

Floating Pennies

A Cool Drink! An Introduction to Concentrations

Colorful Acid-Base Indicators

Rainbow Colors with Mixed Acid-Base Indicators

Acid-Base Indicators Extracted from Plants

Teas as Natural Indicators

4.4 Oxidation-Reduction Reactions

4.5 Concentrations of Solutions

4.6 Solution Stoichiometry and Chemical Analysis

4.6 Solution Stoichiometry and Chemical Analysis

4.6 Solution Stoichiometry and Chemical Analysis

4.6 Solution Stoichiometry and Chemical Analysis

Chapter 4. Reactions in Aqueous Solution

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 $M_{\text{initial}}V_{\text{initial}} = M_{\text{final}}V_{\text{final}}$.
- Students often disregard rules for significant figures when calculating or using molarities.
- 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 think that nonelectrolytes produce no ions in aqueous solution at all.
- Students sometimes cannot tell the difference between dissolution and dissociation.
- The symbols \rightleftharpoons (equilibrium) and \leftrightarrow (resonance) are often confused.
- Students often do not see that the net ionic equation for the reaction between strong acids and strong bases is always $\text{H}^+(\text{aq}) + \text{OH}^-(\text{aq}) \rightleftharpoons \text{H}_2\text{O}(\text{l})$.
- Students try to split polyatomic ions into smaller ions when they write net ionic equations.
- Students often think that a compound consisting of nonmetals only must be molecular (counter-example: $(\text{NH}_4)_2\text{SO}_4$ which is ionic!)
- Students do not realize that *insoluble* really means *poorly soluble*.
- Students do not appreciate the difference between equivalence point and end point.
- Students usually think that an oxidation necessarily involves a reaction with oxygen and/or addition of an atom of oxygen to the formula.
- Students often think that all atoms of the same element must have the same oxidation number and that this number is uniquely related to the atom's location in the periodic table.
- The equivalence point of a titration is the point where the stoichiometrically correct number of moles of each reactant is present. The end point of a titration is the point where the indicator changes. They are not the same although we choose an indicator that will change as close to the equivalence point as possible. As a consequence, students often equate the terms *equivalence point* and *end point*.

Teaching Tips

- Weaknesses in recollection of ionic nomenclature and the structure of common ions often make it difficult for students to write molecular, complete ionic, and net ionic equations for metathesis reactions. A brief review of ionic nomenclature is often useful prior to covering metathesis reactions.
- Writing the net ionic equation makes it easier for students to focus on the ions participating in a chemical reaction. It is important for students to remember that, although they do not appear in the net ionic equation, the spectator ions are still present in the solution.
- Students may be familiar with two good mnemonics for redox reactions: (1) LEO the lion says GER: *L*ose *e*lectrons *o*xidation, *g*ain *e*lectrons *r*eduction and (2) OIL RIG: *O*xidation *i*nvolves *l*oss of electrons, *r*eduction *i*nvolves *g*ain of electrons.

Lecture Outline

4.1 General Properties of Aqueous Solutions¹

- 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 the smallest amount.
- Solutions in which water is the solvent are called **aqueous solutions**.

Electrolytic Properties²

- All aqueous solutions can be classified in terms of whether or not they conduct electricity.
- If a substance forms ions in solution, then the substance is an **electrolyte** and the solution conducts electricity. An example is NaCl.
- If a substance does not form ions in solution, then the substance is a **nonelectrolyte** and the solution does not conduct electricity. Examples are sucrose and water.

Ionic Compounds in Water^{3,4}

- When an ionic compound dissolves in water, the ions are said to *dissociate*.
 - 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 several water molecules; it is called an *aqueous ion* denoted using the abbreviation “*aq*”.
 - These ions are said to be **solvated**.
 - This tends to stabilize the ions in solution and prevent cations and anions from recombining.
 - The positive ions have the oxygen atoms of water pointing towards the ion; negative ions have the hydrogen atoms of water pointing towards the ion.
 - The transport of ions through the solution causes electric current to flow through the solution.

Molecular Compounds in Water^{5,6}

- When a molecular compound (e.g. CH₃OH) dissolves in water, the solution usually consists of intact molecules dispersed homogeneously in the solution.
- Therefore, there is nothing in the solution to transport electric charge and the solution does not conduct electricity and most molecular compounds are nonelectrolytes.
- There are some important exceptions.
 - For example, HCl(g) in water *ionizes* to form H⁺(*aq*) and Cl⁻(*aq*).

Strong and Weak Electrolytes^{7,8,9,10,11}

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

¹ “Water” 3-D Model from Instructor’s Resource CD/DVD

² “Electrolytes and Nonelectrolytes” Animation from Instructor’s Resource CD/DVD

³ “Dissolution of NaCl in Water” Animation from Instructor’s Resource CD/DVD

⁴ “Sodium Chloride (1 × 1 Unit Cell)” 3-D Model from Instructor’s Resource CD/DVD

⁵ “Hydrogen Chloride” 3-D Model from Instructor’s Resource CD/DVD

⁶ “Ethanol” 3-D Model from Instructor’s Resource CD/DVD

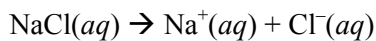
⁷ “Strong Electrolytes” Movie from Instructor’s Resource CD/DVD

⁸ “Conductivity and Extent of Dissociation of Acids in Aqueous Solution” from Live Demonstrations

⁹ “Acetic Acid” 3-D Model from Instructor’s Resource CD/DVD

¹⁰ “Formic Acid” 3-D Model from Instructor’s Resource CD/DVD

¹¹ “Nitric Acid” 3-D Model from Instructor’s Resource CD/DVD



- The single arrow indicates that the Na^+ and Cl^- ions have no tendency to recombine to form NaCl .
- In general, soluble ionic compounds are strong electrolytes.
- Other strong electrolytes include strong acids and soluble strong bases.
- 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.
 - The predominant form of the solute is the un-ionized molecule.
 - Example: acetic acid, $\text{HC}_2\text{H}_3\text{O}_2$

$$\text{HC}_2\text{H}_3\text{O}_2(aq) \rightleftharpoons \text{H}^+(aq) + \text{C}_2\text{H}_3\text{O}_2^-(aq)$$
 - The half-arrows in both directions 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**.

FORWARD REFERENCES:

- Double arrows (\rightleftharpoons) will be used in the chapter on chemical equilibria (Chapter 15) and beyond.
- Strong and weak electrolytes will come up in chapters on acid-base and solubility equilibria (Chapters 16 and 17) as well as in electrochemistry (Chapter 20).
- Equilibria involving *insoluble* or *poorly soluble* compounds and their ions will be discussed in more detail in Chapter 17 (section 4).
- Dissolving of substances in solvents and properties of solutions will be discussed in Chapter 13.
- Interactions between ions and molecules of a solvent (ion-dipole interactions) will be further discussed in Chapters 11 and 13.

4.2 Precipitation Reactions^{12,13,14}

- 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.
 - Example: $\text{Pb}(\text{NO}_3)_2(aq) + 2\text{KI}(aq) \rightarrow \text{PbI}_2(s) + 2\text{KNO}_3(aq)$

Solubility Guidelines for Ionic Compounds^{15,16,17,18,19}

- The **solubility** of a substance at a particular temperature is the amount of that substance that can be dissolved in a given quantity of solvent at that temperature.
- 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:
 - Compounds containing alkali metal ions or ammonium ions are soluble.
 - Compounds containing NO_3^- or $\text{C}_2\text{H}_3\text{O}_2^-$ are soluble.
 - Compounds containing Cl^- , Br^- or I^- are soluble.
 - Exceptions are the compounds of Ag^+ , Hg_2^{2+} , and Pb^{2+} .
 - Compounds containing SO_4^{2-} are soluble.

¹² “Precipitation Reactions” Movie from Instructor’s Resource CD/DVD

¹³ Figure 4.4 from Transparency Pack

¹⁴ “Name That Precipitate” from Live Demonstrations

¹⁵ Table 4.1 from Transparency Pack

¹⁶ “Ionic Compounds” Activity from Instructor’s Resource CD/DVD

¹⁷ “Solubility of Some Silver Compounds” from Live Demonstrations

¹⁸ “Solubility Rules: Three Suggestions for Improved Understanding” from Further Readings

¹⁹ “An Analogy for Solubility: Marbles and Magnets” from Further Readings

- Exceptions are the compounds of Sr^{2+} , Ba^{2+} , Hg_2^{2+} , and Pb^{2+} .
- Compounds containing S^{2-} are insoluble.
 - Exceptions are the compounds of NH_4^+ , the alkali metal cations, and Ca^{2+} , Sr^{2+} , and Ba^{2+} .
- Compounds of CO_3^{2-} or PO_4^{3-} are insoluble.
 - Exceptions are the compounds of NH_4^+ and the alkali metal cations.
- Compounds of OH^- are insoluble.
 - Exceptions are the compounds of NH_4^+ , the alkali metal cations, and Ca^{2+} , Sr^{2+} , and Ba^{2+} .

Exchange (Metathesis) Reactions

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

$$\text{AX} + \text{BY} \rightarrow \text{AY} + \text{BX}.$$
- Many precipitation and acid-base reactions exhibit this pattern.

Ionic Equations^{20,21}

- Consider $2\text{KI}(aq) + \text{Pb}(\text{NO}_3)_2(aq) \rightarrow \text{PbI}_2(s) + 2\text{KNO}_3(aq)$.
- Both $\text{KI}(aq)$ and $\text{Pb}(\text{NO}_3)_2(aq)$ 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 complete chemical forms:

$$\text{Pb}(\text{NO}_3)_2(aq) + 2\text{KI}(aq) \rightarrow \text{PbI}_2(s) + 2\text{KNO}_3(aq)$$
- The **complete ionic equation** lists all strong soluble electrolytes in the reaction as ions:

$$\text{Pb}^{2+}(aq) + 2\text{NO}_3^-(aq) + 2\text{K}^+(aq) + 2\text{I}^-(aq) \rightarrow \text{PbI}_2(s) + 2\text{K}^+(aq) + 2\text{NO}_3^-(aq)$$
 - Only strong electrolytes dissolved in aqueous solution are written in ionic form.
 - Weak electrolytes and nonelectrolytes are written in their complete chemical form.
- The **net ionic equation** lists only those ions which are not common on both sides of the reaction:

$$\text{Pb}^{2+}(aq) + 2\text{I}^-(aq) \rightarrow \text{PbI}_2(s)$$
 - 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.
- A suitable procedure for writing net ionic equations is the following:
 - Write a balanced molecular equation for the reaction.
 - Rewrite the equation to show the ions that forming solution when each soluble strong electrolyte dissociates into its component ions.
 - Only strong electrolytes dissolved in aqueous solution are written in ionic form
 - Identify and cancel **spectator ions**.

FORWARD REFERENCES:

- Net ionic equations will be frequently used in chapters dealing with acid-base reactions (Chapters 16 and 17) as well as in electrochemistry (Chapter 20, Appendix E)
- Equilibria involving *insoluble* or *poorly soluble* compounds and their ions will be discussed in more detail in Chapter 17 (section 17.4).

4.3 Acids, Bases, and Neutralization Reactions

Acids

- **Acids** are substances that are able to ionize in aqueous solution to form H^+ .
 - Ionization occurs when a neutral substance forms ions in solution.
An example is $\text{HC}_2\text{H}_3\text{O}_2$ (acetic acid).
- Since H^+ is a naked proton, we refer to acids as *proton donors* and bases as *proton acceptors*.

²⁰ “Writing a Net Ionic Equation” Activity from Instructor’s Resource CD/DVD

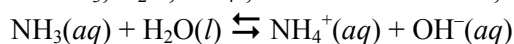
²¹ “Reinforcing Net Ionic Equation Writing” from Further Readings

- Acids that ionize to form *one* H^+ ion are called *monoprotic acids*.
 - Common monoprotic acids include HCl , HNO_3 and $\text{HC}_2\text{H}_3\text{O}_2$.
- Acids that ionize to form *two* H^+ ions are called *diprotic acids*.
 - A common diprotic acid is H_2SO_4 .

Bases^{22,23,24,25}

- **Bases** are substances that accept or react with the H^+ ions.
- Hydroxide ions, OH^- , react with the H^+ ions to form water:

$$\text{H}^+(aq) + \text{OH}^-(aq) \rightarrow \text{H}_2\text{O}(l)$$
- Common bases are NaOH , KOH , and $\text{Ca}(\text{OH})_2$.
- Compounds that do not contain OH^- ions can also be bases.
 - Proton transfer to NH_3 (a weak base) from water (a weak acid) is an example of an acid–base reaction.
 - Since there is a mixture of NH_3 , H_2O , NH_4^+ , and OH^- in solution, we write

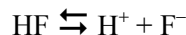


Strong and Weak Acids and Bases^{26,27,28,29,30}

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



- **Weak acids** and **weak bases** are weak electrolytes.
 - They are partially ionized in aqueous solution.
 - $\text{HF}(aq)$ is a weak acid; most acids are weak acids.
 - We write the ionization of HF as:



Identifying Strong and Weak Electrolytes^{31,32}

- Compounds can be classified as strong electrolytes, weak electrolytes, or nonelectrolytes by looking at their solubility.
- Strong electrolytes:
 - Soluble ionic compounds are strong electrolytes.
 - Molecular compounds that are strong acids are strong electrolytes.
- Weak electrolytes:
 - Weak acids and bases are weak electrolytes.
- Nonelectrolytes:

²² “The Origin of the Term Base” from Further Readings

²³ “Hydroxide Ion” 3-D Model from Instructor’s Resource CD/DVD

²⁴ “Ammonia” 3-D Model from Instructor’s Resource CD/DVD

²⁵ “Ammonium Ion” 3-D Model from Instructor’s Resource CD/DVD

²⁶ “Significance, Concentration Calculations, Weak and Strong Acids” from Further Readings

²⁷ “Introduction to Aqueous Acids” Animation from Instructor’s Resource CD/DVD

²⁸ “Introduction to Aqueous Bases” Animation from Instructor’s Resource CD/DVD

²⁹ “Factors that Influence Relative Acid Strength in Water: A Simple Model” from Further Readings

³⁰ “When Is a Strong Electrolyte Strong?” from Further Readings

³¹ “Pictorial Analogies X: Solutions of Electrolytes” from Further Readings

³² “Food Is Usually Acidic, Cleaners Are Usually Basic” from Live Demonstrations

- All other compounds, including water.

Neutralization Reactions and Salts^{33,34,35,36,37}

- A **neutralization reaction** occurs when an acid and a base react:
 - $\text{HCl}(aq) + \text{NaOH}(aq) \rightarrow \text{H}_2\text{O}(l) + \text{NaCl}(aq)$
 - (acid) + (base) (water) + (salt)
- In general, an acid and a base react to form a **salt**.
- 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 non-electrolyte.
- A typical example of a neutralization reaction is the reaction between an acid and a metal hydroxide:
 - $\text{Mg}(\text{OH})_2$ (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:

$$\text{Mg}(\text{OH})_2(s) + 2\text{HCl}(aq) \rightarrow \text{MgCl}_2(aq) + 2\text{H}_2\text{O}(l)$$
 - Net ionic equation:

$$\text{Mg}(\text{OH})_2(s) + 2\text{H}^+(aq) \rightarrow \text{Mg}^{2+}(aq) + 2\text{H}_2\text{O}(l)$$
 - Note that the magnesium hydroxide is an insoluble solid; it appears in the net ionic equation.
 - Note that the ions exchange partners thus neutralization reactions between acids and metal hydroxides are metathesis reactions

Neutralization Reactions with Gas Formation^{38,39,40,41}

- There are many bases besides OH^- that react with H^+ to form molecular compounds.
 - Reaction of sulfides with acid gives rise to $\text{H}_2\text{S}(g)$.
 - Sodium sulfide (Na_2S) reacts with HCl to form $\text{H}_2\text{S}(g)$:
 - Molecular equation:

$$\text{Na}_2\text{S}(aq) + 2\text{HCl}(aq) \rightarrow \text{H}_2\text{S}(g) + 2\text{NaCl}(aq)$$
 - Net ionic equation:

$$2\text{H}^+(aq) + \text{S}^{2-}(aq) \rightarrow \text{H}_2\text{S}(g)$$
 - Carbonates and hydrogen carbonates (or bicarbonates) will form $\text{CO}_2(g)$ when treated with an acid.
 - Sodium bicarbonate (NaHCO_3 ; baking soda) reacts with HCl to form bubbles of $\text{CO}_2(g)$:
 - Molecular equation:

$$\text{NaHCO}_3(s) + \text{HCl}(aq) \rightarrow \text{NaCl}(aq) + \text{H}_2\text{CO}_3(aq) \rightarrow \text{H}_2\text{O}(l) + \text{CO}_2(g) + \text{NaCl}(aq)$$
 - Net ionic equation:

$$\text{H}^+(aq) + \text{HCO}_3^-(aq) \rightarrow \text{H}_2\text{O}(l) + \text{CO}_2(g)$$

FORWARD REFERENCES:

- Strong acids and bases will be revisited in Chapter 16.
- Strong acids and bases will be used as titrants in acid-base titrations (Chapter 17)
- Equilibria involving weak acids and bases will be further discussed in Chapters 16 and 17.
- Environmental impact of weak acid equilibria will be discussed on Chapter 18.

³³ “A Hand-Held Reaction: Production of Ammonia Gas” from Live Demonstrations

³⁴ “Fizzing and Foaming: Reactions of Acids with Carbonates” from Live Demonstrations

³⁵ “Demonstrations with Red Cabbage Indicator” from Live Demonstrations

³⁶ “Dissolution of $\text{Mg}(\text{OH})_2$ by Acid” Animation from Instructor’s Resource CD/DVD

³⁷ “Alka Seltzer Poppers: An Interactive Exploration” from Live Demonstrations

³⁸ “Determination of Neutralizing Capacity of Antacids” from Live Demonstrations

³⁹ “Milk of Magnesia versus Acid” from Live Demonstrations

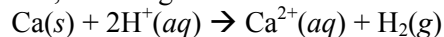
⁴⁰ “Carbon Dioxide” 3-D Model from Instructor’s Resource CD/DVD

⁴¹ “Hydrogen Carbonate Ion” 3-D Model from Instructor’s Resource CD/DVD

4.4 Oxidation-Reduction Reactions

Oxidation and Reduction^{42,43,44,45}

- **Oxidation-reduction**, or *redox*, reactions involve the transfer of electrons between reactants.
- When a substance loses electrons, it undergoes **oxidation**:



- 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**:

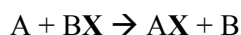
$$2\text{Ca}(s) + \text{O}_2(g) \rightarrow 2\text{CaO}(s)$$
 - In this reaction the neutral O_2 has gained electrons from the Ca to become O^{2-} in CaO.
 - We say O_2 has been reduced to O^{2-} .
- In all redox reactions, one species is reduced at the same time as another is oxidized.

Oxidation Numbers^{46,47,48,49,50}

- Electrons are not explicitly shown in chemical equations.
- **Oxidation numbers** (or *oxidation states*) help us 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; positive for metals and negative for nonmetals.
 - 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 an increase in its oxidation number; reduction is accompanied by a decrease in an oxidation number.

Oxidation of Metals by Acids and Salts^{51,52,53,54}

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



⁴² “Reduction of CuO” Movie from Instructor’s Resource CD/DVD

⁴³ “Oxidation-Reduction Reactions: Part I” Animation from Instructor’s Resource CD/DVD

⁴⁴ “Oxidation-Reduction Reactions: Part II” Animation from Instructor’s Resource CD/DVD

⁴⁵ “Oxidation and Reduction” from Further Readings

⁴⁶ “Oxidation Numbers I” Activity from Instructor’s Resource CD/DVD

⁴⁷ “Oxidation Numbers II” Activity from Instructor’s Resource CD/DVD

⁴⁸ “Oxidation Numbers” from Further Readings

⁴⁹ “Oxidation States of Manganese: Mn^{7+} , Mn^{6+} , Mn^{4+} , and Mn^{2+} ” from Live Demonstrations

⁵⁰ “Simple Method for Determination of Oxidation Numbers of Atoms in Compounds” from Further Readings

⁵¹ “Oxidation-Reduction Chemistry of Tin and Zinc” Movie from Instructor’s Resource CD/DVD

⁵² “Producing Hydrogen Gas from Calcium Metal” from Live Demonstrations

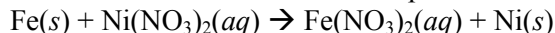
⁵³ “Making Hydrogen Gas an Acid and a Base” from Live Demonstrations

⁵⁴ “Precipitation, Redox, and Neutralization Reactions” Activity from Instructor’s Resource CD/DVD

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



- 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:

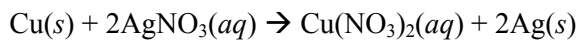


- The net ionic equation shows the redox chemistry well:

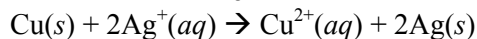
$$\text{Fe}(s) + \text{Ni}^{2+}(aq) \rightarrow \text{Fe}^{2+}(aq) + \text{Ni}(s)$$
- In this reaction iron has been oxidized to Fe^{2+} , while the Ni^{2+} has been reduced to Ni.
- Always keep in mind that whenever one substance is oxidized, some other substance *must* be reduced.

The Activity Series^{55,56,57,58,59}

- 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 only be oxidized 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:



or



FORWARD REFERENCES:

- Oxidation numbers will be frequently used in electrochemistry (Chapter 20, Appendix E).
- Balancing of redox reactions will be covered in Chapter 20.

4.5 Concentrations of Solutions^{60,61}

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

Molarity^{62,63}

- 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** or molar concentration (symbol *M*) 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.

⁵⁵ “Activity Series for Some Metals” from Live Demonstrations

⁵⁶ “An Activity Series: Zinc, Copper, and Silver Half-Cells” from Live Demonstrations

⁵⁷ “Floating Pennies” from Live Demonstrations

⁵⁸ Table 4.5 from Transparency Pack

⁵⁹ “Formation of Silver Crystals” Movie from Instructor’s Resource CD/DVD

⁶⁰ “What Makes Gold Such a Noble Metal?” from Further Readings

⁶¹ “A Cool Drink! An Introduction to Concentrations” from Live Demonstrations

⁶² “Solution Formation from a Solid” Animation from Instructor’s Resource CD/DVD

⁶³ “Dissolution of KMnO_4 ” Animation from Instructor’s Resource CD/DVD

- 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 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 $0.250 \text{ mol CuSO}_4 / 0.250 \text{ L solution} = 1.00 \text{ M}$.

Expressing the Concentration of an Electrolyte⁶⁴

- When an ionic compound dissolves, the relative concentrations of the ions in the solution depend on the chemical formula of the compound.
 - 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 can be helpful in these calculations.

Dilution^{65,66,67}

- 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 alternate 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 diluted forms of the solution, we can show:

$$M_{\text{conc}}V_{\text{conc}} = M_{\text{dil}}V_{\text{dil}}$$

- An alternate form of this equation is:

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

FORWARD REFERENCES:

- Molarity will be used throughout the course as the most common form of concentration.
- The concept of molarity is not limited to solutions; one can calculate molarity for gases and use them in K_c expressions in Chapter 15 and beyond.
- Molarity can be converted into other concentrations (molality, normality, ppm, etc.) as shown in Chapter 13; molarity will be used to calculate osmotic pressure (section 13.5).
- In some later chapters (14 and beyond) molarity will be also symbolized as [solute].
- Dilutions will come up in select acid-base equilibrium problems in Chapter 16 and in titrations (Chapter 17).

4.6 Solution Stoichiometry and Chemical Analysis⁶⁸

- In approaching stoichiometry problems:
 - 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).

⁶⁴ “A Cyclist’s Guide to Ionic Concentration” from Further Readings

⁶⁵ “Solution Formation by Dilution” Animation from Instructor’s Resource CD/DVD

⁶⁶ “Teaching Dilutions” from Further Readings

⁶⁷ “On the Use of Intravenous Solutions to Teach Some Principles of Solution Chemistry” from Further Readings

⁶⁸ Figure 4.18 from Transparency Pack

- Always convert the laboratory units into chemical units first.
 - Convert grams to moles using molar mass.
 - Convert volume or molarity into moles using $M = \text{mol/L}$.
- Use the stoichiometric coefficients to move between reactants and products.
 - ***This step requires the balanced chemical equation.***
- Convert the laboratory units back into the required units.
 - Convert moles to grams using molar mass.
 - Convert moles to molarity or volume using $M = \text{mol/L}$.

Titration^{69,70,71,72,73,74,75}

- A common way to determine the concentration of a solution is via **titration**.
- We determine the concentration of one substance by allowing it to undergo a specific chemical reaction, of known stoichiometry, with a substance with known concentration (**standard solution**).
- *Monoprotic acids and bases react with each other in a stoichiometric ratio of 1:1.*
- Example: Suppose we know the molarity of an NaOH solution and we want to find the molarity of an HCl solution.
 - What do 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 (i.e., 20.00 mL) and measure the number of mL of 0.100 M 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.
 - The equivalence point is a theoretical concept that can be calculated “on paper” only.
 - 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.
 - 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; the end point is determined experimentally.
 - What do we get?
 - We get the volume of NaOH. Since we already have the molarity of the NaOH, we can calculate moles of NaOH.
 - What is the next step?
 - We also know $\text{HCl} + \text{NaOH} \rightarrow \text{NaCl} + \text{H}_2\text{O}$ (note the 1:1 stoichiometric ratio between HCl and NaOH).
 - Therefore, we know moles of HCl.
 - Can we finish?
 - Knowing mol (HCl) and volume of HCl, we can calculate the molarity.

FORWARD REFERENCES:

- Acid-base titrations will be discussed in detail in Chapter 17.

⁶⁹ “Acid-Base Titration” Animation from Instructor’s Resource CD/DVD

⁷⁰ “Colorful Acid-Base Indicators” from Live Demonstrations

⁷¹ “Acid-Base Titration” Activity from Instructor’s Resource CD/DVD

⁷² “Rainbow Colors with Mixed Acid-Base Indicators” from Live Demonstrations

⁷³ “Acid-Base Indicators Extracted from Plants” from Live Demonstrations

⁷⁴ “Acid-Base Indicators: A New Look at an Old Topic” from Further Readings

⁷⁵ “Teas as Natural Indicators” from Live Demonstrations

Further Readings:

1. Bob Blake, "Solubility Rules: Three Suggestions for Improved Understanding," *J. Chem. Educ.*, Vol. 80, **2003**, 1348–1349.
2. Richard A. Kjonaas, "An Analogy for Solubility: Marbles and Magnets," *J. Chem. Educ.*, Vol. 61, **1984**, 765.
3. Betty J. Wruck, "Reinforcing Net Ionic Equation Writing," *J. Chem. Educ.*, Vol. 73, **1996**, 149–150.
4. William B. Jensen, "The Origin of the Term "Base","" *J. Chem. Educ.*, Vol. 83, **2006**, 1130.
5. H. van Lubeck, "Significance, Concentration Calculations, Weak and Strong Acids," *J. Chem. Educ.*, Vol. 60, **1983**, 189.
6. Reading Michael J. Moran, "Factors that Influence Relative Acid Strength in Water: A Simple Model," *J. Chem. Educ.*, Vol. 83, **2006**, 800-803.
7. Albert Kowalak, "When Is a Strong Electrolyte Strong?" *J. Chem. Educ.*, Vol. 65, **1988**, 607.
8. John J. Fortman, "Pictorial Analogies X: Solutions of Electrolytes," *J. Chem. Educ.*, Vol. 71, **1994**, 27–28.
9. Gian Calzaferri, "Oxidation Numbers," *J. Chem. Educ.*, Vol. 76, **1999**, 362–363.
10. R. Lipkin, "What Makes Gold Such a Noble Metal?" *Science News*, July 22, **1995**, 62.
11. Arthur M. Last, "A Cyclist's Guide to Ionic Concentration," *J. Chem. Educ.*, Vol. 75, **1998**, 1433.
12. Lloyd J. McElroy, "Teaching Dilutions," *J. Chem. Educ.*, Vol. 73, **1996**, 765–766.
13. Irwin L. Shapiro, "On the Use of Intravenous Solutions to Teach Some Principles of Solution Chemistry," *J. Chem. Educ.*, Vol. 59, **1982**, 725.
14. Ara S. Kooser, Judith L. Jenkins, and Lawrence E. Welch, "Acid-Base Indicators: A New Look at an Old Topic," *J. Chem. Educ.*, Vol. 78, **2001**, 1504–1506.
15. Marten J. ten Hoor and Aletta Jacobs Scholengemeenschap, "Oxidation and Reduction," *J. Chem. Educ.*, Vol. 60, **1983**, 132. An analogy for remembering oxidation and reduction.
16. Joel M. Kauffman, "Simple Method for Determination of Oxidation Numbers of Atoms in Compounds," *J. Chem. Educ.*, Vol. 63, **1986**, 474–475.

Live Demonstrations:

1. Bassam Z. Shakhashiri, "Conductivity and Extent of Dissociation of Acids in Aqueous Solution," *Chemical Demonstrations: A Handbook for Teachers of Chemistry, Volume 3* (Madison: The University of Wisconsin Press, 1989), pp. 140–145. Universal indicator and a conductivity probe are used to explore the relative acidity and conductivity 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. Lee R. Summerlin, Christie L. Borgford, and Julie B. Ealy, "Name That Precipitate," *Chemical Demonstrations, A Sourcebook for Teachers, Volume 2* (Washington: American Chemical Society, 1988), pp. 121–123. Students explore a variety of ionic reactions that result in the formation of colored precipitates.
4. Lee. R. Summerlin, Christie L. Borgford, and Julie B. Ealy, "Solubility of Some Silver Compounds," *Chemical Demonstrations, A Sourcebook for Teachers, Volume 2* (Washington: American Chemical Society, 1988), pp. 83–85. The solubility of a series of silver salts and complexes is explored in this colorful demonstration.
5. Bassam Z. Shakhashiri, "Food is Usually Acidic, Cleaners Are Usually Basic," *Chemical Demonstrations: A Handbook for Teachers of Chemistry, Volume 3* (Madison: The University of Wisconsin Press, 1989), pp. 65–69. The pH of a variety of household chemicals is determined using indicators and pH meters.
6. Lee R. Summerlin, Christie L. Borgford, and Julie B. Ealy, "A Hand-Held Reaction: Production of Ammonia Gas," *Chemical Demonstrations, A Sourcebook for Teachers, Volume 2* Washington: American Chemical Society, 1988), p. 38. An example of a reaction involving two solids (NH_4Cl and $\text{Ca}(\text{OH})_2$) is demonstrated.
7. Bassam Z. Shakhashiri, "Fizzing and Foaming: Reactions of Acids with Carbonates," *Chemical Demonstrations: A Handbook for Teachers of Chemistry, Volume 3* (Madison: The University of Wisconsin Press, 1989), pp. 96–99.
8. 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.
9. Bassam Z. Shakhashiri, "Determination of Neutralizing Capacity of Antacids," *Chemical Demonstrations: A Handbook for Teachers of Chemistry, Volume 3* (Madison: The University of Wisconsin Press, 1989), pp. 162–166.
10. Lee. R. Summerlin, Christie L. Borgford, and Julie B. Ealy, "Milk of Magnesia versus Acid," *Chemical Demonstrations, A Sourcebook for Teachers, Volume 2* (Washington: American Chemical Society, 1988), p. 173. An antacid, milk of magnesia, is mixed with acid in this demonstration.
11. Lee. R. Summerlin, and James. L. Ealy, Jr., "Oxidation States of Manganese: Mn^{7+} , Mn^{6+} , Mn^{4+} , and Mn^{2+} ," *Chemical Demonstrations, A Sourcebook for Teachers, Volume 1* (Washington: American Chemical Society, 1988), p.133–134 .

12. Lee. R. Summerlin, Christie L. Borgford, and Julie B. Ealy, “Producing Hydrogen Gas from Calcium Metal,” *Chemical Demonstrations, A Sourcebook for Teachers, Volume 2* (Washington: American Chemical Society, **1988**), pp. 51–52.
13. Lee. R. Summerlin, and James. L. Ealy, Jr., “Activity Series for Some Metals,” *Chemical Demonstrations, A Sourcebook for Teachers, Volume 1* (Washington: American Chemical Society, **1988**), p. 150. An overhead projector demonstration employing hydrogen gas formation.
14. Lee. R. Summerlin, Christie L. Borgford, and Julie B. Ealy, “Making Hydrogen Gas from an Acid and a Base,” *Chemical Demonstrations, A Sourcebook for Teachers, Volume 2* (Washington: American Chemical Society, **1988**), pp. 33–34. Hydrogen gas is collected as a product of the reaction of aluminum with either HCl or NaOH.
15. Bassam Z. Shakhshiri, “An Activity Series: Zinc, Copper, and Silver Half Cells,” *Chemical Demonstrations: A Handbook for Teachers of Chemistry, Volume 4* (Madison: The University of Wisconsin Press, **1992**), pp. 101–106.
16. Lee. R. Summerlin, Christie L. Borgford, and Julie B. Ealy, “Floating Pennies,” *Chemical Demonstrations, A Sourcebook for Teachers, Volume 2* (Washington: American Chemical Society, **1988**), p. 63. The zinc core of copper-coated pennies reacts with acid to form pennies that float in this demonstration.
17. Mindy Bedrossian, “A Cool Drink! An Introduction to Concentrations,” *J. Chem. Educ.*, Vol. 85, **2005**, 240A.
18. Bassam Z. Shakhshiri, “Colorful Acid-Base Indicators,” *Chemical Demonstrations: A Handbook for Teachers of Chemistry, Volume 3* (Madison: The University of Wisconsin Press, **1989**), pp. 33–40.
19. Bassam Z. Shakhshiri, “Rainbow Colors with Mixed Acid-Base Indicators,” *Chemical Demonstrations: A Handbook for Teachers of Chemistry, Volume 3* (Madison: The University of Wisconsin Press, **1989**), pp. 41–46.
20. Bassam Z. Shakhshiri, “Acid-Base Indicators Extracted from Plants,” *Chemical Demonstrations: A Handbook for Teachers of Chemistry, Volume 3* (Madison: The University of Wisconsin Press, **1989**), pp. 50–57.
21. 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.