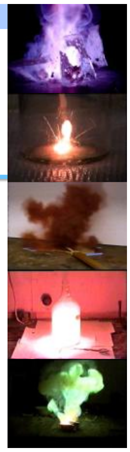
CHEM 1411 General Chemistry — *Chemistry: An Atoms First Approach*

10 Types of Chemical Reactions and Solution Stoichiometry



Chapter Objectives:

- Learn how to use molarity to perform calculations involving solution stoichiometry.
- Learn how to recognize and predict products in precipitation, neutralization, redox, and other types of chemical reactions.

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Introduction

- Much of the chemistry (both biological and non-biological) that takes place on Earth involves **water** in some fashion:
  - Almost 75% of the Earth's surface is covered by water or ice.
  - About 66% of the human body consists of water.
  - A lot of important chemistry takes place in **aqueous solution**, in which the solvent is water.
- In this chapter, we'll see how some types of chemical reactions take place and how we can organize chemical reactions into different types. Most of these reactions will take place in aqueous solutions.

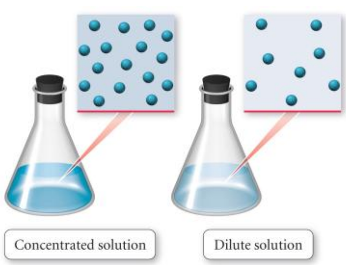
Solution Stoichiometry

Solutions

- For a chemical reaction to occur, the reacting species have to come in close contact with each other. Most chemical reactions are performed in a *solution* (or in the gas phase) rather than in the solid state.
- A **solution** consists of a smaller amount of one substance, the **solute** (usually a liquid or solid), dissolved in a larger amount of another substance, the **solvent** (usually a liquid).
  - Other kinds of solutions, such as of two or more solids (e.g., metal alloys), or gases dissolved in solids, or gases dissolved in other gases (e.g., the atmosphere), are also possible.
- Solutions in which water is the solvent are known as **aqueous solutions**.

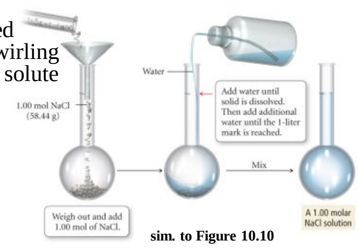
Dilute and Concentrated Solutions

- A solution that contains a small amount of solute relative to the solvent is a **dilute solution**.
- A solution that contains a large amount of solute relative to the solvent is a **concentrated solution**.



Making Solutions of a Desired Molarity

- Because the volume of a solution comes from the solute *and the* solvent, a 1 molar solution cannot be made by adding one mole of solute to 1 L of solvent.
- Solutions of a desired molarity are usually prepared by placing the appropriate amount of solute in a *volumetric flask*, and adding solvent until a calibrated final volume is reached (with frequent swirling to make sure the solute dissolves).



sim. to Figure 10.10

Solution Dilution

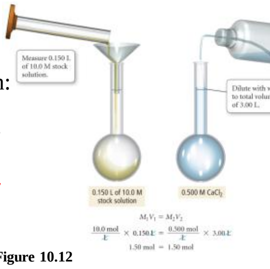
- Solutions can also be prepared by **diluting** a more concentrated stock solution.

Concentrated solution + Solvent → Dilute solution

- The initial molarity (M_1) and volume (V_1) of a concentrated solution are related to the final molarity (M_2) and volume (V_2) of a dilute solution by the equation:

$$M_1 V_1 = M_2 V_2$$

Note that the units for volume and concentration don't actually matter in this equation.



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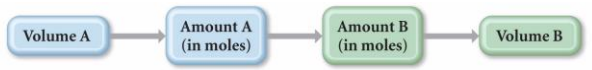
Solution Stoichiometry — Molarity

- We must know the amount of material present in a certain volume of solution — the **concentration** — in order to perform measurements and calculations.
- A common unit of concentration is **molarity** (***M***), defined as the number of *moles of solute* per *liter of solution* (that's solution, not solvent!):

$$\text{Molarity}(M) = \frac{\text{moles of solute}}{\text{liters of solution}} = \text{mol/L} = \text{mol L}^{-1}$$

- The molarity of a solution can be used as a conversion factor to relate the solution volume to the number of moles of solute present.



Examples: Molarity

1. What is the molarity of a solution made by dissolving 2.355 g of sulfuric acid in water and diluting to a final volume of 50.0 mL?

Answer: 0.480 M H_2SO_4

Examples: Molarity

2. How many grams of solute are in 1.75 L of 0.460 M sodium monohydrogen phosphate?

Answer: 114 g

Examples: Molarity

3. How many liters of a 0.125 M NaOH solution contains 0.255 mol of NaOH? (Ex. 4.6)

Answer: 2.04 L soln.

Examples: Solution by Dilution

4. Isotonic saline is a 0.150 M aqueous solution of NaCl that simulates the total concentration of ions found in many cellular fluids. Its uses range from a cleansing rinse for contact lenses to a washing medium for red blood cells. How would you prepare 800. mL of isotonic saline from a 6.00 M stock solution?

Answer: 20.0 mL

Examples: Solution by Dilution

5. To what volume should you dilute 0.200 L of a 15.0 M NaOH solution to obtain a 3.00 M NaOH solution?

Answer: 1.00 L

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Examples: Stoichiometry of Reactions in Soln

6. Stomach acid, a dilute solution of HCl in water, can be neutralized by reaction with sodium bicarbonate according to the equation

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

How many mL of 0.125 M $NaHCO_3$ solution are needed to neutralize 18.0 mL of 0.100 M HCl?

Answer: 14.4 mL

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Strong Electrolytes, Weak Electrolytes, and Nonelectrolytes

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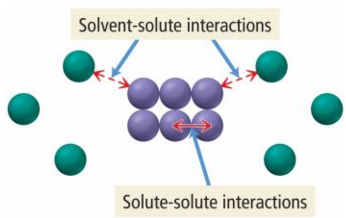
Solute and Solvent Interactions

- When a solid is put into a liquid solvent such as water, there is a competition between the forces of attraction among the particles of the solute (*solute-solute interactions*) and the forces of attraction between the solvent molecules and the particles in the solute (*solvent-solute interactions*). Which interactions are stronger determines whether the solute dissolves.



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The Role of Water as a Solvent

- Many reactions take place in *aqueous solution*. Water interacts with many substances, and plays an active role in many chemical processes.
- Many ionic substances dissolve in water (to at least some extent); water is very good at **solvating** (dissolving) cations and anions.
  - The O end of a water molecule has a slight negative charge (a **partial negative charge**, δ^-) while the H ends have slight positive charges (a **partial positive charge**, δ^+).
  - Water also has an overall **bent** shape.
  - The combined effects of polar bonds in a bent shape make water a **polar molecule**, having an uneven distribution of electrons.

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The Shape and Polarity of Water

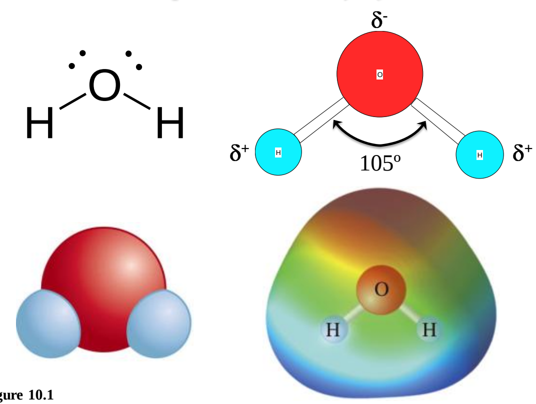

Figure 10.1

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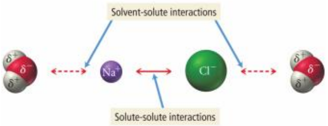
Water as a Solvent

- When sodium chloride is dissolved in water, the Na^+ and Cl^- ions are attracted to each other, and both are also attracted to water molecules.
  - The Na^+ ions are attracted to the partially negative oxygen atom.
  - The Cl^- ions are attracted to the partially positive hydrogen atoms.
  - The attraction between the water and the ions is greater than that of the ions for each other, and the sodium chloride dissolves.



Water as a Solvent

- The ions are separated from each other, and are free to move randomly through the solution, surrounded by a crowd of water molecules.

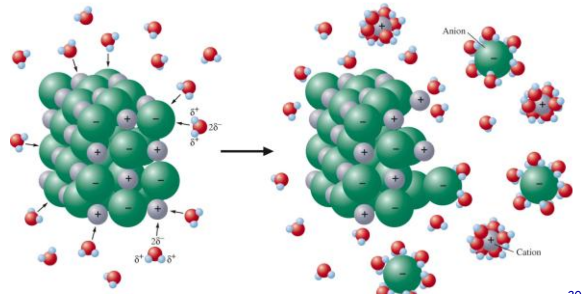


Figure 10.2 MOV: Dissolution of NaCl

Water as a Solvent

- Ethanol (ethyl alcohol, C_2H_5OH) dissolves in water because the partial charges on the H and O atoms in the molecule interacts strongly with the partial charges on the atoms in water.
  - The ethanol molecules themselves don't break apart.

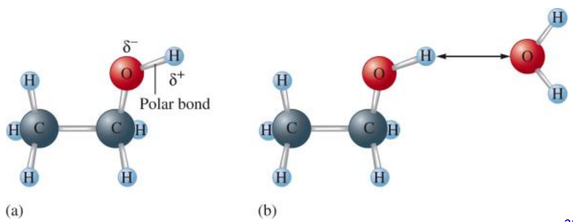
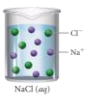


Figure 10.3

Electrolytes in Aqueous Solution

- Substances that dissolve in water can be classified as *electrolytes* or *nonelectrolytes*:
  - **Electrolytes** dissolve in water to form solutions which *conduct electricity*, because they dissociate to form mobile, solvated cations and anions:

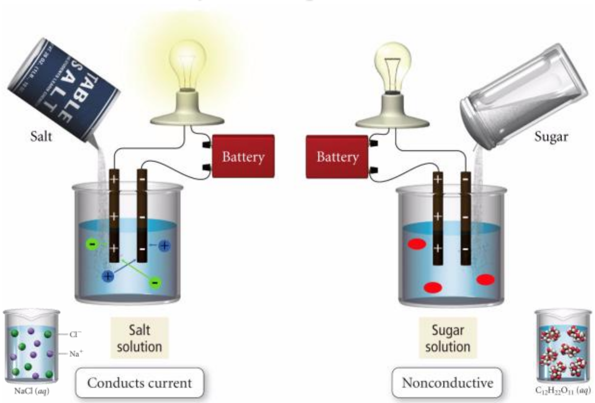


$$NaCl(s) \xrightarrow{H_2O} Na^+(aq) + Cl^-(aq)$$

  - **Nonelectrolytes** do *not conduct electricity* when dissolved in water, because they are molecules with don't dissociate into ions (water itself is a nonelectrolyte):

$$C_{12}H_{22}O_{11}(s) \xrightarrow{H_2O} C_{12}H_{22}O_{11}(aq)$$

Electrolytes in Aqueous Solution



Strong and Weak Electrolytes

- Electrolytes are divided into two categories:
  - **Strong electrolytes** are dissociated to a very large extent (70-100%); that is, virtually all of the units of the original substance are separated into ions.
  - **Weak electrolytes** are dissociated to a very small extent; that is, only a small percentage of the substance is dissociated into ions at any one time.
- Dissociation of weak electrolytes are often written using back-and-forth double arrows, $\rightleftharpoons$ indicating that a **dynamic equilibrium** is taking place, in which there is a forward and a backward direction to the process.

Strong and Weak Electrolytes

- For strong electrolytes, the equilibrium goes almost completely to the right, to the dissociated species.

$$\rightleftharpoons$$

- For weak electrolytes, the equilibrium lies to the left, with the undissociated species.

$$\leftrightharpoons$$

- Since most of the particles in a weak electrolyte solution are not ionized, weak electrolytes conduct electricity in solution only weakly.
- Equilibrium processes are important in many areas of chemistry, and will be discussed in excruciating detail later.

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Strong and Weak Electrolytes

Figure 10.4 (a) Many ions (b) Few ions (c) No ions

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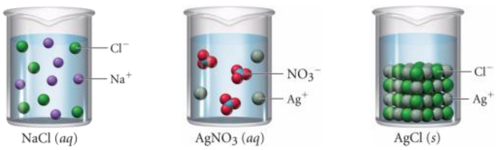
Soluble and Insoluble

- In some cases, the electrostatic forces in ionic compounds are too great to be overcome by water, and the substance is **insoluble** in water.
- This is greatly oversimplified, though; in reality, solubility is a continuum, and even "insoluble" substances dissolve in water to some slight extent.
  - Solubility of NaCl in H_2O at 25ºC = 357 g/L
  - Solubility of $AgNO_3$ in H_2O at 25ºC = 216 g/L
  - Solubility of AgCl in H_2O at 25ºC = 0.0019 g/L



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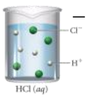
Acids as Electrolytes

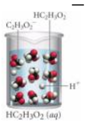
- Most molecular compounds are nonelectrolytes, except for the **acids**.
- Acids of the general formula HA **ionize** into H^+ ions and A^- ions when they dissolve in water:

$$HA(g, l) \xrightarrow{H_2O} H^+(aq) + A^-(aq)$$



- A **strong acid** such as HCl *ionizes completely*. Every molecule of HCl splits up into H^+ and Cl^-:

$$HCl(g) \rightarrow H^+(aq) + Cl^-(aq)$$



- A **weak acid** such as $HC_2H_3O_2$ *does not ionize completely*. Only a small percentage of the molecules split up into H^+ and $C_2H_3O_2^-$:

$$HC_2H_3O_2(g) \rightleftharpoons H^+(aq) + C_2H_3O_2^-(aq)$$

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Classifying Electrolytes

- **Strong Electrolytes:**
  - Soluble ionic compounds.
  - Strong acids: HCl, HBr, HI, HNO_3, $HClO_4$, H_2SO_4

memorize!

- **Weak Electrolytes:**
  - All other acids are weak acids/weak electrolytes: HF, $HC_2H_3O_2$, HNO_2, H_2SO_3, etc.
- **Nonelectrolytes:**
  - Molecular compounds (except for acids): water, sugar, most organic compounds, etc.

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Examples: Dissociation of Ionic Compounds

1. Write the equation for the dissociation in water of Na_2S, HBr, $AlCl_3$, and $Pb(NO_3)_2$.

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Examples: Dissociation of Ionic Compounds

2. How many moles of each ion are in 1 mol of $CaCl_2$?

Examples: Dissociation of Ionic Compounds

3. How many moles of each ion are in 23.4 g of Na_2S?

Answer: 0.600 mol Na^+, 0.300 mol S^{2-}

Examples: Dissociation of Ionic Compounds

4. What is the molar concentration of iron(III) ions and bromide ions in a 0.225 M aqueous solution of iron(III) bromide, assuming complete dissociation? What is the total molar concentration of ions?

Answer: 0.225 M Fe^{3+}, 0.675 M Br^-, 0.900 M ions

Examples: Classifying Electrolytes

5. Classify the following compounds as strong electrolytes, weak electrolytes, or nonelectrolytes.

- **a.** HCl
- **b.** SCl_2
- **c.** $AgNO_3$
- **d.** H_2O
- **e.** CH_3CH_2OH
- **f.** Na_2SO_4
- **g.** HNO_2
- **h.** HNO_3

Precipitation Reactions

Precipitation Reactions

- A **precipitation reaction** occurs when two ionic compounds react to produce a **precipitate**, an insoluble substance which falls out of the solution.
- These reactions are also known as **double-displacement** or **metathesis** reactions, because the cations and anions of the reactants "change partners" in the products:

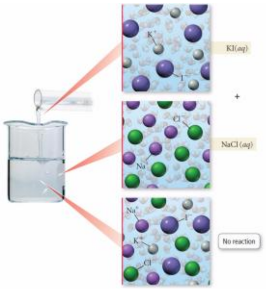
$$AB + CD \rightarrow AD + CB$$

- These reactions occur when the solute-solute attractions between the ions in the precipitate are stronger than the solvent-solute attractions.
- The other ions stay in solution as **spectator ions**.

Precipitation Reactions

- If we mix solutions of NaCl and KI, all we get is a mixture of dissolved ions; no reaction takes place because none of the products are insoluble in water:

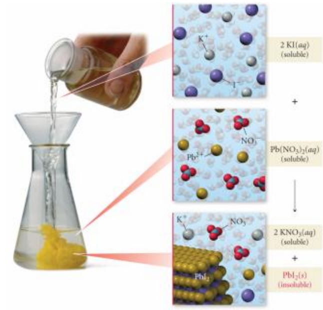
$NaCl(aq) + KI(aq) \rightarrow KCl(aq) + NaI(aq)$: **NR**



Precipitation Reactions

- On the other hand, if we mix $Pb(NO_3)_2$ and KI, we get a solid product of PbI_2:

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



Predicting Precipitation Reactions

$$AB(aq) + CD(aq) \rightarrow AD(aq) + CB(s)$$

- In order to predict whether a precipitation reaction occurs, we use the **solubility rules** on the next slides to determine whether any of the potential products of the reaction are insoluble.
  - If one of the products of the reaction is insoluble (or is a weak electrolyte, gas, or nonelectrolyte), the reaction occurs, and must be balanced appropriately.
  - If no insoluble product (or weak electrolyte, gas, or nonelectrolyte) form, then no reaction (NR) takes place, and all of the ions remain in solution as spectator ions.

Solubility Rules for Ionic Compounds in Water

TABLE 10.1 › Simple Rules for the Solubility of Salts in Water

1. Most nitrate (NO_3^-) salts are soluble.
2. Most salts containing the alkali metal ions (Li^+, Na^+, K^+, Cs^+, Rb^+) and the ammonium ion (NH_4^+) are soluble.
3. Most chloride, bromide, and iodide salts are soluble. Notable exceptions are salts containing the ions Ag^+, Pb^{2+}, and Hg_2^{2+}.
4. Most sulfate salts are soluble. Notable exceptions are $BaSO_4$, $PbSO_4$, Hg_2SO_4, and $CaSO_4$.
5. Most hydroxides are only slightly soluble. The important soluble hydroxides are NaOH and KOH. The compounds $Ba(OH)_2$, $Sr(OH)_2$, and $Ca(OH)_2$ are marginally soluble.
6. Most sulfide (S^{2-}), carbonate (CO_3^{2-}), chromate (CrO_4^{2-}), and phosphate (PO_4^{3-}) salts are only slightly soluble, except for those containing the cations in Rule 2.

Solubility Rules for Ionic Compounds in Water

Soluble Ionic Compounds

1. All common compounds of the Group 1A ions (Li^+, Na^+, K^+, etc.) and the ammonium ion (NH_4^+) are soluble.

2. All common nitrates (NO_3^-), acetates ($C_2H_3O_2^-$), and most perchlorates (ClO_4^-) are soluble.

3. All common chlorides (Cl^-), bromides (Br^-), and iodides (I^-) are soluble, *except* those of Ag^+, Pb^{2+}, and Hg_2^{2+}.

4. All common sulfates (SO_4^{2-}) are soluble, *except* those of Ca^{2+}, Sr^{2+}, Ba^{2+}, Pb^{2+}, and Ag^+.

memorize!

Solubility Rules for Ionic Compounds in Water

Insoluble Ionic Compounds

5. Most metal hydroxides (OH^-) are insoluble, *except* those of Group 1A, NH_4^+, which are soluble, and Ca^{2+}, Sr^{2+}, and Ba^{2+}, which are slightly soluble.

6. Most sulfides (S^{2-}), carbonates (CO_3^{2-}), chromates (CrO_4^{2-}), and phosphates (PO_4^{3-}) are insoluble, *except* those of Group 1A and NH_4^+.

memorize!

Examples: Predicting Solubility

1. Predict whether the following compounds are **soluble** or **insoluble** in water.

a. $CaCO_3$	**g.** $HgCl_2$
b. $Mg(OH)_2$	**h.** Hg_2Cl_2
c. Na_2S	**i.** $AgC_2H_3O_2$
d. $PbSO_4$	**j.** $Pb(NO_3)_2$
e. $(NH_4)_3PO_4$	**k.** K_2CO_3
f. $NH_4C_2H_3O_2$	**l.** AgCl

Representing Aqueous Reactions

- Aqueous reactions involving ionic compounds can be represented by three types of equations:
  - **Molecular equation (formula equation** in book): formulas are written as if they were intact molecules (e.g., NaCl)
  - **Complete ionic equation**: soluble strong electrolytes (soluble ionic compounds and strong acids) are written as ions (e.g., Na^+, Cl^-)
    - Insoluble precipitates, weak electrolytes, and molecules are left intact.
    - Ions that are not involved in the actual chemical change are called **spectator ions**.
  - **Net ionic equation**: shows only the ions which undergo a change, eliminating the spectator ions.

Molecular, Ionic, and Net Ionic Equations

Molecular equation:

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

Ionic equation: (phase labels omitted for clarity)

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

Spectator ions: K^+ and NO_3^-

Net ionic equation:

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

Examples: Molec., Ionic, and Net Ionic Eqns.

2. Write ionic and net ionic equations corresponding to the following molecular equations.

$$2AgNO_3(aq) + Na_2CrO_4(aq) \rightarrow Ag_2CrO_4(s) + 2NaNO_3(aq)$$

$$H_2SO_4(aq) + MgCO_3(s) \rightarrow MgSO_4(aq) + H_2O(l) + CO_2(g)$$

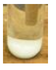
Examples: Predicting Precipitation Reactions

3. Predict whether a precipitation reaction will occur when aqueous solutions of silver nitrate and sodium chloride are mixed. Write the ionic and net ionic equation for the reaction (if any). [movie]

Examples: Predicting Precipitation Reactions

4. Predict whether a reaction occurs when each of the following pairs of solutions are mixed. If a reaction does occur, write balanced molecular, ionic, and net ionic equations, and identify the spectator ions.

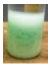
a. __ $CaCl_2(aq)$ + __ $Na_2CO_3(aq) \rightarrow$



b. __ $KMnO_4(aq)$ + __ $NaCl(aq) \rightarrow$

c. __ $Pb(NO_3)_2(aq)$ + __ $Na_2CrO_4(aq) \rightarrow$

d. __ $NiCl_2(aq)$ + __ $Na_3PO_4(aq) \rightarrow$



Examples: Predicting Precipitation Reactions

4. Predict whether a reaction occurs when each of the following pairs of solutions are mixed. If a reaction does occur, write balanced molecular, ionic, and net ionic equations, and identify the spectator ions.

e. __ $Pb(NO_3)_2$ + __ $Fe_2(SO_4)_3$ →

f. __ $Ca(OH)_2$ + __ Na_2SO_4 →

g. Potassium nitrate and ammonium chloride →

h. Iron(III) chloride and cesium phosphate →

Examples: Predicting Precipitation Reactions

5. How might you use a precipitation reaction to prepare a sample of $CuCO_3$? Write the net ionic equation for the reaction.

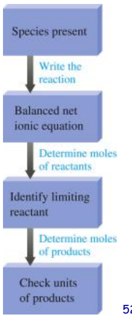
Examples: Stoichiometry of Precipitation Rxns

6. Calculate the mass of solid NaCl that must be added to 1.50 L of 0.100 M $AgNO_3$ to precipitate all of the Ag^+ ions in the form of AgCl(s).

Answer: 8.77 g NaCl

Examples: Stoichiometry of Precipitation Rxns

7. If 15.0 mL of 2.00 M $Pb(NO_3)_2$ and 45.0 mL of 1.15 M NaCl are mixed, how many grams of precipitate will be produced?



Answer: 7.20 g

Neutralization Reactions

Acids and Bases

- The concept of acidity and basicity has existed for a long time, and some compounds have long been designated as acids or bases:
  - **Acids** turn *blue* litmus *red*; their aqueous solutions have a tart taste.
  - **Bases** turn *red* litmus *blue*; their aqueous solutions have a bitter taste and a “soapy” feel.
- The level of acidity is measured on a logarithmic scale called the **pH scale** ($pH = -\log[H^+]$):
  - pH < 7 is **acidic**
  - pH > 7 is **basic**
  - pH = 7 is **neutral**

The Arrhenius Definition of Acids and Bases

- In the **Arrhenius definition** (Svante Arrhenius, 1887) of acids and bases:
  - An **acid** is a substance that produces **H^+ ions** (protons) when dissolved in water.
  - A **base** is a substance that produces **OH^- ions** when dissolved in water.

$$HA(aq) \rightarrow H^+(aq) + A^-(aq)$$

an **acid**

$$MOH(aq) \rightarrow M^+(aq) + OH^-(aq)$$

a **base**

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Brønsted-Lowry Acids and Bases

- The Arrhenius definition of acids and bases applies only to aqueous solutions, but acid-base reactions also take place in other solvents; many species can also act as bases without containing OH^- ions. An extended definition was developed independently in 1923 by Johannes Brønsted and Thomas Lowry.
- The **Brønsted-Lowry definition** of acids and bases is based on **proton (H^+) transfer**:
  - An **Brønsted acid** is a **proton (H^+) donor**.
  - A **Brønsted base** is a **proton (H^+) acceptor**.
  - An acid-base (neutralization) reaction is a **proton transfer process**.
- Any Arrhenius acid or base is also a Brønsted acid or base (but not vice versa).

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Brønsted-Lowry Acids and Bases

$$HCl(g) + H_2O(l) \rightarrow H_3O^+(aq) + Cl^-(aq)$$

$HCl(g)$	$H_2O(l)$
proton donor	*proton acceptor*
Brønsted acid	**Brønsted base**

$$NH_3(aq) + H_2O(l) \rightleftharpoons NH_4^+(aq) + OH^-(aq)$$

$NH_3(aq)$	$H_2O(l)$
proton acceptor	*proton donor*
Brønsted base	**Brønsted acid**

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Lewis Acids and Bases

- The acid-base concept was extended even further by a definition developed by G. N. Lewis in 1923.
- The **Lewis definition** of acids and bases is based on **electron-pair bonding**:
  - A **Lewis base** is an **electron-pair donor**.
  - A **Lewis acid** is an **electron-pair acceptor**.
  - An acid-base reaction is **the donation and acceptance of an electron pair** to form a covalent bond.
- The Lewis definition allows many substances which do not contain "H^+" to be treated as acids.
- *Any Brønsted acid or base is also a Lewis acid or base* (but not necessarily vice versa).

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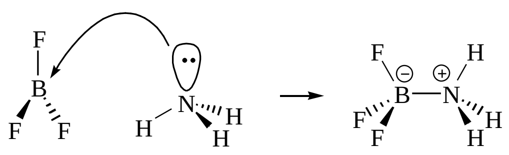
Lewis Acids and Bases

- The key feature of Lewis base is a lone pair of electrons to donate; the key feature of a Lewis acid is a vacant orbital, or the ability to form one, to accept that lone pair and form a new covalent bond.

$$HCl(g) + H_2O(l) \rightarrow H_3O^+(aq) + Cl^-(aq)$$

$HCl(g)$	$H_2O(l)$
e^--pair acceptor	*e^--pair donor*
Lewis acid	**Lewis base**



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Acids

- Acids in their pure form (i.e., not mixed with water) are molecular compounds, but when they are dissolved in water, they *dissociate* into ions:

$$HCl(g) \rightarrow H^+(aq) + Cl^-(aq)$$

- **Polyprotic acids** contain more than one acidic proton. For instance, sulfuric acid is **diprotic**:

$$H_2SO_4(aq) \rightarrow H^+(aq) + HSO_4^-(aq)$$
$$HSO_4^-(aq) \rightleftharpoons H^+(aq) + SO_4^{2-}(aq)$$

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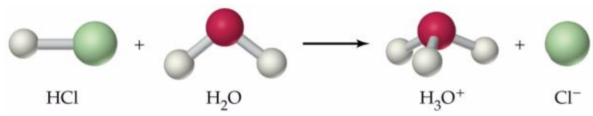
The Hydronium Ion

- An "H⁺" cation is too reactive to exist freely in an aqueous solution. The "H⁺" is always attached to one (or more) water molecules as a **hydronium ion**, $\mathbf{H_3O^+}$ [movie].
- The dissociation reaction for HCl

$$HCl(g) \rightarrow H^+(aq) + Cl^-(aq)$$

is more correctly represented by the reaction:

$$HCl(aq) + H_2O(l) \rightarrow H_3O^+(aq) + Cl^-(aq)$$



Bases

- Under the Arrhenius definition, sodium hydroxide is a base because it produces hydroxide ions in water:

$$NaOH(s) \rightarrow Na^+(aq) + OH^-(aq)$$

- Hydroxide salts in their pure form are ionic solids, but when they are dissolved in water, the cation and anion dissociate from each other, releasing hydroxide into the solution:

$$KOH(s) \rightarrow K^+(aq) + OH^-(aq)$$

$$Ba(OH)_2(s) \rightarrow Ba^{2+}(aq) + 2OH^-(aq)$$

Strong Acids and Weak Acids

- Acids that dissociate to a large extent are strong electrolytes, and are therefore called **strong acids**. There are six major strong acids:

Hydrochloric acid, HCl	Sulfuric acid, H_2SO_4
Hydrobromic acid, HBr	Nitric acid, HNO_3
Hydriodic acid, HI	Perchloric acid, $HClO_4$

memorize!

- Acids that dissociate to only a small extent are weak electrolytes, and are therefore called **weak acids**. Examples of weak acids include:

  Hydrofluoric acid, HF
  Acetic acid, $HC_2H_3O_2$ (or CH_3CO_2H)
  Phosphoric acid, H_3PO_4
  etc.

Strong Bases and Weak Bases

- Bases can also be either strong or weak.
- Most metal hydroxides are **strong bases**, which are virtually completely dissociated in water. Some examples include:

  Sodium hydroxide, NaOH
  Potassium hydroxide, KOH
  Calcium hydroxide, $Ca(OH)_2$
  Barium hydroxide, $Ba(OH)_2$

- **Weak bases** are only partially dissociated. Weak bases do not contain OH^- ions, but produce them in an equilibrium reaction with water:

$$NH_3(aq) + H_2O(l) \rightleftharpoons NH_4^+(aq) + OH^-(aq)$$

Neutralization Reactions

- **Acid-base reactions** or **neutralization reactions** occur when an acid and base react to form *water*, a *weak electrolyte*, or a *gas*. (These reactions are also a form of double-displacement reaction.)

$$\underset{\text{acid}}{HA(aq)} + \underset{\text{base}}{MOH(aq)} \rightarrow \underset{\text{a salt}}{MA(aq)} + \underset{\text{water}}{H_2O(l)}$$

- For a strong acid and base, this can be written as the ionic equation:

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

- Canceling the spectator ions gives the net ionic equation:

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

$$\{ H_3O^+(aq) + OH^-(aq) \rightarrow 2H_2O(l) \}$$

Examples: Neutralization Reactions

1. Write the molecular equation, complete ionic equation, and net ionic equation for the following neutralization reactions:

$HCl(aq) + NaOH(aq) \rightarrow$

$HI(aq) + Ba(OH)_2(aq) \rightarrow$

Neutralization Reactions

- We will see the same net ionic equation for the reaction of *any* strong acid with *any* strong base.
- Neutralization reactions are essentially a type of double-displacement (metathesis) reaction where one of the products is water.
- When one of the species is a weak acid or base, one of the products is still water, but the net ionic equation will be slightly different, since weak acids and bases do not dissociate:

$$HF(aq) + NaOH(aq) \rightarrow NaF(aq) + H_2O(l)$$

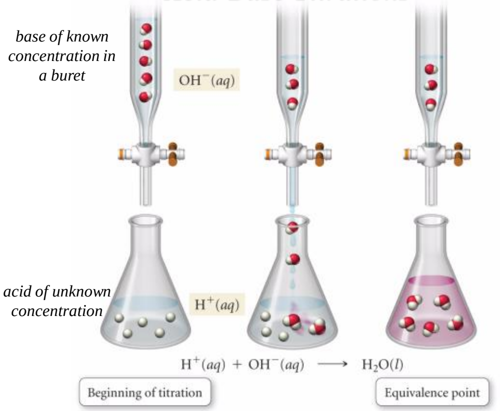
NIE: $HF + OH^- \rightarrow F^- + H_2O(l)$

Acid-Base Titrations

- A **titration** is a procedure for determining the concentration of a solution:
  - A *standard solution* of known concentration is reacted with a solution of unknown concentration.
  - By measuring the volume of standard solution that reacts with a known volume of the unknown solution, the concentration can be calculated from the reaction stoichiometry.
- A common example of this process is an **acid-base titration**, in which an acid or base of unknown concentration reacts with a base or acid of known concentration:

$$HCl(aq) + NaOH(aq) \rightarrow NaCl(aq) + H_2O(l)$$

Acid-Base Titrations



Acid-Base Titrations

- In the previous figure, OH- from the buret is added to the acid of unknown concentration in the flask until the **equivalence point** is reached, when the number of moles of OH^- added equals the number of moles of H^+ that were originally present.

M, L NaOH → **mol NaOH** → **mol HCl** → **M HCl**

$\text{mol} = \frac{\text{mol}}{\text{L}} \times \text{L}$ $\frac{1\,\text{mol HCl}}{1\,\text{mol NaOH}}$ $\text{M} = \frac{\text{mol}}{\text{L}}$

sim. to Figure 10.18. The equivalence point is usually signaled by the color change of an acid-base **indicator**, such as phenolphthalein, which is colorless in an acidic solution but pink in a basic solution.

Examples: Acid-Base Titrations

2. A 25.0 mL sample of vinegar (dilute acetic acid, $HC_2H_3O_2$) is titrated and found to react with 94.7 mL of 0.200 M NaOH. What is the molarity of the acetic acid solution?

$$HC_2H_3O_2(aq) + NaOH(aq) \rightarrow NaC_2H_3O_2(aq) + H_2O(l)$$

Answer: 0.758 M $HC_2H_3O_2$

Other Neutralization Reactions

- Neutralization reactions may produce other products besides water:
  - Some acid-base reactions result in the **formation of gases**, either directly or in the decomposition of an unstable intermediate. This removes ions from the solution, driving the reaction to occur.

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

  - Neutralization reactions can also result in the **formation of weak electrolytes**, which also do not dissociate. Usually, these will be weak acids:

$$HCl(aq) + NaF(aq) \rightarrow HF(aq) + NaCl(aq)$$

Neutralization Reactions that Produce Gases

- Carbonic acid, H_2CO_3, is unstable, and falls apart to give water and $CO_2(g)$.

$2HCl(aq) + K_2CO_3(aq) \rightarrow 2KCl(aq) + \mathbf{H_2CO_3(aq)}$

$2HCl(aq) + K_2CO_3(aq) \rightarrow 2KCl(aq) + \mathbf{H_2O(l) + CO_2(g)}$

- Other gases that can be produced in neutralization reactions include $\mathbf{H_2S(g)}$, **HCN(g)**, $\mathbf{SO_2(g)}$ (from the decomposition of H_2SO_3), and $\mathbf{NH_3(g)}$ (from the dissociation of NH_4OH).

$HCl(aq) + NaCN(aq) \rightarrow NaCl(aq) + \mathbf{HCN(g)}$

$HCl(aq) + Na_2SO_3(aq) \rightarrow 2NaCl(aq) + \mathbf{H_2SO_3(aq)}$
$HCl(aq) + Na_2SO_3(aq) \rightarrow 2NaCl(aq) + \mathbf{H_2O(l) + SO_2(g)}$

$NH_4Cl(aq) + KOH(aq) \rightarrow KCl(aq) + \mathbf{NH_4OH(aq)}$
$NH_4Cl(aq) + KOH(aq) \rightarrow KCl(aq) + \mathbf{NH_3(g) + H_2O(l)}$

Double-Displacement Reactions

- Both precipitation and acid-base reactions are examples of **double-displacement** or **metathesis** reactions. In these reactions, the cations and anions in two ionic compounds "change partners":

$$AB + CD \rightarrow AD + CB$$

- In order for the reaction to occur, one of four things must occur, otherwise, there is no reaction (NR).

1. In **precipitation reactions**, an ***insoluble ionic compound*** forms (solubility rules).

2. In **acid-base** (**neutralization**) **reactions**, an acid and a base form a salt and ***water***, a ***gas***, or a ***weak electrolyte***.

Examples: Predicting Neutralization Rxns

3. Predict whether a reaction occurs when each of the following pairs of solutions are mixed. If a reaction does occur, write balanced molecular, ionic, and net ionic equations, and identify the spectator ions.

a. HCl and calcium carbonate [movie]

b. Hydrobromic acid and barium hydroxide

c. Nitric acid and sodium fluoride

d. Sulfuric acid and sodium cyanide

e. Sodium bicarbonate and acetic acid

f. CdS(aq) + HCl

Oxidation-Reduction (Redox) Reactions

Oxidation and Reduction

- **Oxidation-Reduction (redox) reactions**, are another major class of chemical reactions.
  - Historically, "oxidation" meant the combination of an element with oxygen, as in the rusting of iron:

$$4Fe(s) + 3O_2(g) \rightarrow 2Fe_2O_3(s)$$

  - "Reduction" meant the removal of oxygen from an oxide, as in the heating of iron ore with charcoal in a blast furnace to form iron metal:

$$2Fe_2O_3(s) + 3C(s) \rightarrow 4Fe(s) + 3CO_2(g)$$

- Now, the terms "oxidation" and "reduction" have more specific meanings:

Oxidation-Reduction Reactions

- An **oxidation-reduction** (**redox**) **reaction** is a process in which *electrons are transferred* from one substance to another.
  - **Oxidation** is the **loss** of electrons.
  - **Reduction** is the **gain** of electrons.

Oxidation: $Fe \rightarrow Fe^{2+} + 2e^-$

Reduction: $Cu^{2+} + 2e^- \rightarrow Cu$ *half-reactions*

oxidation

$$Fe(s) + Cu^{2+}(aq) \longrightarrow Fe^{2+}(aq) + Cu(s)$$

reduction

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Oxidizing and Reducing Agents

- The **oxidizing agent** is the substance that causes oxidation to occur by accepting electrons. The oxidizing agent itself becomes *reduced*.
- The **reducing agent** is the substance that causes reduction to occur by losing electrons. The reducing agent itself becomes *oxidized*.
- In general, metals tend to act as reducing agents (lose electrons), and nonmetals tend to act as oxidizing agents (gain electrons).

oxidation

$$Fe(s) + Cu^{2+}(aq) \longrightarrow Fe^{2+}(aq) + Cu(s)$$

reduction

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Oxidation Numbers

- In the following reaction, it is easy to see that a redox reaction is taking place, because the charges are explicitly written out:

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

- But what about these reactions?

$$2Mg(s) + O_2(g) \rightarrow 2MgO(s)$$

$$H_2(g) + Cl_2(g) \rightarrow 2HCl(g)$$

$$MnO_4^-(aq) + 5Fe^{2+}(aq) + 8H^+(aq) \rightarrow Mn^{2+}(aq) + 5Fe^{3+}(aq) + 4H_2O$$

- We can keep track of whether a reaction is a redox reaction, and what specifically is being oxidized or reduced, by using **oxidation numbers** (aka *oxidation state*). Note that ***oxidation number is not the same thing as charge***!!!!!

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Rules for Assigning Oxidation Numbers

These rules are hierarchical. If any two rules conflict, follow the rule that is higher on the list.

1. The oxidation number of an atom in its elemental state is zero.

Na	H_2	Br_2	S	Ne
0	0	0	0	0

2. The oxidation number of a monatomic ion is the same as its charge.

Na^+	Ca^{2+}	Al^{3+}	Cl^-	O^{2-}	N^{3-}
+1	+2	+3	-1	-2	-3

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Rules for Assigning Oxidation Numbers

3. The sum of the oxidation numbers is 0 for a neutral compound and is equal to the net charge for a polyatomic ion.

H_2SO_4: +1 ? -2

ClO_4^-: ? -2

ClO_2^-: ? -2

NH_4^+: ? +1

4. In their compounds, metals have positive oxidation numbers:

a. Group 1A metals *always* have oxidation numbers of +1.

b. Group 2A metals *always* have oxidation numbers of +2.

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Rules for Assigning Oxidation Numbers

5. In their compounds, nonmetals are assigned oxidation states according to the table below. Entries at the top of the table take precedence over entries at the bottom of the table.

a. Fluorine -1

b. Hydrogen +1 (-1 when bonded with a metal)

NaH: +1 -1

CaH_2: +2 -1

HCl: +1 -1

H_2S: +1 -2

c. Oxygen -2 (-1 in peroxides, O_2^{2-})

H_2O: +1 -2

OH^-: -2 +1

H_2O_2: +1 ?

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Rules for Assigning Oxidation Numbers

d. Group 7A -1 (except in compounds with O)

NaCl: +1, ?
Cl_2O: ?, -2
HOBr: +1, -2, ?

e. Group 6A -2 (except in compounds with O)

Na_2S: +1, ?
SO_2: ?, -2
SO_4^{2-}: ?, -2

f. Group 5A -3 (except in compounds with O)

NH_3: ?, +1
NO_2: ?, -2
NO_3^-: ?, -2

Examples: Assigning Oxidation Numbers

1. Determine the oxidation number of each element in the following species.

a. Zinc chloride
b. Sulfur trioxide
c. Sulfite ion
d. Chlorite ion
e. Nitric acid
f. $S_2O_3^{2-}$
g. Sodium azide (NaN_3)
h. Potassium dichromate

Identifying Redox Reactions

- Another way to define a redox reaction is one in which the oxidation numbers of species change:
  - **oxidation** is an **increase** in oxidation number.
  - **reduction** is a **decrease** in oxidation number.
- Any reaction in which oxidation numbers change is a redox reaction. If no oxidation numbers change, the reaction is *not* a redox reaction.

$$S + O_2 \rightarrow SO_2$$

S: 0; O: 0; SO_2: +4, -2

S: 0 → +4 *oxidized*
O: 0 → -2 *reduced*

Examples: Identifying Redox Reactions

2. Identify the substance oxidized, the substance reduced, the oxidizing agent, and the reducing agent in each of the following reactions.

a. $2Al(s) + 3H_2SO_4(aq) \rightarrow Al_2(SO_4)_3(aq) + 3H_2(g)$

b. $PbO(s) + CO(g) \rightarrow Pb(s) + CO_2(g)$

c. $H_2(g) + Cl_2(g) \rightarrow 2HCl(g)$

d. $CH_4(g) + 2O_2(g) \rightarrow CO_2(g) + 2H_2O(g)$

e. $2KCl + MnO_2 + 2H_2SO_4 \rightarrow K_2SO_4 + MnSO_4 + Cl_2 + 2H_2O$

f. $Na_2SO_4(aq) + Pb(NO_3)_2(aq) \rightarrow 2NaNO_3(aq) + PbSO_4(s)$

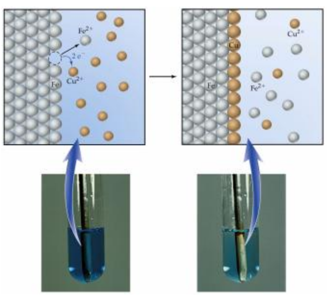
Single-Displacement Reactions; Activity Series

- A particular kind of redox reaction, called a **single-displacement reaction** occurs when an uncharged metal (A^0) "displaces" a charged metal (B^+) in an ionic compound (B^+C^-):

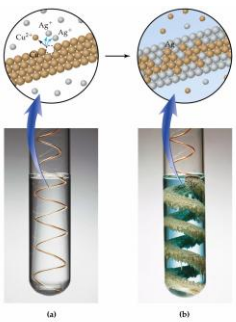
$$A^0 + B^+C^- \rightarrow A^+C^- + B^0$$

  - The reacting metal (A^0) becomes charged (A^+), forming a new ionic compound (A^+C^-), and the displaced metal (B^+) becomes uncharged (B^0):
- This reaction only occurs when the metal A^0 is a *more active* (i.e., higher on the **activity series**) metal than the metal cation B^+. The products of the reaction are written using ionic charge rules.

A Single-Displacement Reaction



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



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

The Activity Series of the Metals

Strength as a reducing agent	Metal	Behavior toward H^+
Most active ↑	Li	
	K	
	Ba	
	Sr	
	Ca	
	Na	↑ displace H^+ from water
	Mg	
	Al	
	Mn	
	Zn	
	Cr	
	Fe	
	Cd	↑↑ displace H^+ from steam
	Co	
	Ni	
	Sn	
	Pb	↑↑↑ displace H^+ from acid
	H_2	
	Cu	↓ do not displace H^+
	Ag	
	Pt	
	Hg	
Least active	Au	

Examples: Single-Displacement Reactions

3. Predict whether the following reactions will occur.

a. $Hg^{2+}(aq) + Zn(s) \rightarrow Hg(l) + Zn^{2+}(aq)$

b. $2H^+(aq) + 2Ag(s) \rightarrow H_2(g) + 2Ag^+(aq)$

c. $CaCl_2(aq) + Mg(s) \rightarrow Ca(s) + MgCl_2(aq)$

d. $Ca(s) + MgCl_2(aq) \rightarrow CaCl_2(aq) + Mg(s)$

e. $Zn(s) + SnCl_2(aq) \rightarrow ZnCl_2(aq) + Sn(s)$ [movie]

Examples: Single-Displacement Reactions

4. Predict the products (if any) of the following reactions.

a. $Al(s) + Pb(NO_3)_2(aq) \rightarrow$

b. $Ni(s) + Zn(NO_3)_2(aq) \rightarrow$

c. $Al(s) + CuCl_2(aq) \rightarrow$

d. $AlCl_3(aq) + Cu(s) \rightarrow$

Examples: Single-Displacement Reactions

5. Predict the products (if any) of the following reactions.

a. $Zn(s) + Cu^{2+}(aq) \rightarrow$

b. $Al(s) + H_2O(g) \rightarrow$

c. $K(s) + H_2O(l) \rightarrow$

d. $Au(s) + H_2O(l) \rightarrow$

Combustion Reactions

- **Combustion reactions** are a type of redox reaction in which a substance reacts with molecular oxygen (O_2) to form one or more oxygen-containing products. These reactions are usually accompanied by the release of large amounts of heat and light (*burning*).

$2Mg(s) + O_2(g) \rightarrow 2MgO(s)$ [movie]

$4Fe(s) + 3O_2(g) \rightarrow 2Fe_2O_3(s)$

$S(s) + O_2(g) \rightarrow SO_2(s)$ [or $SO_3(g)$]

$2H_2(g) + O_2(g) \rightarrow 2H_2O(g)$ [movie]

$C(s) + O_2(g) \rightarrow CO_2(g)$ [burning of coal]

Combustion Reactions

- The primary type of combustion reaction is the **combustion of hydrocarbons** to produce carbon dioxide and water (as well as heat energy, which can be converted to other forms of energy):

Hydrocarbon + $O_2(g) \rightarrow CO_2(g) + H_2O(g)$
(C,H,O)

$CH_4(g) + 2O_2(g) \rightarrow CO_2(g) + 2H_2O(g)$
methane

$C_3H_8(g) + 5O_2(g) \rightarrow 3CO_2(g) + 4H_2O(g)$

$2C_8H_{18}(l) + 25O_2(g) \rightarrow 16CO_2(g) + 18H_2O(g)$
~ gasoline

$C_2H_5OH(l) + 3O_2(g) \rightarrow 2CO_2(g) + 3H_2O(g)$
ethanol

Combustion Reactions

- In a limited oxygen supply, carbon monoxide or elemental carbon may form instead; hot elemental carbon glows with an orange color, producing the colors associated with burning wood or candles.
- The metabolism of glucose ($C_6H_{12}O_6$) in the body is similar to combustion reactions:

  $C_6H_{12}O_6 + 6O_2 \rightarrow 6CO_2 + 6H_2O$ + energy

  In living cells the process occurs slowly, in a complex series of steps that release energy in a controlled way.

Other Applications of Redox Reactions

- **Bleaching** uses redox reactions to decolorize or lighten colored materials (hair, clothes, paper, etc.)
- **Batteries** use spontaneous redox reactions which occur in separated compartments; the electrons that move between the oxidation compartment (anode) and the reduction compartment (cathode) can be used to deliver electrical energy .
- **Metallurgy** is the processes involved in extracting and purifying metals from their ores. Roasting ores with carbon (coke) carries away oxygen in the form of CO or CO_2, yielding the pure metal:

  $ZnO(s) + C(s) \rightarrow Zn(s) + CO(g)$
- **Corrosion** is the deterioration of a metal by oxidation, as in the rusting of iron in moist air.

  $4Fe(s) + 3O_2(g) + H_2O \rightarrow 2Fe_2O_3{\bullet}H_2O(s)$

Synthesis and Decomposition Reactions

Synthesis Reactions

- In a **synthesis** or **combination reaction**, two or more reactants combine to form one product.

$$A + B \rightarrow C$$

Combination Reactions that *Are* Redox Reactions

1. Metal + nonmetal → ionic compound

$Zn(s) + S(s) \rightarrow ZnS(s)$

$2Al(s) + 3Cl_2(g) \rightarrow 2AlCl_3(s)$

Synthesis Reactions

Combination Reactions that *Are* Redox Reactions

2. Nonmetal + nonmetal → molecular compound

$S(s) + O_2(g) \rightarrow SO_2(g)$

$2S(s) + 3O_2(g) \rightarrow 2SO_3(g)$

3. Compound + element → molecular compound

$2NO(g) + O_2(g) \rightarrow 2NO_2(g)$

$PCl_3(l) + Cl_2(g) \rightarrow PCl_5(s)$

Synthesis Reactions

Combination Reactions that *Are Not* Redox Rxns

1. Metal oxide + nonmetal oxide → ionic compound w/polyatomic ion

$Li_2O(s) + CO_2(g) \rightarrow Li_2CO_3(s)$

2. Metal oxide + water → metal hydroxides (bases)

Metal oxides are also known as *basic anhydrides*.

$CaO(s) + H_2O(l) \rightarrow Ca(OH)_2(s)$

$Na_2O(s) + H_2O \rightarrow 2NaOH(aq)$

$Al_2O_3(s) + 3H_2O \rightarrow 2Al(OH)_3(aq)$

Synthesis Reactions

Combination Reactions that *Are Not* Redox Rxns

3. Nonmetal oxide + water → acid

Add the elements of water into the nonmetal oxide formula and reduce the subscripts to lowest terms. For phosphorus oxides, add three molecules of water.

Nonmetal oxides are also known as *acidic anhydrides*.

$$CO_2(g) + H_2O \rightarrow H_2CO_3(aq)$$
$$N_2O_5(g) + H_2O \rightarrow 2HNO_3(aq)$$
$$P_2O_5(s) + 3H_2O \rightarrow 2H_3PO_4(aq)$$

Decomposition Reactions

- In a **decomposition reaction**, one reactant breaks down to form two or more products.

$$C \rightarrow A + B$$

Thermal Decomposition

1. Ionic compound with oxoanion →

metal oxide + gaseous nonmetal oxide

$$CaCO_3(s) \xrightarrow{\Delta} CaO(s) + CO_2(g)$$

2. Many metal oxides, chlorates, and perchlorates release O_2 when heated.

$$2KClO_3(s) \xrightarrow{\Delta} 2KCl(s) + 3O_2(g)$$
$$2NaNO_3(s) \xrightarrow{\Delta} 2NaNO_2(s) + O_2(g)$$

Decomposition Reactions

Electrolytic Decomposition — occurs when compounds absorb electrical energy and undergo *electrolysis*.

1. Decomposition of water.

$$2H_2O(l) \xrightarrow{\text{electricity}} 2H_2(g) + O_2(g)$$

2. Molten ionic compounds decompose into their elements.

$$MgCl_2(l) \xrightarrow{\text{electricity}} Mg(l) + Cl_2(g)$$
$$2NaCl(l) \xrightarrow{\text{electricity}} 2Na(l) + Cl_2(g)$$

Decomposition Reactions

Decomposition of Binary Compounds — may occur by thermal decomposition, electrolysis, or other conditions.

1. Binary compound → element + element

$$2NI_3(s) \xrightarrow{\text{tickle}} N_2(g) + 3I_2(g) \quad (+\text{bang!})$$

- The decomposition of binary compounds are the only decomposition reactions for which you'll be asked to predict the products. Be able to recognize and balance the other types of decompositions.

Examples: Predicting Chemical Reactions

1. Complete the following reactions and classify them as synthesis, decomposition, single-displacement, acid-base, precipitation, or combustion reactions.

a. Magnesium + nitrogen →

b. $H_2SO_4(aq) + NaF(aq) \rightarrow$

c. $Na_2S(aq) + HCl(aq) \rightarrow$

d. Hydrogen peroxide(l) →

e. $(NH_4)_2SO_4(aq) + K_3PO_4(aq) \rightarrow$

f. Cadmium + lead(II) nitrate →

The End