reactions
Reinforcing Net Ionic Equation Writing	4.3 Acid-Base Reactions
Oxidation and Reduction	4.4 Oxidation-Reduction Reactions
Oxidation Numbers	4.4 Oxidation-Reduction Reactions
Simple Method for Determination of Oxidation Numbers of Atoms in Compounds	4.4 Oxidation-Reduction Reactions
What Makes Gold Such a Noble Metal?	4.4 Oxidation-Reduction Reactions
A Cyclist's Guide to Ionic Concentration	4.5 Concentrations of Solutions
Teaching Dilutions	4.5 Concentrations of Solutions
On the Use of Intravenous Solutions to Teach Some Principles of Solution Chemistry	4.5 Concentrations of Solutions

Literature Demonstrations:	**Section**
Conductivity and Extent of Dissociation of Acids in Aqueous Solution.	4.1 General Properties of Aqueous Solutions
Name That Precipitate	4.2 Precipitation Reactions
Alka Seltzer Poppers: An Interactive Exploration	4.3 Acid-Base Reactions
Demonstrations with Red Cabbage Indicator	4.3 Acid-Base Reactions
A Handheld Reaction: Production of Ammonia Gas	4.3 Acid-Base Reactions
Teas as Natural Indicators Chemical	4.6 Solution Stoichiometry and Analysis

Student CD	**Section**
Activities:	
Ionic Compounds	4.2 Precipitation Reactions
Titration	4.6 Solution Stoichiometry and Chemical Analysis
Movies:	
Electrolytes and Nonelectrolytes Solutions	4.1 General Properties of Aqueous
Introduction to Aqueous Acids	4.3 Acid-Base Reactions
Introduction to Aqueous Bases	4.3 Acid-Base Reactions
Formation of Silver Crystals	4.4 Oxidation-Reduction Reactions
Oxidation-Reduction Reactions I	4.4 Oxidation-Reduction Reactions
Oxidation-Reduction Reactions II	4.4 Oxidation-Reduction Reactions
Solution Formation by a Solid	4.5 Concentration of Solutions
Solution Formation by Dilution	4.5 Concentration of Solutions
Titration	4.6 Solution Stoichiometry and Chemical Analysis

Common Student Misconceptions

- Molarity is moles of solute per *liter of solution* not per liter of solvent.
- Students sometimes use moles instead of molarity in $M_{initial}V_{initial} = M_{final}V_{final}$.
- Students sometimes think that water is a good conductor.
- Students sometimes have a problem with the arbitrary difference between strong and weak electrolytes.
- Students often confuse the symbols —— (equilibrium) and ℑ (resonance).
- Students often do not see that the net ionic equation for the reaction between strong acids and strong bases is always $H^+(aq) + OH^-(aq)$ □ $H_2O(l)$.
- Students try to split polyatomic ions into smaller ions when they write net ionic equations.
- Students do not appreciate the difference between equivalence point and end point.

Lecture Outline

4.1 General Properties of Aqueous Solutions

- A solution is a homogeneous mixture of two or more substances.
- A solution is made when one substance (the **solute**) is dissolved in another (the **solvent**).
- The solute is the substance that is present in smallest amount.
- Solutions in which water is the solvent are called **aqueous solutions**.

Electrolytic Properties[1]

- All aqueous solutions can be classified in terms of whether or not they conduct electricity.
- If a substance forms ions in solution, the substance is an **electrolyte**, and the solution conducts electricity. Example: NaCl.
- If a substance does not form ions in solution, the substance is a **nonelectrolyte**, and the solution does not conduct electricity. Example: sugar.

Ionic Compounds in Water

- When an ionic compound dissolves in water, the ions *dissociate*.[2, 3]
  - This means that in solution the solid no longer exists as a well-ordered arrangement of ions in contact with one another.
  - Instead, each ion is surrounded by water molecules.
  - This tends to stabilize the ions in solution and prevent cations and anions from recombining.
  - The positive ions have the surrounding oxygen atoms of water pointing toward the ion; the negative ions have the surrounding hydrogen atoms of water pointing toward the ion.
  - The transport of ions through the solution causes electric current to flow through the solution.

Molecular Compounds in Water[4]

- When a molecular compound (e.g., CH_3OH) dissolves in water there are no ions formed.

---

[1] “Electrolytes and Nonelectrolytes” Movie from Student CD

[2] “Dissolution of NaCl in Water” from Matter

[3] “Dissolution of $KMnO_4$” from Matter

[4] Figure 4.3 from Matter

- Therefore, there is nothing in the solution to transport electric charge, and the solution does not conduct electricity.
- There are some important exceptions.
  - For example, $NH_3(g)$ reacts with water to form $NH_4^+(aq)$ and $OH^-(aq)$.
  - For example, $HCl(g)$ in water ionizes to form $H^+(aq)$ and $Cl^-(aq)$.

Strong and Weak Electrolytes[5]

- Compounds whose aqueous solutions conduct electricity well are called **strong electrolytes.**[6]
  - These substances exist only as ions in solution.
  - Example: NaCl:

$$NaCl(aq) \square Na^+(aq) + Cl^-(aq)$$

  - The single arrow indicates that the Na^+ and Cl^- ions have no tendency to recombine to form NaCl molecules.
- Compounds whose aqueous solutions conduct electricity poorly are called **weak electrolytes.**
  - These substances exist as a mixture of ions and un-ionized molecules in solution.
  - Example: acetic acid, $HC_2H_3O_2$.

$$HC_2H_3O_2(aq) \text{—} H^+(aq) + C_2H_3O_2^-(aq)$$

  - The double arrow means that the reaction is significant in both directions.
  - It indicates that there is a balance between the forward and reverse reactions.
  - This balance produces a state of **chemical equilibrium.**

4.2 Precipitation Reactions

- Reactions that result in the formation of an insoluble product are known as **precipitation reactions**.
- A **precipitate** is an insoluble solid formed by a reaction in solution.[7]
  - Example: $Pb(NO_3)_2(aq) + 2KI(aq) \square PbI_2(s) + 2KNO_3(aq)$

Solubility Guidelines for Ionic Compounds[8, 9]

- The **solubility** of a substance is the amount of that substance that can be dissolved in a given quantity of solvent.
- A substance with a solubility of less than 0.01 mol/L is regarded as being *insoluble*.
- Experimental observations have led to empirical guidelines for predicting solubility.
- Solubility Guidelines for Common Ionic Compounds in Water:[10]
  - Compounds containing alkali metal ions or ammonium ions are soluble.
  - Compounds containing NO_3^- or $C_2H_3O_2^-$ are soluble.
  - Compounds containing Cl^-, Br^-, or I^- are soluble.
    - Exceptions: compounds of Ag^+, Hg_2^{2+}, and Pb^{2+}.
  - Compounds containing SO_4^{2-} are soluble.
    - Exceptions: compounds of Sr^{2+}, Ba^{2+}, Hg_2^{2+}, and Pb^{2+}.
  - Compounds containing S^{2-} are insoluble.
    - Exceptions: compounds of NH_4^+, the alkali metal cations, and Ca^{2+}, Sr^{2+}, and Ba^{2+}.

---

[5] "Strong and Weak Electrolytes" from Matter

[6] "Conductivity and Extent of Dissociation of Acids in Aqueous Solution" from Literature Demonstrations

[7] "Name That Precipitate" from Literature Demonstrations

[8] "Ionic Compounds" Activity from Student CD

[9] "An Analogy for Solubility: Marbles and Magnets" from Further Readings

[10] Table 4.1 from Matter and Transparency Pack

- Compounds of CO_3^{2-} or PO_4^{3-} are insoluble.
    - Exceptions: compounds of NH_4^+ and the alkali metal cations.
- Compounds of OH^- are insoluble.
    - Exceptions: compounds of the alkali metal cations, and Ca^{2+}, Sr^{2+}, and Ba^{2+}.

Exchange (Metathesis) Reactions

- **Exchange reactions** or **metathesis reactions** involve the swapping of ions in solution:

$$\mathbf{A}\mathit{X} + \mathit{B}\mathbf{Y} \rightarrow \mathbf{AY} + \mathit{BX}$$

- Many precipitation and acid-base reactions exhibit this pattern.

Ionic Equations

- Consider $2KI(aq) + Pb(NO_3)_2(aq) \rightarrow PbI_2(s) + 2KNO_3(aq)$.
- Both $KI(aq) + Pb(NO_3)_2(aq)$ are colorless solutions. When mixed, they form a bright yellow precipitate of PbI_2 and a solution of KNO_3.
- The final product of the reaction contains solid PbI_2, aqueous K^+, and aqueous NO_3^- ions.
- Sometimes we want to highlight the reaction between ions.
- **The molecular equation** lists all species in their molecular forms:

$$Pb(NO_3)_2(aq) + 2KI(aq) \rightarrow PbI_2(s) + 2KNO_3(aq)$$

- **The complete ionic equation** lists all strong soluble electrolytes in the reaction as ions:

$$Pb^{2+}(aq) + 2NO_3^-(aq) + 2K^+(aq) + 2I^-(aq) \rightarrow PbI_2(s) + 2K^+(aq) + 2NO_3^-(aq)$$

- **The net ionic equation** lists only those ions that are not common on both sides of the reaction:

$$Pb^{2+}(aq) + 2I^-(aq) \rightarrow PbI_2(s)$$

- Note that **spectator ions**, ions that are present in the solution but play no direct role in the reaction, are omitted in the net ionic equation.

4.3 Acid-Base Reactions

Acids

- **Acids** are substances that are able to ionize in aqueous solution to form H^+.[11]
    - Ionization occurs when a neutral substance forms ions in solution.
      Example: $HC_2H_3O_2$ (acetic acid).
- Since H^+ is a naked proton, we refer to acids as proton donors and bases as proton acceptors.
- Common acids are HCl, HNO_3, vinegar, and vitamin C.[12]
- Acids that ionize to form *one* H^+ ion are called *monoprotic* acids.
- Acids that ionize to form *two* H^+ ions are called *diprotic acids*.

Bases

- **Bases** are substances that react with the H^+ ions formed by acids.[13]
- Hydroxide ions, OH^-, react with the H^+ ions to form water:

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

- Common bases are NH_3 (ammonia), Drano, milk of magnesia.[14]

---

[11] "Introduction to Aqueous Acids" from Matter and Student CD

[12] Figure 4.601 from Matter

[13] "Introduction to Aqueous Bases" from Matter and Student CD

[14] Table 4.4 from Matter

- Compounds that do not contain OH^- ions can also be bases.
  - Proton transfer between NH_3 (a weak base) and water (a weak acid) is an example of an acid-base reaction.[15]
  - Since there is a mixture of NH_3, H_2O, NH_4^+, and OH^- in solution, we write

$$NH_3(aq) + H_2O(l) \text{ — } NH_4^+(aq) + OH^-(aq)$$

Strong and Weak Acids and Bases[16]

- **Strong acids** and **strong bases** are strong electrolytes.
  - They are completely ionized in solution.
  - Strong bases include Group 1A metal hydroxides, $Ca(OH)_2$, $Ba(OH)_2$, and $Sr(OH)_2$.
  - Strong acids include HCl, HBr, HI, $HClO_3$, $HClO_4$, H_2SO_4, and HNO_3.
  - We write the ionization of HCl as:

$$HCl \text{ □ } H^+ + Cl^-$$

- **Weak acids** and **weak bases** are weak electrolytes.
  - Therefore, they are partially ionized in solution.
  - HF(aq) is a weak acid; most acids are weak acids.
  - We write the ionization of HF as:

$$HF \text{ — } H^+ + F^-$$

Identifying Strong and Weak Electrolytes[17, 18, 19, 20, 21]

- Compounds can be classified as strong electrolytes, weak electrolytes, or nonelectrolytes from their solubility.
- Strong electrolytes:
  - If a compound is water soluble and ionic, then it is probably a strong electrolyte.
  - If a compound is water soluble and not ionic, and is a strong acid, then it is a strong electrolyte.
  - Similarly, if a compound is water soluble and not ionic, but is a strong base, then it is a strong electrolyte.
- Weak electrolytes:
  - If a compound is water soluble and not ionic, and is a weak acid or weak base, then it is a weak electrolyte.
- Nonelectrolytes:
  - Otherwise, the compound is probably a nonelectrolyte.

Neutralization Reactions and Salts

- A **neutralization reaction** occurs when an acid and a base react:[22, 23]
  - $HCl(aq) + NaOH(aq)$ □ $H_2O(l) + NaCl(aq)$
  - (acid) + (base) □ (water) + (salt)
- In general, an acid and base react to form a **salt**.

---

[15] Figure 4.6 from Matter

[16] Table 4.2 from Matter

[17] Table 4.3 from Matter

[18] Figure 4.7 from Matter and Transparency Pack

[19] "Significance, Concentration Calculations, Weak and Strong Acids," from Further Readings

[20] "Pictorial Analogies X: Solutions of Electrolytes" from Further Readings

[21] "When Is a Strong Electrolyte Strong?" from Further Readings

[22] "Demonstrations with Red Cabbage Indicator" from Literature Demonstrations

[23] "Reinforcing Net Ionic Equation Writing" from Further Readings

- A salt is any ionic compound whose cation comes from a base, and anion from an acid.
- The other product, H_2O, is a common weak electrolyte.
- Typical examples of neutralization reactions:
  - Reaction between an acid and a metal hydroxide:[24]
    - $Mg(OH)_2$ (milk of magnesia) is a suspension.
    - As HCl is added, the magnesium hydroxide dissolves, and a clear solution containing Mg^{2+} and Cl^- ions is formed.
    - Molecular equation:

$$Mg(OH)_2(s) + 2HCl(aq) \rightarrow MgCl_2(aq) + 2H_2O(l)$$

    - Net ionic equation:

$$Mg(OH)_2(s) + 2H^+(aq) \rightarrow Mg^{2+}(aq) + 2H_2O(l)$$

      - Note that the magnesium hydroxide is an insoluble solid; it appears in the net ionic equation.

Acid-Base Reactions with Gas Formation[25]

- There are many bases besides OH^- that react with H^+ to form molecular compounds.
  - Carbonates and hydrogen carbonates (or bicarbonates) will form $CO_2(g)$ when treated with an acid.[26]
    - Sodium hydrogen carbonate ($NaHCO_3$; baking soda) reacts with HCl to form bubbles of $CO_2(g)$:
    - Molecular equation:

$$NaHCO_3(s) + HCl(aq) \rightarrow NaCl(aq) + H_2CO_3(aq) \rightarrow H_2O(l) + CO_2(g) + NaCl(aq)$$

    - Net ionic equation:

$$H^+(aq) + HCO_3^-(aq) \rightarrow H_2O(l) + CO_2(g)$$

- Reaction of sulfides with acid gives rise to H_2S, the substance that gives rotten eggs their foul odor.
  - Sodium sulfide (Na_2S) reacts with HCl to form $H_2S(g)$:
    - Molecular equation:

$$Na_2S(aq) + 2HCl(aq) \rightarrow H_2S(g) + 2NaCl(aq)$$

    - Net ionic equation:

$$2H^+(aq) + S^{2-}(aq) \rightarrow H_2S(g)$$

4.4 Oxidation-Reduction Reactions

Oxidation and Reduction[27, 28, 29]

- When a substance loses electrons, it undergoes **oxidation**:[30]

$$Ca(s) + 2H^+(aq) \rightarrow Ca^{2+}(aq) + H_2(g)$$

- The neutral Ca has lost two electrons to $2H^+$ to become Ca^{2+}.
- We say Ca has been oxidized to Ca^{2+}.
- When a substance gains electrons, it undergoes **reduction**:

$$2Ca(s) + O_2(g) \rightarrow 2CaO(s)$$

---

[24] "Dissolution of $Mg(OH)_2$ by Acid" from Matter

[25] "A Handheld Reaction: Production of Ammonia Gas" from Literature Demonstrations

[26] "Alka Seltzer Poppers: An Interactive Exploration" Literature Demonstrations

[27] "Oxidation-Reduction Reactions — Part I" from Matter and Student CD

[28] "Formation of Silver Crystals" Movie from Student CD

[29] "Oxidation-Reduction Reactions II" Movie from Student CD

[30] "Oxidation and Reduction" from Further Readings

- In this reaction the neutral O_2 has gained electrons from the Ca to become O^{2-} in CaO.
- We say O_2 has been reduced to O^{2-}.
- In all reduction-oxidation (redox) reactions, one species is reduced at the same time as another is oxidized.

Oxidation Numbers[31, 32]

- Electrons are not explicitly shown in chemical equations.
- **Oxidation numbers** (or *oxidation states*) help up keep track of electrons during chemical reactions.
- Oxidation numbers are assigned to atoms using specific rules.
  - For an atom in its *elemental form*, the oxidation number is always zero.
  - For any *monatomic ion*, the oxidation number equals the charge on the ion.
  - *Nonmetals* usually have negative oxidation numbers.
    - The oxidation number of *oxygen* is usually -2.
      - The major exception is in peroxides (containing the O_2^{2-} ion).
    - The oxidation number of *hydrogen* is +1 when bonded to nonmetals and –1 when bonded to metals.
    - The oxidation number of *fluorine* is -1 in all compounds. The other *halogens* have an oxidation number of -1 in most binary compounds.
  - *The sum of the oxidation numbers* of all atoms in a neutral compound is zero.
  - The sum of the oxidation numbers in a polyatomic ion equals the charge of the ion.
- The oxidation of an element is evidenced by its increase in oxidation number; reduction is accompanied by a decrease in oxidation number.

Oxidation of Metals by Acids and Salts[33]

- The reaction of a metal with either an acid or a metal salt is called a **displacement reaction.**
- General pattern:

$$\mathrm{A + B\mathbf{X}} \ \square\ \mathrm{A\mathbf{X} + B}$$

  - Example: It is common for metals to produce hydrogen gas when they react with acids. Consider the reaction between Mg and HCl:

$$Mg(s) + 2HCl(aq)\ \square\ MgCl_2(aq) + H_2(g)$$

    - In the process the metal is oxidized and the H^+ is reduced.
  - Example: It is possible for metals to be oxidized in the presence of a salt:

$$Fe(s) + Ni(NO_3)_2(aq)\ \square\ Fe(NO_3)_2(aq) + Ni(s)$$

    - The net ionic equation shows the redox chemistry well:

$$Fe(s) + Ni^{2+}(aq)\ \square\ Fe^{2+}(aq) + Ni(s)$$

    - In this reaction iron has been oxidized to Fe^{2+}, while the Ni^{2+} has been reduced to Ni.
- Always keep in mind that whenever one substance is oxidized, some other substance *must* be reduced.

The Activity Series[34, 35]

---

[31] "Oxidation Numbers" from Further Readings

[32] "Simple Method for Determination of Oxidation Numbers of Atoms in Compounds" from Further Readings

[33] "Oxidation-Reduction Reactions — Part II" from Matter

[34] Table 4.5 from Matter and Transparency Pack

- We can list metals in order of decreasing ease of oxidation.
    - This list is an **activity series**.
- The metals at the top of the activity series are called active metals.
- The metals at the bottom of the activity series are called noble metals.
- A metal in the activity series can be oxidized only by a metal ion below it.
- If we place Cu into a solution of Ag^+ ions, then Cu^{2+} ions can be formed because Cu is above Ag in the activity series:[36]

$$Cu(s) + 2AgNO_3(aq) \rightarrow Cu(NO_3)_2(aq) + 2Ag(s)$$

or

$$Cu(s) + 2Ag^+(aq) \rightarrow Cu^{2+}(aq) + 2Ag(s)$$

4.5 Concentrations of Solutions

- The term **concentration** is used to designate the amount of solute dissolved in a given quantity of solvent or solution.

Molarity

- Solutions can be prepared with different concentrations by adding different amounts of solute to solvent.
- The amount (moles) of solute per liter of solution is the **molarity** of the solution:

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

- By knowing the molarity of a quantity of liters of solution, we can easily calculate the number of moles (and, by using molar mass, the mass) of solute.[37]
- Consider weighed copper sulfate, $CuSO_4$ (39.9 g, 0.250 mol) placed in a 250 mL volumetric flask. A little water is added and the flask swirled to ensure that the copper sulfate dissolves. When all the copper sulfate has dissolved, the flask is filled to the mark with water. The molarity of the solution is 0.250 mol $CuSO_4$/0.250 L solution = 1.00 *M*.

Expressing the Concentration of an Electrolyte

- When an ionic compound dissolves, the relative concentrations of the ions in the solution depend on the chemical formula of the compound.[38]
- Example: For a 1.0 *M* solution of NaCl:
    - The solution is 1.0 *M* in Na^+ ions and 1.0 *M* in Cl^- ions.
- Example: For a 1.0 *M* solution of Na_2SO_4:
    - The solution is 2.0 *M* in Na^+ ions and 1.0 *M* in SO_4^{2-} ions.

Interconverting Molarity, Moles, and Volume

- The definition of molarity contains three quantities: molarity, moles of solute, and liters of solution.
    - If we know any two of these, we can calculate the third.
    - Dimensional analysis is very helpful in these calculations.

Dilution[39, 40, 41]

---

[35] "What Makes Gold Such a Noble Metal?" from Further Readings

[36] "Formation of Silver Crystals" from Matter

[37] "Solution Formation from a Solid" from Matter and Student CD

[38] "A Cyclist's Guide to Ionic Concentration" from Further Readings

[39] "Solution Formation by Dilution" from Matter and Student CD

[40] "Teaching Dilutions" from Further Readings

- A solution in concentrated form (*stock solution*) is mixed with solvent to obtain a solution of lower solute concentration.
  - This process is called **dilution**.
- An alternative way of making a solution is to take a solution of known molarity and dilute it with more solvent.
- Since the number of moles of solute remains the same in the concentrated and dilute forms of the solution, we can show:

$$M_{initial}V_{initial} = M_{final}V_{final}$$

- An alternative form of this equation is:

$$M_{concentrated}V_{concentrated} = M_{dilute}V_{dilute}$$

4.6 Solution Stoichiometry and Chemical Analysis

- In approaching stoichiometry problems:[42]
  - Recognize that there are two different types of units:
    - Laboratory units (the macroscopic units that we measure in lab) and
    - Chemical units (the microscopic units that relate to moles).
  - Always convert the laboratory units into chemical units first.
    - Convert grams to moles using molar mass.
    - Convert volume or molarity into moles using M = mol/L.
  - Use the stoichiometric coefficients to move between reactants and product.
    - ***This step requires the balanced chemical equation***
  - Convert the chemical units back into the required units.
    - Convert moles to grams using molar mass.
    - Convert moles to molarity or volume using M = mol/L.

Titrations

- A common way to determine the concentration of a solution is via **titration**.[43, 44, 45]
- We determine the concentration of one substance by allowing it to undergo a specific chemical reaction, of known stoichiometry, with another substance whose concentration is known (**standard solution**).
- Example: Suppose we know the molarity of an NaOH solution and we want to find the molarity of an HCl solution.
  - First, let us examine what we know:
    - Molarity of NaOH, volume of HCl.
  - What do we want?
    - Molarity of HCl.
  - What do we do?
    - Take a known volume of the HCl solution (e.g., 20.0 mL) and measure the number of milliliters of NaOH solution required to react completely with the HCl solution.
    - The point at which stoichiometrically equivalent quantities of NaOH and HCl are brought together is known as the **equivalence point** of the titration.
    - In a titration we often use an acid-base **indicator** to allow us to determine when the equivalence point of the titration has been reached.[46, 47]

---

[41] "On the Use of Intravenous Solutions to Teach Some Principles of Solution Chemistry" from Further Readings

[42] Figure 4.20 from Matter and Transparency Pack

[43] Figure 4.22 from Matter

[44] "Titration" Movie from Student CD

[45] "Titration" Activity from Student CD

    - Acid-base indicators change color at the *end point* of the titration.
    - The indicator is chosen so that the end point corresponds to the equivalence point of the titration.
- What do we get?
    - Volume of NaOH. Since we already have the molarity of the NaOH, we can calculate moles of NaOH.
- Next step?
    - We also know HCl + NaOH □ NaCl + H_2O.
    - Therefore, we know moles of HCl.
- Can we finish?
    - Knowing mol (HCl) and volume of HCl (20.0 mL above), we can calculate the molarity.

---

[46] "Natural Indicators" from Matter

[47] "Teas as Natural Indicators" from Literature Demonstrations

Further Readings:

1. Richard A. Kjonaas, "An Analogy for Solubility: Marbles and Magnets," *J. Chem. Educ.*, Vol. 61, **1984**, 765.

2. H. van Lubeck, "Significance, Concentration Calculations, Weak and Strong Acids," *J. Chem. Educ.*, Vol. 60, **1983**, 189.

3. Albert Kowalak, "When Is a Strong Electrolyte Strong?" *J. Chem. Educ.*, Vol. 65, **1988**, 607.

4. John J. Fortman, "Pictorial Analogies X: Solutions of Electrolytes," *J. Chem. Educ.*, Vol. 71, **1994**, 27–28.

5. Betty J. Wruck, "Reinforcing Net Ionic Equation Writing," *J. Chem. Educ.*, Vol. 73, **1996**, 149–150.

6. Marten J. ten Hoor and Aletta Jacobsscholengemeenschap, "Oxidation and Reduction," *J. Chem. Educ.*, Vol. 60, **1983**, 132. An analogy for remembering the difference between oxidation and reduction.

7. Gian Calzaferri, "Oxidation Numbers," *J. Chem. Educ.*, Vol. 76, **1999**, 362–263.

8. Joel M. Kauffman, "Simple Method for Determination of Oxidation Numbers of Atoms in Compounds," *J. Chem. Educ.*, Vol. 63, **1986**, 474–475.

9. R. Lipkin, "What Makes Gold Such a Noble Metal?" *Science News*, **July 22, 1995**, 62.

10. Arthur M. Last, "A Cyclist's Guide to Ionic Concentration," *J. Chem. Educ.*, Vol. 75, **1998**, 1433.

11. Lloyd J. McElroy, "Teaching Dilutions," *J. Chem. Educ.*, Vol. 73, **1996**, 765–766.

12. Irwin L. Shaprio, "On the Use of Intravenous Solutions to Teach Some Principles of Solution Chemistry," *J. Chem. Educ.*, Vol. 59, **1982**, 725.

Literature Demonstrations:

1. Bassam Z. Shakhashiri, "Conductivity and Extent of Dissociation of Acids in Aqueous Solution" in *Chemical Demonstrations: A Handbook for Teachers of Chemistry*, Vol.3, (Madison: The University of Wisconsin Press, 1989), pp. 140–145. The universal indicator and a conductivity probe are used to explore the relative acidities and conductivities of a series of aqueous acids.

2. A. M. Sarquis and L. M. Woodward, "Alka Seltzer Poppers: An Interactive Exploration," *J. Chem. Educ.*, Vol. 76, **1999**, 386–386. An interactive exercise involving the addition of water to Alka Seltzer®. This demonstration may be used to introduce a variety of concepts such as acid-base chemistry, kinetics, and solubility.

3. L. R. Summerlin, Christie L. Borgford, and Julie B. Ealy, "Name That Precipitate" in *Chemical Demonstrations: A Sourcebook for Teachers*, Vol. 2 (Washington, DC: American Chemical Society, 1987), pp. 121–123. Students explore a variety of ionic reactions that result in the formation of colored precipitates.

4. John J. Fortman and Katherine M. Stubbs, "Demonstrations with Red Cabbage Indicator," *J. Chem. Educ.*, Vol. 69, **1992**, 66–67. The acidic or basic nature of solutions of gases is investigated.

5. L. R. Summerlin, Christie L. Borgford, and Julie B. Ealy, "A Hand-Held Reaction: Production of Ammonia Gas" in *Chemical Demonstrations: A Sourcebook for Teachers,* Vol. 2 (Washington, DC: American Chemical Society, 1987), p. 38. An example of a reaction involving two solids ($NH_4Cl(s)$ and $Ca(OH)_2(s)$) is demonstrated.

6. Dianne N. Epp, "Teas as Natural Indicators," *J. Chem. Educ.*, Vol. 70, **1993**, 326. Infusions from a series of herbal teas provide a source of natural pH indicators in this simple demonstration.