

0.1 Outline

- Aqueous solutions (electrolytes, concentrations)
- Precipitation reactions (soluble and insoluble compounds)
- Molecular, ionic, and net ionic equations
- Oxidation-Reduction reactions

0.2 Terminology

In **solutions**, we need to define the following terms:

- **solvent** – The medium (e.g., water, ethanol, benzene, etc.) in which a solute is dissolved to form a solution.
- **solute** – The substance (e.g. NaCl, glucose, etc.) dissolved in a solvent to form a solution.

0.3 Concentration of Solute

The amount of solute in a solution is given by its **concentration**.

$$\text{Molarity(M)} = \frac{\text{moles of solute}}{\text{liters of solution}} \quad (1)$$

$$\begin{aligned} [\text{NaCl}] &= 0.1M \\ &= \frac{0.1 \text{ moles of NaCl}}{1 \text{ L of solution}} \end{aligned}$$

0.4 Preparing Solutions

- Weigh out a solid solute and dissolve in a given quantity of solvent.
- **Dilute** a concentrated solution to give one that is less concentrated.

0.5 Using Molarity

What mass of oxalic acid, $\text{H}_2\text{C}_2\text{O}_4$, is required to make 250.00 mL of a 0.0500 M solution?

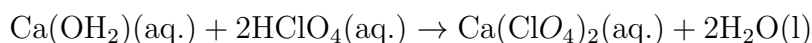
$$\begin{aligned} \text{molar mass} &= (2 * 1.008) + (2 * 12.011) + (4 * 15.999) \\ &= 2.016 + 24.022 + 63.996 \\ &= 90.034 \text{ g mol}^{-1} \\ 250.00 \text{ mL} &\times \frac{1 \text{ L}}{1,000 \text{ mL}} \times \frac{0.05 \text{ mol}}{1\text{L}} \times \frac{90.034 \text{ g H}_2\text{C}_2\text{O}_4}{1 \text{ mol H}_2\text{C}_2\text{O}_4} = 1.125425 \text{ g H}_2\text{C}_2\text{O}_4 \end{aligned}$$

0.6 Preparing a Solution by Dilution

You have 50.0 mL of 3.0 M NaOH and you want 0.50 M NaOH. What does one do?

$$\begin{aligned}
 M_1 V_1 &= M_2 V_2 \\
 (3.0)(50.0) &= (0.5)V_2 \\
 V_2 &= \frac{(3.0)(50.0)}{0.5} \\
 V_2 &= (6.0)(50.0) \\
 V_2 &= 300.0 \text{ mL} \\
 &= 3.0 \times 10^2 \text{ mL}
 \end{aligned}$$

In an acid-base titration, it takes 38.55 mL of 0.650 M perchloric acid (HClO_4) to completely neutralize 25.00 mol calcium hydroxide ($\text{Ca}(\text{OH}_2)$) solution.



A) How many moles of HClO_4 are needed for the complete neutralization?

$$38.55\text{mL perchloric acid} \times \frac{1 \text{ L perchloric acid}}{1,000 \text{ mL perchloric acid}} \times \frac{0.650 \text{ mol}}{1 \text{ L}} = 0.0251 \text{ mol perchloric acid}$$

B) How many moles of $\text{Ca}(\text{OH}_2)_2$ got consumed during the neutralization?

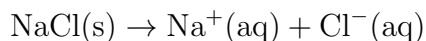
$$0.0251 \text{ mol perchloric acid} \times \frac{1 \text{ mol Ca}(\text{OH}_2)_2}{2 \text{ mol perchloric acid}} = 0.01255 \text{ Ca}(\text{OH}_2)_2$$

C) What is the concentration of $\text{Ca}(\text{OH}_2)_2$ in the original solution before titration?

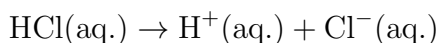
$$\frac{25.00 \text{ mol Ca}(\text{OH}_2)_2}{38.55\text{mL}} \times \frac{1,000 \text{ mL}}{1 \text{ mL}} =$$

0.7 Dissociation

- When an ionic compound dissolves in water, the solvent pulls the individual ions from the crystal and solvates them.
- This process is called **dissociation**.



- An **electrolyte** is a substance that dissociates into ions when dissolved in water.
- Ionic compounds dissociate in water (e.g., NaCl , BaCl_2 , etc.)
- Only a few molecular compounds are capable of dissociating in water.
- For example,



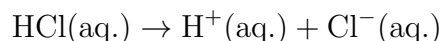
0.8 Electrolytes

- An **electrolyte** is a substance that dissociates into ions when dissolved in water.
- A **nonelectrolyte** may dissolve in water, but it does not dissociate into ions when it does so.
- There are many examples of molecular compounds (e.g., methanol (CH_3OH), table sugar (sucrose; $\text{C}_{12}\text{H}_{22}\text{O}_{11}$)) that serve as nonelectrolytes in water.

Table 1: Summary

	Strong Electrolyte	Weak Electrolyte	Nonelectrolyte
Ionic	All	None	None
Molecular	Strong acids (see Table 4.2)	Weak acids Weak bases	All other compounds

e.g., the molecular compound HCl is a strong acid that dissociates completely in water. As a result, HCl is a strong electrolyte.



e.g., the molecular compound acetic acid, CH_3COOH , is a weak acid that is only partially dissociated. As a result, acetic acid is a weak electrolyte.



- Ion concentration can be measured using conductivity. Ions carry electrical charge from one electrode to the other, completing the electrical circuit.

No ions A nonelectrolyte solution does not contain ions, and the bulb does not light.

Few ions If the solution contains a small number of ions, the bulb will only dimly lit.

Many ions If the solution contains a large number of ions, the bulb will be brightly lit.

0.9 Strong Electrolytes Are...

- Strong acids

HCl Hydrochloric

HBr Hydrobromic

HI Hydroiodic

HClO_3 Chloric

HClO₄ Perchloric**HNO₃** Nitric**H₂SO₄** Sulfuric

- Strong bases
 - Group 1A metal hydroxides
 - * LiOH
 - * NaOH
 - * KOH
 - * RbOH
 - * CsOH
 - Heavy group 2A metal hydroxides
 - * Ca(OH)₂
 - * Sr(OH)₂
 - * Ba(OH)₂
- Also soluble ionic salts

Table 2: Solubility Guidelines for Common Ionic Compounds in Water

Soluble Ionic Compounds		Important Exceptions
Compounds containing	NO ₃ ⁻	None
	CH ₃ COO ⁻	None
	Cl ⁻	Compounds of Ag ⁺ , Hg ₂ ²⁺ , and Pb ²⁺
	Br ⁻	Compounds of Ag ⁺ , Hg ₂ ²⁺ , and Pb ²⁺
	I ⁻	Compounds of Ag ⁺ , Hg ₂ ²⁺ , and Pb ²⁺
	SO ₄ ²⁻	Compounds of Sr ²⁺ , Ba ²⁺ , Hg ₂ ²⁺ , and Pb ²⁺
Insoluble Ionic Compounds		Important Exceptions
Compounds containing	S ²⁻	Compounds of NH ₄ ⁺ , the alkali metal cations, Ca ²⁺ , Sr ²⁺
	CO ₃ ²⁻	Compounds of NH ₄ ⁺ and the alkali metal cations
	PO ₄ ³⁻	Compounds of NH ₄ ⁺ and the alkali metal cations
	OH ⁻	Compounds of NH ₄ ⁺ , the alkali metal cations, Ca ²⁺ , Sr ²⁺

- Note that all common ionic compounds of the alkali metal ions (group 1A of the periodic table) and the ammonium ion (NH₄⁺) are soluble in water.

0.10 Precipitation Reactions

When ions that are insoluble (as could be predicted by the solubility guidelines in Table 2) are mixed and form compounds, a **precipitate** is formed.

The insoluble yellow solid formed on the right is lead iodide (PbI_2).

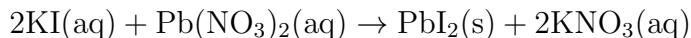
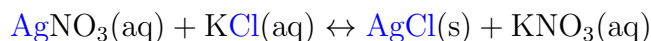


Figure 1: Lead Iodide

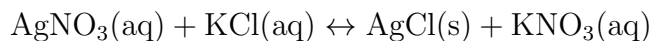
0.11 Metathesis (Exchange) Reactions

- Metathesis comes from a Greek word that means “to transpose.”
- It appears that the ions in the reactants exchange, or transpose, ions.
- This is like switching dance partners!



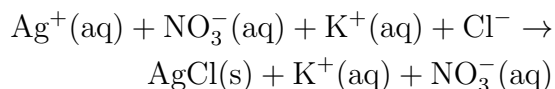
0.12 Molecular Equation

The molecular equation lists the reactants and products in their molecular forms.



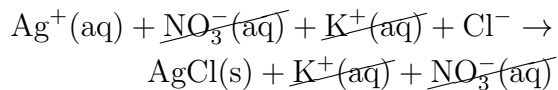
0.13 Ionic Equation

- In the ionic equation, all strong electrolytes (strong acids, strong bases, and soluble ionic salts) are dissociated into their ions.
- This more accurately reflects the species that are found in the reaction mixture.

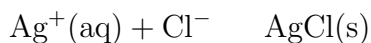


0.14 Net Ionic Equation

- To form the net ionic equation, cross out anything that does not change from the left side of the equation to the right.



- The only things left in the equation are those things that change (i.e., react) during the course of the reaction.



0.15 Oxidation-Reduction Reactions

- An **oxidation** occurs when an atom or ion *loses* electrons.
- A **reduction** occurs when an atom or ion *gains* electrons.
- One cannot occur without the other.

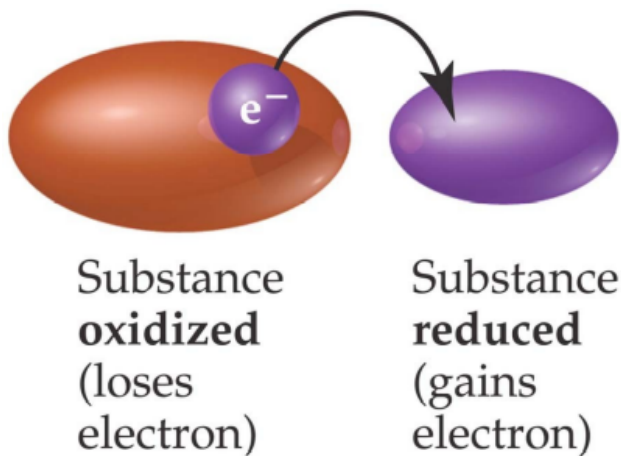


Figure 2: Oxidation-Reduction Reactions

0.16 Oxidation Numbers

- Elements in their elemental form have an oxidation number of 0.
- The oxidation number of a monatomic ion is the same as its charge.
- Nonmetals tend to have negative oxidation numbers, although some are positive in certain compounds or ions.
- Oxygen has an oxidation number of -2 , except in the peroxide ion in which it has an oxidation number of -1 .
- Hydrogen is -1 when bonded to a metal, $+1$ when bonded to a nonmetal.
- Fluorine always has an oxidation number of -1 .
- The other halogens have an oxidation number of -1 when they are negative; they can have positive oxidation numbers; however, most notably in oxyanions (e.g., ClO_4^-).
- The sum of the oxidation numbers in a neutral compound is 0.
- The sum of the oxidation numbers in a polyatomic ion is the charge on the ion.

0.17 Acids and Bases

There are 3 definitions of acids and bases:

- Arrhenius
- Brønsted
- Lewis

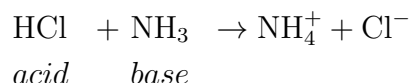
The original one was Arrhenius. This describes the strong acids and bases, which was all they knew of at the time.

0.17.1 Arrhenius

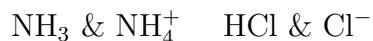
Acids contain H^+ , bases contain OH^- . This worked well until it was found that NH_3 was a base - but no OH^- in sight. So ... redefine bases.

0.17.2 Brønsted

Acids are H^+ donors, bases are H^+ acceptors. So acids have stayed essentially the same, but bases are now defined by the way they interact with acids. Acid base reactions are now seen as the transfer of an H^+ ion.



This also means that acids and bases come in pairs, called conjugate pairs. Notice that the two conjugates differ by an H^+ ion.



Some chemicals can act as both an acid & a base – we call them amphoteric. For example: The Brønsted definition works for most acid-base reactions, so to predict the products of an acid-base reaction, you transfer an H^+ from one chemical (the acid) to another (the base). And then ... the found acids that do not contain H^+ , like AlCl_3 .

0.17.3 Lewis definition

It seemed reasonable that acid-base chemistry should be described by electrons, all other reactions are. So ...

Lewis acid e^- pair acceptor

Lewis base e^- pair donor

Notice that all the bases we have seen before have lone pairs of e^- . When you draw the Lewis structure

So things that are only Lewis acids are those with empty orbitals, metal cations (Fe^{2+} , Cu^+) and Group 13 elements (Al, B, etc.)

These three definitions are a hierarchy. All Arrhenius acids are also Brønsted acids and Lewis acids. But not all Lewis acids are Brønsted and/or Arrhenius acids; the same is true for bases.