

## 4.1 Outline

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

## 4.2 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.

## 4.3 Concentration of Solute

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

$$\text{Molarity(M)} = \frac{\text{moles of solute}}{\text{liters of solution}} \quad (4.1)$$

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

## 4.4 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.

## 4.5 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?

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

## 4.6 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 ( $\text{HClO}_4$ ) to completely neutralize 25.00 mol calcium hydroxide ( $\text{Ca}(\text{OH}_2)$ ) solution.



A) How many moles of  $\text{HClO}_4$  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?

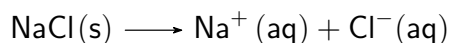
$$0.0251 \text{ mol perchloric acid} \times \frac{1 \text{ mol Ca}(\text{OH}_2)_2}{2 \text{ mol perchloric acid}} = 0.01255 \text{ Ca}(\text{OH}_2)_2$$

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

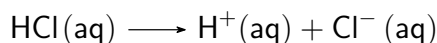
$$\frac{25.00 \text{ mol Ca}(\text{OH}_2)_2}{38.55 \text{ mL}} \times \frac{1,000 \text{ mL}}{1 \text{ L}} =$$

## 4.7 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.,  $\text{NaCl}$ ,  $\text{BaCl}_2$ , etc.)
- Only a few molecular compounds are capable of dissociating in water.
- For example,



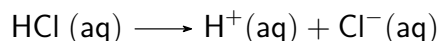
## 4.8 Electrolytes

- 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 ( $\text{CH}_3\text{OH}$ ), table sugar (sucrose;  $\text{C}_{12}\text{H}_{22}\text{O}_{11}$ )) that serve as nonelectrolytes in water.

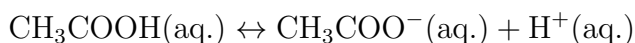
Table 4.1: Summary

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

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



e.g., the molecular compound acetic acid,  $\text{CH}_3\text{COOH}$ , 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 by only dimly lit.

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

## 4.9 Strong Electrolytes Are...

- Strong acids

**HCl** Hydrochloric

**HBr** Hydrobromic

**HI** Hydroiodic

**HClO<sub>3</sub>** Chloric

**HClO<sub>4</sub>** Perchloric**HNO<sub>3</sub>** Nitric**H<sub>2</sub>SO<sub>4</sub>** Sulfuric

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

Table 4.2: Solubility Guidelines for Common Ionic Compounds in Water

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

- Note that all common ionic compounds of the alkali metal ions (group 1A of the periodic table) and the ammonium ion (NH<sub>4</sub><sup>+</sup>) are soluble in water.

## 4.10 Precipitation Reactions

When ions that are insoluble (as could be predicted by the solubility guidelines in Table 4.2) are mixed and form compounds, a **precipitate** is formed.

The insoluble yellow solid formed on the right is lead iodide ( $\text{PbI}_2$ ).

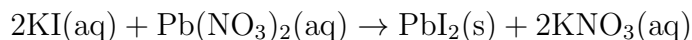
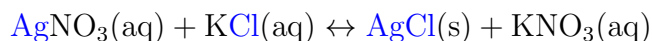


Figure 4.1: Lead Iodide

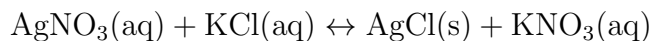
## 4.11 Metathesis (Exchange) Reactions

- 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!



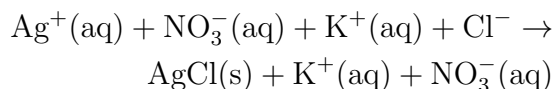
## 4.12 Molecular Equation

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



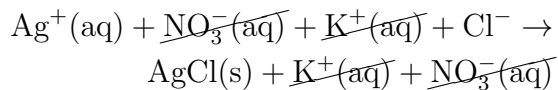
## 4.13 Ionic Equation

- 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.

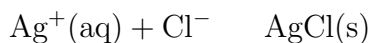


## 4.14 Net Ionic Equation

- 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.



## 4.15 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.

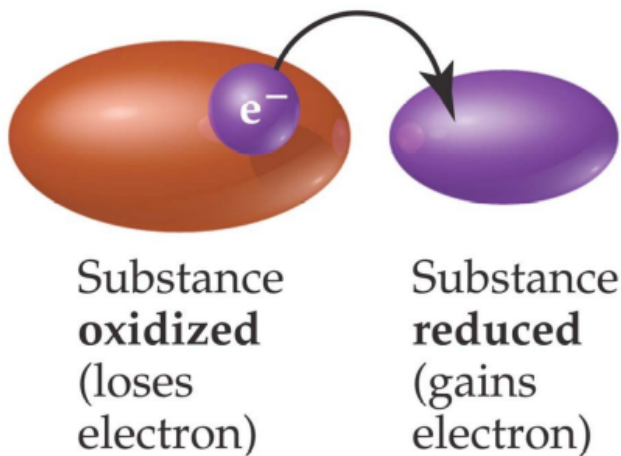


Figure 4.2: Oxidation-Reduction Reactions

## 4.16 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.,  $\text{ClO}_4^-$ ).
- 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.

## 4.17 Acids and Bases

There are 3 definitions of acids and bases:

- Arrhenius
- Brønsted
- Lewis

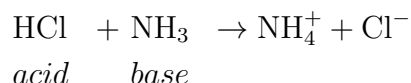
The original one was Arrhenius. This describes the strong acids and bases, which was all they knew of at the time.

### 4.17.1 Arrhenius

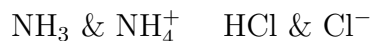
Acids contain  $\text{H}^+$ , bases contain  $\text{OH}^-$ . This worked well until it was found that  $\text{NH}_3$  was a base - but no  $\text{OH}^-$  in sight. So ... redefine bases.

### 4.17.2 Brønsted

Acids are  $\text{H}^+$  donors, bases are  $\text{H}^+$  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  $\text{H}^+$  ion.



This also means that acids and bases come in pairs, called conjugate pairs. Notice that the two conjugates differ by an  $\text{H}^+$  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  $\text{H}^+$  from one chemical (the acid) to another (the base). And then ... the found acids that do not contain  $\text{H}^+$ , like  $\text{AlCl}_3$ .

### 4.17.3 Lewis definition

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

**Lewis acid**  $\text{e}^-$  pair acceptor

**Lewis base**  $\text{e}^-$  pair donor

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

So things that are only Lewis acids are those with empty orbitals, metal cations ( $\text{Fe}^{2+}$ ,  $\text{Cu}^+$ ) 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.