

REACTIONS IN AQUEOUS

AP CHEMISTRY 20



4.1

GENERAL PROPERTIES OF AQUEOUS SOLUTIONS

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GENERAL PROPERTIES

- A **solution** in which water is the dissolving medium is called an **aqueous solution**.
- In this unit, we examine the chemical reactions taking place in this aqueous solutions.
- A solution is a homogeneous mixture of two or more substances.
 - A **solvent** is the substance in the greatest quantity (often water)
 - A **solute** is the substance that is dissolved in the solvent.

ELECTROLYTIC PROPERTIES

- We learn, very early, that electrical devices in water is a bad combination – which is useful as most of the water in our lives allows an electrical charge to travel through it and cause damage.
- However, pure water is a bad conductor, it is only the substances dissolved in water that make it dangerous when electrical currents are nearby.

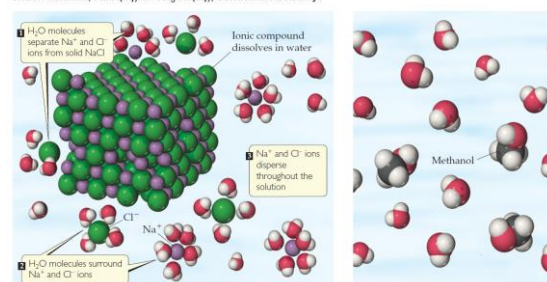
ELECTROLYTIC PROPERTIES

- For electrical charge to travel through liquid, there must be a **current between the two electrodes**. The conductivity of pure water is not enough to complete the circuit.
- It is only when **ions** are introduced that electrical charge can travel through water. These ions, carry electrical charge from one another and complete the circuit.
 - This is why sodium chloride (formed of ions) make water more conductive, than sugar (a molecule). Sugar dissolves but does not form a charged species like sodium (i.e., Na^+ and Cl^-)

ELECTROLYTIC PROPERTIES

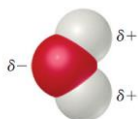
- A substance (such as NaCl) whose aqueous solution contains ions is called an **electrolyte**.
 - When NaCl breaks down it creates **solvated ions**.
- A substance (such as glucose [sugar]) that does not form ions in an solution is called a **nonelectrolyte**.

Which solution, NaCl(aq) or CH₃OH(aq), conducts electricity?



ELECTROLYTIC PROPERTIES

- Water, despite being an electrically neutral molecule, is an **effective solvent for ionic compounds**.
- This is due to the electron rich oxygen (creating a partial negative charge denoted by δ^-), and deficient hydrogen denoted (creating a partial positive δ^+).
- **Cations** are removed from the lattice by the '**negative**' oxygen, while the **anions** are attracted to the '**positive**' hydrogen.



ELECTROLYTIC PROPERTIES

- **Solvation** of ionic compounds increases the stability (meaning it's favored energetically – a driving force for all of chemistry) and prevents the return to the solid ionic structure.
- For most **ionic solids**, the ions are easily predicted from the **chemical formula**.
 - Sodium Sulfate (Na_2SO_4) breaks down in 2Na^+ and SO_4^{2-} . To show these are dissolved in water they are written as $\text{Na}^+(\text{aq})$ and $\text{SO}_4^{2-}(\text{aq})$.

EXAMPLE

- Write the solvation equation for the following substances:
 - potassium chloride
 - Lead(II) nitrate
 - Iron(III) iodate
 - Propanol

MOLECULAR COMPOUNDS

- **Most molecular compounds** (e.g., sugar, methanol, butanol) **do not form ions** when dissolved in water. They remain neutral and make non-electrolytes.
- One group of **molecular compounds that can dissolve in water to make ions are acids**.
 - For example $\text{HCl}(\text{g})$ dissolves in water to form hydrochloric acid, $\text{HCl}(\text{aq})$, it **ionizes**, that is, it dissociates into $\text{H}^+(\text{aq})$ and $\text{Cl}^-(\text{aq})$.

STRONG AND WEAK ELECTROLYTES

- **Strong electrolytes** that exist in solution completely or nearly completely as ions (e.g., NaCl or HCl).
- **Weak electrolytes** are compounds that exist in solution mostly as neutral solutes (e.g., acetic acid – only about 1% of AA will ionize).

STRONG AND WEAK ELECTROLYTES

- It is important to note the **solubility** (the amount of substance that will dissolve in a given amount of water) is not a measure of a substance's **electrolytic** property.
- A.A. is highly soluble, but weak electrolyte, while calcium hydroxide is a strong electrolyte but has low solubility.

WEAK ELECTROLYTES

- When **weak electrolytes dissolve in water** and ionize to some extent we write the equation using a **double arrow**. This indicates that some products are reforming to make the original reactant.
- This creates a *chemical equilibrium* in which the relative number of reactants and products are stable over time.



STRONG ELECTROLYTES

- For **strong electrolytes**, the products are far more stable than the reactants, and therefore will not reform the reactants.
- For strong electrolytes chemists use a **single arrow** showing the reaction.
- Again, *water soluble ionic compounds are strong electrolytes*.
 - Ionic compounds are characterized by the presence of a metal and a non-metal.
 - Compounds that include ammonium (NH_4^+) are considered to be strong electrolytes too.



EXAMPLE

A circuit is set up with a solution that is part of the electrical pathway. Three solutions are tested and cause a light bulb to have different levels of brightness (unlighted, dim and bright). The solutions were CH_3OH , NaOH and HCH_3COO – **which corresponds to each?**

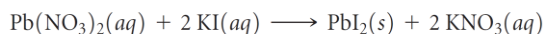
4.2

PRECIPITATION REACTIONS

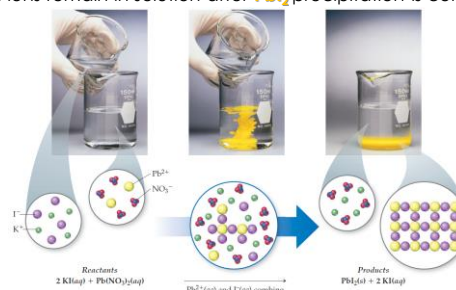
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PRECIPITATION REACTIONS

- Reactions that result in the product of an insoluble product are called **precipitation reactions**.
- A **precipitate** is an insoluble solid formed by a reaction in solution.
 - A common exemplar of precipitation reaction is the formation of $\text{PbI}_2(s)$ which is insoluble in water.



Which ions remain in solution after PbI_2 precipitation is complete?



PRECIPITATION REACTIONS

- Precipitates form** when the **attractive forces** of the ionic solid are **greater than the water's** ability to pull apart the solid.
- The **solubility** of a substance is how much will dissolve in a given solvent at the given temperature.
 - Any substance that has a solubility of 0.01 mol/L is considered to be insoluble.

SOLUBILITY GUIDELINES

- Unfortunately there are no rules basic** on simple physical properties such as ionic charge to help.
- There are some trends but each rule must be memorized.
 - Fun fact:** AP expects students to have their solubility rules memorized!

Soluble Ionic Compounds	Important Exceptions
Compounds containing	NO_3^- None
	CH_3COO^- 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+} , and Ba^{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+} , and Ba^{2+}

Solubility Rules

EXAMPLE

Predict if the following ionic compounds are soluble in water:

- Sodium carbonate
- Calcium hydroxide
- Mercury(I) iodide
- Lead(II) sulfate

EXAMPLE

Predict if the following ionic compounds are soluble in water:

- a. cobalt(II) hydroxide
- b. barium nitrate
- c. ammonium phosphate

PREDICTING PRECIPITATES

1. Note the ions present in the reactants.
2. Consider the possible cation-anion combinations.
3. Use solubility table to determine the solubility of the products.
 - Will a precipitate form when magnesium nitrate is mixed with sodium hydroxide?

EXCHANGE (METATHESIS) REACTIONS

- These are reactions in which ions are exchanged.
 - Metathesis, root is in Greek, meaning to 'transpose'.
- These reactions can only be balanced after the products are predicted.



METATHESIS REACTIONS

To complete and balance metathesis reactions, we follow these steps:

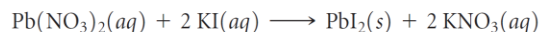
1. Use the chemical formula to **determine which ions** are present.
2. Write the **chemical formula of the products** by combining the cation from one reactant with the anion of the other, using the ionic charges to determine the subscripts in the chemical formula.
3. **Check the solubility**, for a precip. reaction there must be at least one insoluble product.
4. **Balance the equation**.

EXAMPLE

(a) Predict the identity of the precipitate that forms when barium chloride and potassium sulfate are mixed. (b) Write the balanced chemical equation for the reaction.

IONIC EQUATIONS

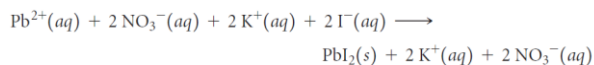
- When working with solutions it is useful to consider whether substances exist predominantly as ions or molecules. Consider the reaction between lead(II) nitrate and potassium iodide again.



- The reaction shown above, which shows the **complete chemical formula of reactants** and products is called a **molecular equation** because it shows the chemical formulas without indicating ionic character.

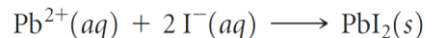
IONIC EQUATIONS

- Because lead(II) nitrate, potassium iodide and potassium nitrate are all strong electrolytes we can write them in the following form, called **complete ionic equation**.
- Notice that $K^+(aq)$ and $I^-(aq)$ appear on both sides, we call ions that do not participate in the reaction **spectator ions**.



NET IONIC EQUATION

- We can remove these spectator ions (they cancel out like algebraic expressions) leaving the **net ionic equation** which shows only the ions involved in the reaction.
- If there is no product that is insoluble then no reaction occurs.



EXAMPLE

- Write the net ionic equation for the precipitation reaction that occurs when aqueous solutions of calcium chloride and sodium carbonate are mixed.

EXAMPLE

- Write the net ionic equation for the precipitation reaction that occurs when aqueous solutions of silver nitrate and potassium phosphate are mixed.

PRACTICE PROBLEMS

Section 4.2 – Precipitation Reactions

4.19, 4.21, 4.23, 4.24, 4.25 & 4.27

HWC Friday Nov. 16th

4.3

ACID, BASES AND NEUTRALIZATION REACTIONS

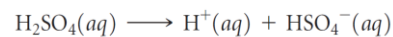
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ACIDS, BASES AND NEUTRALIZATION REACTIONS

- **Acids** are substances that ionize in aqueous solution to form **hydrogen ions $H^+(aq)$** . Because a hydrogen consists of just a proton and electron, H^+ , is simply a proton, thusly, acids are called **proton donors**.
- Acids are solvated in water similar to that of ionic compounds.
- The **number of hydrogen ions released** depends on the type of material donating the proton.

ACIDS

- For example: both HCl and HNO_3 are monoprotic acids, while sulfuric is diprotic.
- The ionization of diprotic acids (e.g., H_2SO_4) occurs in two steps.
 - While H_2SO_4 is a strong electrolyte, HSO_4^- is not and shown by the double arrow. When sulfuric acid breaks down it will become a mixture of H^+ , HSO_4^- , and SO_4^{2-}



MONOPROTIC ACIDS

Hydrochloric acid, HCl Nitric acid, HNO_3

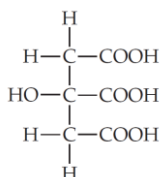
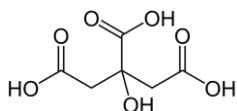
ACIDS

- While many organic (molecular) compounds will not ionize some do. Notably carboxylic acids (molecules with the 'COOH' structure) will donate a single hydrogen.
- Acetic Acid (CH_3COOH) donates the hydrogen part of the COOH group but not the hydrogen bonded to the hydrogen.

Acetic acid, CH_3COOH

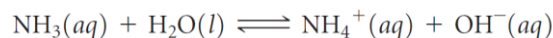
DISCUSSION

How many protons would citric acid release?



BASES

- Bases are substances that accept (react with) H^+ ions. Bases produce hydroxide ions (OH^-) when they dissolve in water.
 - $NaOH$, KOH , $Ca(OH)_2$ are all examples of bases.
- Some compounds that do not have OH^- can be a base (e.g., NH_3) if they are able to accept an H^+ ion.
 - $NH_3(aq)$ is a weak base because only about 1% forms $NH_4^+(aq)$



STRONG ACIDS AND BASES

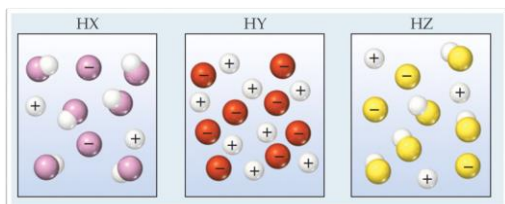
- Acids and bases that **completely ionize** in water are considered to be **strong electrolytes**.
- Acids and bases that **partially ionize** in water are considered to be **weak electrolytes**.
 - These must be memorized!

STRONG ACIDS & BASES

Strong Acids	Strong Bases
Hydrochloric, HCl	Group 1A metal hydroxides [LiOH, NaOH, KOH, RbOH, CsOH]
Hydrobromic, HBr	Heavy group 2A metal hydroxides [Ca(OH) ₂ , Sr(OH) ₂ , Ba(OH) ₂]
Hydroiodic, HI	
Chloric, HClO ₃	
Perchloric, HClO ₄	
Nitric, HNO ₃	
Sulfuric, H ₂ SO ₄	

EXAMPLE

Which acid is the strongest?



IDENTIFYING STRONG OR WEAK

- Using the information we have learned so far we are able to classify matter into three groups: strong electrolytes, weak electrolytes and non-electrolytes.

TABLE 4.3 • Summary of the Electrolytic Behavior of Common Soluble Ionic and Molecular Compounds

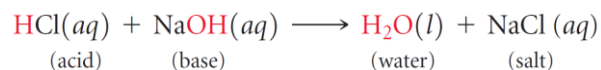
	Strong Electrolyte	Weak Electrolyte	Nonelectrolyte
Ionic	All	None	None
Molecular	Strong acids (see Table 4.2)	Weak acids, weak bases	All other compounds

EXAMPLE

- Classify these dissolved substances as strong, weak or non-electrolytes: CaCl₂, HNO₃, C₂H₅OH (ethanol), HCOOH(formic acid), KOH

NEUTRALIZATION REACTIONS

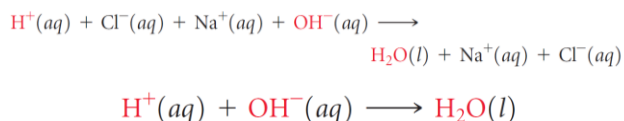
- When a solution of an acid and the solution of a base are mixed, a **neutralization reaction** occurs.
- The products of a neutralization reaction have none of the properties of an acid (sour) or base (bitter)



PRODUCTS

- Salts is an ionic compound that whose cation comes from a base (e.g., Na^+ from NaOH) and an anion that comes from an acid (e.g., Cl^- from HCl).
- In general, neutralization reaction between an acid and a metal hydroxide produces water and a salt.

COMPLETE & NET IONIC EQUATION



EXAMPLE

Consider the reaction between magnesium hydroxide (solid) and hydrochloric acid:

Molecular equation:



Net ionic equation:



EXAMPLE

For reaction between aqueous solutions of acetic acid (CH_3COOH) and barium hydroxide, $\text{Ba}(\text{OH})_2$, write (a) balanced molecular equation, (b) the complete ionic equation, (c) net ionic equation

EXAMPLE

For the reaction between phosphoric acid and potassium hydroxide, write (a) the balanced molecular equation and (b) the net ionic equation.

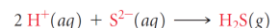
N.R WITH GAS FORMATION

- Some reactions produce low solubility gases as a product.
- These gases (commonly involving the **sulfide ion**, S^{2-} , or **carbonate ion**, CO_3^{2-} .) appear often in specific acid reactions.

Molecular equation:

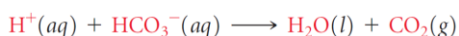
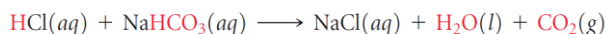
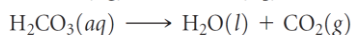


Net ionic equation:



GAS FORMATION

- Carbonates form carbonic acid, which is unstable, and decomposes into water and carbon dioxide according to the following reactions.



PRACTICE PROBLEMS

Section 4.3 – Acid, Bases and Neutralization Reactions

4.29 – 4.31, 4.35, 4.37, 4.39 & 4.41

Due Monday, Nov. 19

4.4

OXIDATION-REDUCTION REACTIONS

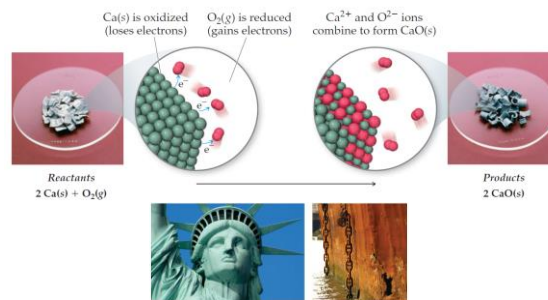
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OXIDATION-REDUCTION REACTIONS

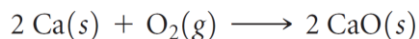
- Precipitation** reactions produce an insoluble product.
- Neutralization** reactions are those in which hydrogen ions combine with hydroxide ions to produce water.
- There is a third type of reaction called a **oxidation-reduction reaction** in which electrons are transferred.
 - These are often referred to as redox reactions.

REDOX REACTIONS

- One of the most common examples of redox reactions is the **corrosion of metals** (the conversion of pure metal to a mixture).
- The color of the dome on the parliament building, rust and even silver tarnish are examples of redox chemistry.
 - The loss of electrons by a substance is called **oxidation**.
 - The gain of electrons by a substance is called **reduction**.
- Oxidization and reduction always occur at the same time, one part of the reaction cannot happen on it's own.



EXAMPLE



- What is the element being oxidized?
- What is the element that is being reduced?
- What is the reducing species?
- What is the oxidizing species?

OXIDATION NUMBERS

- To understand redox reactions we need to be able to determine the electron change in each species in a reaction. In order to do this we assign oxidation numbers using the following rules:
 - For an atom in its elemental form it's always 0.
 - For any monoatomic ion, the oxidation number is equal to the charge.
 - Non-Metals usually have negative oxidation numbers.
 - The oxidation number of oxygen is usually -2 in both ionic and molecular compound. The major exception is peroxides, where it has an oxidation number of -1.
 - Hydrogen is usually +1, unless connected to metals, where it is -1.
 - The oxidation number of fluorine is -1 in all compounds, halogens usually have a -1 in most binary compounds. When they are combined with oxygen, as in oxyanions, they have positive oxidation states.
 - The sum of oxidation numbers of all atoms in a neutral compound is zero.

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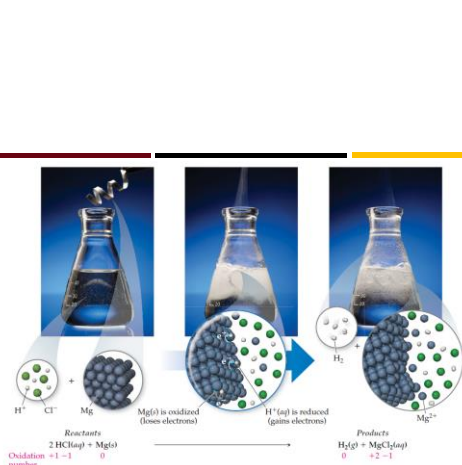
EXAMPLE

Determine the oxidation numbers for all elements in the following:

- | | |
|-----------------------------|-----------------------|
| a) H_2S | e) SO_4^{2-} |
| b) S_8 | f) HClO_4 |
| c) SCl_2 | g) HCl |
| d) Na_2SO_3 | h) HClO_3 |

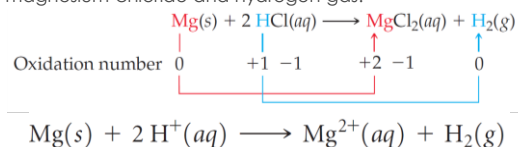
OXIDATION OF METALS

- The reaction between a metal and either an acid or a metal salt conforms to the following pattern, these are called **displacement** reactions:



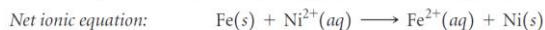
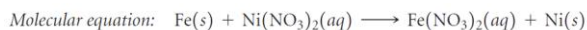
DISPLACEMENT REACTIONS

- Many metals undergo displacement reactions with acids, producing salts and hydrogen gas. In the previous reaction magnesium ribbon reacts with hydrochloric acid to produce magnesium chloride and hydrogen gas.



DISPLACEMENT REACTIONS

- Metals can also be oxidized by various other aqueous solutions of salts.



- In this reaction the iron is oxidized from 0 to +2 while the nickel is reduced from +2 to 0. The nickel is a spectator ion as it is not oxidized nor reduced.

EXAMPLE

- (a) Write the balanced molecular and net ionic for the reaction between magnesium and cobalt(II) sulfate.
 (b) What is oxidized and what is reduced in the reaction?

EXAMPLE

- Write the balanced molecular and net ionic equation for the reaction of aluminium with hydrobromic acid.

EXAMPLE

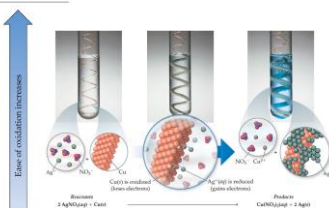
THE ACTIVITY SERIES

- Metals are able to oxidize one another but not every metal is able to oxidize the next. To predict the ease of oxidation we can use an **activity series** (pg. 136).
- Metals near the top are considered to be **active metals** (alkali/alkaline metals) where as group 8B and 1B are considered to be **noble metals** as they resist oxidation.
- Any metal on the list can be oxidized by the ions of elements below it.** (e.g., copper can be oxidized by silver)



TABLE 4.2 Activity Series of Metals in Aqueous Solution

Metal	Oxidation Reaction
Lithium	$\text{Li}(s) \longrightarrow \text{Li}^+(aq) + e^-$
Potassium	$\text{K}(s) \longrightarrow \text{K}^+(aq) + e^-$
Barium	$\text{Ba}(s) \longrightarrow \text{Ba}^{2+}(aq) + 2e^-$
Calcium	$\text{Ca}(s) \longrightarrow \text{Ca}^{2+}(aq) + 2e^-$
Sodium	$\text{Na}(s) \longrightarrow \text{Na}^+(aq) + e^-$
Magnesium	$\text{Mg}(s) \longrightarrow \text{Mg}^{2+}(aq) + 2e^-$
Aluminum	$\text{Al}(s) \longrightarrow \text{Al}^{3+}(aq) + 3e^-$
Manganese	$\text{Mn}(s) \longrightarrow \text{Mn}^{2+}(aq) + 2e^-$
Zinc	$\text{Zn}(s) \longrightarrow \text{Zn}^{2+}(aq) + 2e^-$
Chromium	$\text{Cr}(s) \longrightarrow \text{Cr}^{3+}(aq) + 3e^-$
Iron	$\text{Fe}(s) \longrightarrow \text{Fe}^{2+}(aq) + 2e^-$
Cobalt	$\text{Co}(s) \longrightarrow \text{Co}^{2+}(aq) + 2e^-$
Nickel	$\text{Ni}(s) \longrightarrow \text{Ni}^{2+}(aq) + 2e^-$
Tin	$\text{Sn}(s) \longrightarrow \text{Sn}^{2+}(aq) + 2e^-$
Lead	$\text{Pb}(s) \longrightarrow \text{Pb}^{2+}(aq) + 2e^-$
Hydrogen	$\text{H}_2(g) \longrightarrow 2\text{H}^+(aq) + 2e^-$
Copper	$\text{Cu}(s) \longrightarrow \text{Cu}^{2+}(aq) + 2e^-$
Silver	$\text{Ag}(s) \longrightarrow \text{Ag}^+(aq) + e^-$
Mercury	$\text{Hg}(l) \longrightarrow \text{Hg}^{2+}(aq) + 2e^-$
Platinum	$\text{Pt}(s) \longrightarrow \text{Pt}^{2+}(aq) + 2e^-$
Gold	$\text{Au}(s) \longrightarrow \text{Au}^3+(aq) + 3e^-$



EXAMPLE

Will an aqueous solution of iron(II) chloride oxidize magnesium metal? If so, write the balanced molecular and net ionic equations for the reaction.

EXAMPLE

- Which of the following metals will be oxidized by lead(II) nitrate?
 - Zinc
 - Copper
 - Iron

CHECKPOINT

- In this section there are many types of reactions that may occur producing a number of different products. While it may seem easy with a group or through guided instruction you must be able to predict products on your own. Consider the following:
 - What are the reactants?
 - Are they electrolytes or non-electrolytes
 - Are they acids or bases?
 - If the reactants are electrolytes, will metathesis produce a precipitate? Water? Gas?
 - If metathesis cannot occur, can the reactants engage in an oxidation-reduction reaction? This requires that both a reactant can be oxidized and a reactant be reduced.

PRACTICE PROBLEMS

Section 4.4 – Acid, Bases and Neutralization Reactions

4.45 – 4.51, 4.53, 4.55, 4.58

4.5

CONCENTRATION OF SOLUTIONS

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CONCENTRATION

- **Concentration** is the amount of solute dissolved in a given quantity of solvent or quantity of solution.
- The greater the amount of solute dissolved the more **concentrated** the solution is.
- **Molarity** expresses the concentration of a solution as the number of moles of solute in a liter of solution.

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



DISCUSSION

- Which is more concentrated, a solution prepared by dissolving 21.0 g of NaF (0.500 mol) in enough water to make a 500 mL solution or a solution prepared by adding 10.5 g (0.250 mol) of NaF in enough water to make a 100 mL of solution?

EXAMPLE

- Calculate the molarity of a solution made by dissolving 23.4 g of sodium sulfate in enough water to form 125 mL of solution.

EXPRESSING CONCENTRATION OF ELECTROLYTES

- When electrolytes dissolved in water the strength of the resultant ions depends on the chemical formula.
 - For example a solution of NaCl that is 1.0 M has 1.0 M of Na^+ and Cl^- ions.
 - A solution that is 1.0 M Na_2SO_4 contains 2.0 M of Na^+ and 1.0 M of SO_4^{2-} .

EXAMPLE

- What is the molar concentration of each ion present in 0.0250 M aqueous solution of calcium nitrate?
- What is the molar concentration of potassium ions in a 3.0 M solution of potassium carbonate?

MOLES, MOLARITY AND VOLUME

- When two are known, the third can be solved for.
- We can determine the number of moles in a 2.0 L solution of 0.200 M solution of nitric acid by using a conversion factor.

$$\text{Moles HNO}_3 = (2.0 \text{ L soln}) \left(\frac{0.200 \text{ mol HNO}_3}{1 \text{ L soln}} \right) = 0.40 \text{ mol HNO}_3$$

EXAMPLE

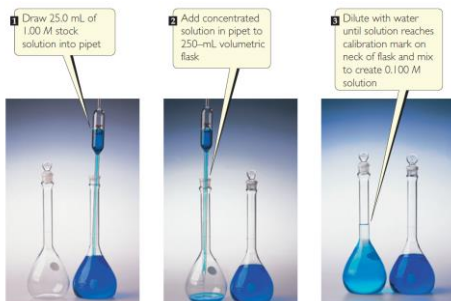
- How many grams of Na_2SO_4 are required to make a 0.350L solution of 0.500 M Na_2SO_4 ?

EXAMPLE

(a) How many grams of Na_2SO_4 are there in 15 mL of a 0.50 M Na_2SO_4 ? (b) How many millilitres of 0.50 M Na_2SO_4 solution are needed to provide 0.038 mol of this salt?

DILUTION

- For the most part laboratories often purchase **stock solution** where are then diluted to the correct strength.
- A **dilution** is the process of making a less concentrated solution from a concentrated one.
 - It is important to note while the concentration changes, the amount of solute (number of moles) does not.



DILUTIONS

- In laboratory situations, calculations of this sort are often made this way using an equation that derived from the understanding the number of moles does not change, only the concentration and the volume.

Moles solute in conc soln = moles solute in dilute soln

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

EXAMPLE

How many millilitres of 3.0 M sulfuric acid is needed to make 450 mL of a 0.10 M solution?

EXAMPLE

- What volume of 2.50 M lead(II) nitrate solution contains 0.500 mol of Pb^{2+} ?
- How many millilitres of 5.0 M potassium dichromate must be diluted to prepare 250 mL of 0.10 M solution?
- If 10.0 mL of a 10.0 M stock solution is NaOH is diluted to 250. mL, what is the concentration of the resulting stock solution?

PRACTICE PROBLEMS

Section 4.5 – Solutions

4.60 – 4.64, 4.71 – 4.73 & 4.75 – 4.78

SOLUTION STOICHIOMETRY

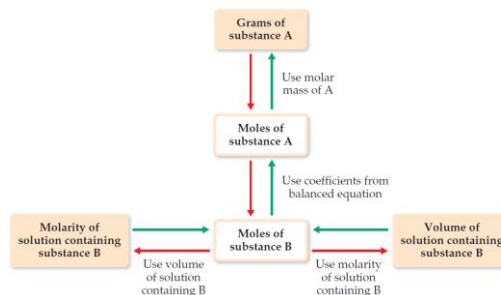
- Using what we know about **balanced chemical reactions** (stoichiometry) and **solutions** (molarity \times volume = moles) we can predict quantities of needed reactant or products.
- In order to achieve this we have to convert to a common unit, **the mole**.

EXAMPLE

- How many grams of $\text{Ca}(\text{OH})_2$ are needed to neutralize 25.0 mL of 0.100 M HNO_3 ?

4.6 SOLUTION STOICHIOMETRY AND CHEMICAL ANALYSIS

AP CHEMISTRY 20

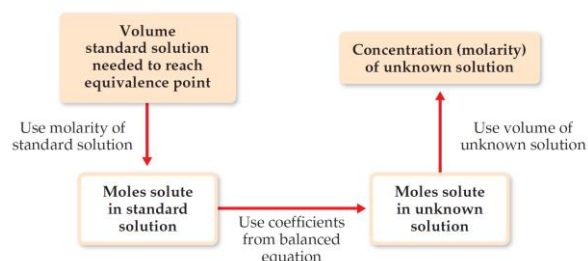
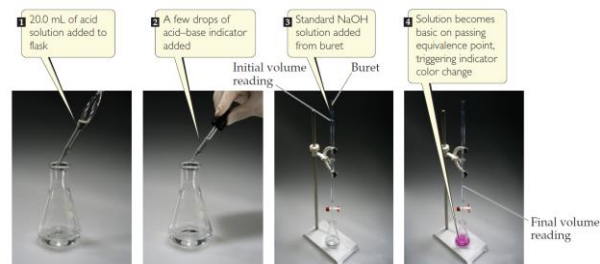


EXAMPLE

- How many grams of NaOH are needed to neutralize 20.0 mL of 0.150 M sulfuric acid solution (assume sulfuric acid is diprotic and completely ionizes).
- How many liters of a 0.500 M $\text{HCl}(\text{aq})$ are needed to react completely with 0.100 mol of $\text{Pb}(\text{NO}_3)_2(\text{aq})$, forming a precipitate of $\text{PbCl}_2(\text{s})$?

TITRATIONS

- To determine the concentration of a particular solution in a solution, chemists often carry out a **titration**, which involves the combination of a solution where the solute concentration is not known with a reagent solution of known concentration, called a **standard solution**.
- The point where the concentration of the solute is completely reacted (stoichiometrically equivalent) is called the **equivalence point**.



TITRATIONS

- Titration**s can be conducted using **neutralization**, **precipitation reactions** or **oxidation-reduction** reactions.
- Color change, gas formed, pH change (a measure of a substance's H⁺ levels) can all be used to indicate the **equivalence point**.

EXAMPLE

One commercial method used to peel potatoes is to soak them in a NaOH solution for a short time, then remove them and spray off the peel. The NaOH concentration is normally 3 to 6 M, and the solution must be analyzed periodically. In one such analysis, 45.7 mL of 0.500 M H₂SO₄ is required to neutralize 20.0 mL of NaOH solution. What is the concentration of the solution?

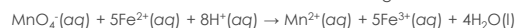
EXAMPLE

The quantity of dissolved chloride ions, Cl⁻(aq), in a municipal water supply is determined by titrating the same with silver ions to create an insoluble product. The end point in this reaction is marked by a change in color by a special type of indicator.

- Write the net ionic equation for this reaction.
- How many grams of chloride ions are in a sample of the water if 20.2 mL of 0.100 M Ag⁺ is needed to react with all the chloride ions in the sample?
- If the sample has a mass of 10.0 g, what is the percent Cl⁻ it contains?

EXAMPLE

- A sample of iron ore is dissolved in acid, and the iron is converted to Fe^{2+} . The same is then titrated with 47.20 mL of 0.02240 MnO_4^- solution. The oxidation-reduction reaction that occurs during the titration is



- How many moles of MnO_4^- were added to the solution?
- How many moles of Fe^{2+} were in the sample?
- How many grams of iron were in the sample?
- If the same had a mass of 0.8890 g, what is the percentage of iron in the sample?

EXAMPLE

A sample of 70.5 mg of potassium phosphate is added to 15.0 mL of 0.50 M silver nitrate resulting in the formation of a precipitate.

- Write the molecular equation for the reaction
- What is the limiting reactant in the reaction?
- Calculate the theoretical yield, in grams, of the precipitate that forms.

PRACTICE PROBLEMS

Section 4.6 – Solution Stoic & Analysis

4.79 – 4.81, 4.83, 4.84, 4.86 & 4.88