Chapter 13: Chemical Reactions

Homework: All questions on the “Multiple-Choice” and the odd-numbered questions on “Exercises” sections at the end of the chapter.
Properties

- **Properties** – the characteristics of a substance
  - Physical Properties – do not describe the chemical reactivity of the substance
    - Density, hardness, phase, color, melting point, electrical conductivity, specific heat
  - Chemical Properties – reflect the ways in which a substance can be transformed into another substance
    - Describe a substance’s chemical reactivity
    - Burning, rusting, decomposition
Physical and Chemical Properties and Changes

<table>
<thead>
<tr>
<th>Physical</th>
<th>Chemical</th>
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</thead>
<tbody>
<tr>
<td>Property</td>
<td>Description that tells how a substance reacts, or fails to react, chemically</td>
</tr>
<tr>
<td>Change</td>
<td>Description such as size, color, odor, density, and melting point</td>
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<td></td>
<td>Change in which no new substance is formed, only a different form of the original substance</td>
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Chemical Reactions

- **Chemical Reaction** – a change that alters the chemical composition of a substance and results in the formation of one or more new substances
  - Decomposition of water $H_2O$ into hydrogen ($H_2$) and oxygen ($O_2$) gases

- A chemical reaction is simply a rearrangement of the atoms. Some of the original chemical bonds are broken and new bonds form.
  - New and different chemical structures are formed.
Chemical Reaction – Rearrangement of Atoms
Chemical Reactions

• Generally, the atom’s valence electrons are the only ones directly involved in the chemical reaction.
  – *The nucleus is unchanged and therefore the identity of the atom is unaffected.*

• Consider the following generalized reaction

  • $A + B \rightarrow C + D$

  • **Reactants** – original substances $A + B$
  • **Products** – the new substances $C + D$
  • “$\rightarrow$” means “*reacts to form*” or “*yields*”
## Common Symbols in Chemical Equations

<table>
<thead>
<tr>
<th>Symbol</th>
<th>Meaning</th>
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<tbody>
<tr>
<td>+</td>
<td>Plus, or and</td>
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<td>→</td>
<td>React to form, or yields</td>
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<td>(g)</td>
<td>Gas</td>
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<td>(l)</td>
<td>Liquid</td>
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<td>(s)</td>
<td>Solid</td>
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<tr>
<td>(aq)</td>
<td>Aqueous (water) solution</td>
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<tr>
<td>MnO₂ →</td>
<td>Catalyst (MnO₂, in this case)</td>
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<td>⇌</td>
<td>Equilibrium</td>
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Chemical Reactions

• In any chemical reaction three things occur:
  1) Reactants disappear or are diminished.
  2) New substances are formed as products.
     • These products have different chemical/physical properties from the original reactants.
  3) Energy is either released or absorbed.
     • Heat, light, electricity, sound
Chemical Equations

• A chemical equation can be written for every chemical reaction.
• The correct chemical formulas must be used and the equation must be balanced.
  – A balanced equation reflects the correct chemical formulas and the correct ratios of reactants and products.
Balancing Chemical Reactions

- Most chemical reactions can be balanced by trial and error, using three simple principles:
  1) An equal number of atoms of each kind must be represented on each side of the reaction arrow.
  2) The formulas may not be changed, only the coefficients in front of the formulas.
  3) The final set of coefficients used should be the smallest whole numbers that will satisfy the equation.
Balancing Chemical Reactions – an Example

• HI $\rightarrow$ H$_2$ + I$_2$  -- *an unbalanced equation*
• The formulas cannot be changed but a coefficient of “2” can be used, as below.
• 2 HI $\rightarrow$ H$_2$ + I$_2$  -- *a balance equation*
• The following equations are also inappropriate:
  – HI $\rightarrow$ $\frac{1}{2}$ H$_2$ + $\frac{1}{2}$ I$_2$  *Fractions should not be used*
  – 4 HI $\rightarrow$ 2H$_2$ + 2I$_2$  *Smallest whole number should be used*
Tips to Balance Equations

• You must be able to properly count the atoms.
  – \(4\text{Al}_2(\text{SO}_4)_3\) -- in this equation there are 8 Al atoms, 12 S atoms, and 48 O atoms

• Start by balancing an element that is present in only one place on both sides of the reaction.
  – In the reaction \(C + \text{SO}_2 \rightarrow \text{CS}_2 + \text{CO}\) start by balancing the S or the O.

• Insert the lowest coefficient possible on either side to get the same number of atoms of that element on each side.
Tips to Balance Equations

• When polyatomic ions remain intact during the reaction, balance them as a unit.
• If both sides balance only by the use of a fractional coefficient in one place, multiply all the coefficients by the denominator of the fraction.

  $C_2H_2 + \frac{5}{2} O_2 \rightarrow 2 CO_2 + H_2O$
  
  Multiply by 2 (the denominator of 5/2)
  $2 C_2H_2 + 5 O_2 \rightarrow 4 CO_2 + 2 H_2O$
  
  Resulting in a balance equation with no fractions
Balancing Equations – An Example

• Mg + O₂ → MgO – an unbalanced equation
• Balance the oxygens first.
• Mg + O₂ → 2 MgO
  – Oxygens are balanced but not magnesiums.
• 2 Mg + O₂ → 2 MgO
  – Now both magnesiums and oxygens are balanced.
Balancing Equations

Confidence Exercise

- $\text{NaN}_3(s) \rightarrow \text{Na}(s) + \text{N}_2(g)$
  - *This is an unbalance reaction.*
- $2 \text{NaN}_3(s) \rightarrow \text{Na}(s) + 3 \text{N}_2(g)$
  - *Nitrogens are now balanced.* (6 on each side)
- $2 \text{NaN}_3(s) \rightarrow 2 \text{Na}(s) + 3 \text{N}_2(g)$
  - *Both nitrogens and sodiums are balanced.*

This reaction shows how car air bags inflate by the electrical ignition of sodium azide ($\text{NaN}_3$) to produce nitrogen gas ($\text{N}_2$)
Combination Reactions

- **Combination Reaction** – at least two reactants combine to form one product
  - \( A + B \rightarrow AB \)
Decomposition Reactions

• **Decomposition Reaction** – only one reactant is present and it breaks into two, or more, products
  \[ AB \rightarrow A + B \]
Chemical Reactions – *A Change in Energy*

- During a chemical reaction, some chemical bonds are broken and other bonds are formed.
- The different bonding energies of the reactants and products result in a change in energy.
- In order to break bonds, energy must be absorbed.
- When new bonds are formed, energy is released.
- Energy from a chemical reaction is released or absorbed in the form of light, heat, electrical energy, or sound.
Exothermic Reactions

- **Exothermic reaction** – a chemical reaction that results in a net release of energy to the surroundings.

  - **Example:** The burning of methane

  \[
  \text{CH}_4 + 2 \text{O}_2 \rightarrow \text{CO}_2 + 2 \text{H}_2\text{O} + \text{energy}
  \]

  - Bonding energy in the products is less than bonding energy in the reactants – *therefore, energy is released.*
Exothermic Reaction

- Net release of energy \((E_R)\) - the bonds in the products have less total energy than the bonds in the reactants

\[
\text{Reactants} \\
\begin{align*}
\text{CH}_4 + 2 \text{O}_2 &\rightarrow \text{CO}_2 + 2 \text{H}_2\text{O} \\
E_{\text{act}} &= \text{activation energy} \\
E_R &= \text{energy of reaction (energy liberated in this case)}
\end{align*}
\]
Endothermic Reactions

• **Endothermic reaction** - a chemical reaction that results in a net absorption of energy from the surroundings

• *Example:* \( 3 \text{O}_2 + \text{energy} \rightarrow 2 \text{O}_3 \)

• Bonding energy in the reactants is less than bonding energy in the products.
  – *Although energy is released when the ozone bonds are formed, the amount is less than is absorbed in breaking the oxygen molecule bonds*
  – *therefore energy is absorbed.*
Endothermic Reaction

• *Net absorption of energy* \((E_R)\) - *the bonds in the reactants have less total energy than the bonds in the products*
Activation Energy

- **Activation energy** \( (E_{\text{act}}) \) – the energy necessary to start a chemical reaction
- In order to burn methane, one must provide an initial spark to break the “first” C–H and O–O bonds.
  - After the initial bonds are broken, the energy released breaks the bonds of still more \( \text{CH}_4 \) and \( \text{O}_2 \), continuously giving off energy as heat and light.
- \( E_{\text{act}} \) – the minimum kinetic energy at colliding molecules must possess to react chemically
The activation energy required for a common match is acquired through friction (heat.)
Heat Flow in Exothermic and Endothermic Reactions

Exothermic – vessel heats up as heat flows to surroundings

Endothermic – vessel cools off as heat flows in from surroundings
Explosive & Combustive Reactions

- **Explosion** – occurs when an exothermic chemical reaction liberates energy almost instantaneously

- **Combustion reaction** – a substance reacts with oxygen by bursting into flames and forming an oxide (exothermic reaction)
  - *Burning of natural gas, coal, paper, wood*
Hydrocarbon Combustion

• All C-H compounds (hydrocarbons) and C-H-O compounds produce energy when they react with oxygen.
  – Exothermic chemical reactions
  – Give off $CO_2$ and $H_2O$ when combustion is complete
Complete Hydrocarbon Combustion

*An Example*

- **One of the components of gasoline is the hydrocarbon named heptane, \( \text{C}_7\text{H}_{16} \). Write the balanced equation for its complete combustion.**

- Write heptane plus oxygen w/ reaction arrow:
  \[
  \text{C}_7\text{H}_{16} + \text{O}_2 \rightarrow \text{CO}_2 + \text{H}_2\text{O}
  \]

- The products are always \( \text{CO}_2 \) & \( \text{H}_2\text{O} \).

- Balance equation:
  \[
  \text{C}_7\text{H}_{16} + 11 \text{O}_2 \rightarrow 7 \text{CO}_2 + 8 \text{H}_2\text{O}
  \]
Complete Hydrocarbon Combustion
Confidence Exercise

• Write and balance the equation for the complete combustion of the hydrocarbon propane, \( C_3H_8 \), a common fuel gas.
• Write propane plus oxygen w/ reaction arrow
  \[ C_3H_8 + O_2 \rightarrow \]
• The products are always \( CO_2 \) & \( H_2O \)
• \( C_3H_8 + O_2 \rightarrow CO_2 + H_2O \)
• Balance equation
• \( C_3H_8 + 5 O_2 \rightarrow 3 CO_2 + 4 H_2O \)
Incomplete Combustion of Heptane

• With insufficient time or oxygen heptane does not completely combust, resulting in the following chemical reaction:

\[
C_7H_{16} + 9 \text{O}_2 \rightarrow 4 \text{CO}_2 + 2 \text{CO} + \text{C} + 8 \text{H}_2\text{O}
\]

• In many cases this incomplete combustion can be seen in the dark exhaust gases given off by some automobiles.

  – *Note that sooty black carbon (C) and poisonous carbon monoxide (CO) are also products of incomplete hydrocarbon combustion.*
Rate of Reaction

• How fast a reaction proceeds (rate of reaction) depends on four variables:
  1) Temperature of the reactants
  2) Concentration of the reactants
  3) Surface Area of the reactants
  4) Catalyst Presence
Temperature affects Rate of Reaction

- Recall that temperature is the average kinetic energy of the molecules.
- For chemical reactions to proceed, the reacting molecules must collide with enough kinetic energy to break bonds. \(E_{\text{act}}\)
- Increased kinetic energy (higher temperature) of the reacting molecules results in more abundant and harder molecular collisions, therefore, the rate of reaction increases.
Concentration affects Rate of Reaction

• Generally, if the concentration of the reactants is greater, the rate of the reaction is also greater.

• A higher concentration of reactants results in more molecular collisions and a faster reaction rate.

• In pure oxygen (100%) a normal cigarette will actually burst into flames.
  – Recall that typical air has only about 21% oxygen.
Concentration affects Rate of Reaction

Phosphorus burning in pure oxygen (left) vs. 21% oxygen (right)
Surface Area affects Rate of Reaction

• As the surface area of reactants increase, so does the reaction rates.
• For example, chemical weathering of rocks progresses much faster as the surface area of the rocks increase.
• Dramatic and devastating explosions have occurred when large amounts of coal dust or grain dust are present.
  – *It is the high surface area of these dusts that causes combustion with explosive speed.*
Catalysts affects the Rate of Reaction

- **Catalyst** – a substance that increases the rate of reaction, but is not consumed in the reaction
- Some catalysts provide a surface to aid in concentrating the reactants.
- Most catalysts provide a new reaction pathway that has a lower activation energy.
- Although not consumed, the catalyst usually forms an intermediate “product” that takes part in the process and then decomposes back to its original form.
Catalysts

- Generally provides a new reaction pathway with a lower activation energy requirement. The presence of NO lowers the activation energy requirement ($E_{\text{cat}}$)
Catalysts affects the Rate of Reaction

An Example

• In the manufacture of H₂SO₄ the reaction is very slow without a catalyst.
  
• 2 SO₂ + O₂ $\rightarrow$ 2 SO₃ (slow)

• When NO is added two fast reactions occur.
  
• 2 NO + O₂ $\rightarrow$ 2 NO₂ (fast)

• 2 SO₂ + 2 NO₂ $\rightarrow$ 2 SO₃ + 2 NO (fast)

• Therefore the following is the net reaction:
  
• 2 SO₂ + O₂ $\rightarrow$ 2 SO₃ net reaction (fast)
Catalysts affects the Rate of Reaction

Another Example

• The decomposition of $\text{H}_2\text{O}_2$ at room temperature is very slow.

• $2 \text{H}_2\text{O}_2 \rightarrow 2 \text{H}_2\text{O} + \text{O}_2 \ (\text{slow})$

• When a small amount $\text{MnO}_2$ is added to the $\text{H}_2\text{O}_2$ the reaction proceeds rapidly.

• $2 \text{H}_2\text{O}_2 \xrightarrow{\text{MnO}_2} 2 \text{H}_2\text{O} + \text{O}_2$
  
  – Catalyst is signified by placing it over the reaction arrow.
Automotive Catalytic Converter

- Beads of Pt, Rh, or Pd serve as catalysts to quickly convert noxious CO and NO into CO$_2$ and N$_2$. 
Enzymes

- In the biological world, living organisms also use catalysts. They are called enzymes.
- These enzymes control various physiologic reactions.
- For example, milk sugar (lactose) is broken down in a reaction catalyzed by the enzyme lactase.
- Individuals who are “lactose intolerant” have a deficiency of lactase.
Bombardier Beetle

*This exothermic reaction uses an enzyme catalyst.*
Acids and Bases

• Early in the history of chemistry, substances were classified as acids or bases

• One of the first theories to explain acids and bases was put forward in 1887 by the Swedish chemist, Svante Arrhenius

• He proposed that the properties of aqueous solutions of acids and bases are due to the hydrogen ion (H+) and the hydroxide ion (OH-), respectively
Acid

• When dissolved in water, acids have the following properties:
  – Conducts electricity
  – Litmus dye $\rightarrow$ blue to red
  – Sour taste (never taste an acid!)
  – Reacts with and neutralizes a base
  – Reacts with active metals, liberating H gas
Base

• When dissolved in water, bases have the following properties:
  – Conduct electricity
  – Litmus dye $\rightarrow$ red to blue
  – Reacts with and neutralizes an acid
An Arrhenius Acid

• When the colorless, gaseous substance hydrogen chloride (HCl) is added to water, virtually all the HCl molecules ionize into $\text{H}^+$ ions and $\text{Cl}^-$ ions.

• The acidic properties of an acid are due to the hydrogen ions ($\text{H}^+$).

• In the case of HCl, virtually all of the atoms ionize into $\text{H}^+$ ions and $\text{Cl}^-$ ions.
An Arrhenius Acid

• As the HCl molecule disassociates in water the hydrogen ions are transferred from the HCl to the water molecules, as shown in the following:

\[
\text{HCl} + \text{H}_2\text{O} \rightarrow \text{H}_3\text{O}^+ + \text{Cl}^-
\]

– The \( \text{H}_3\text{O}^+ \) is called the hydronium ion

• Therefore, an Arrhenius acid is a substance that produces hydrogen ions, \( \text{H}^+ \), (or hydronium ions, \( \text{H}_3\text{O}^+ \)) in water
HCl Reacts with H$_2$O forming -
Hydronium ions ($H_3O^+$) and Chlorine ions ($Cl^-$)
Strong & Weak Acids

• An acid classified as strong, ionizes almost completely in solution
• Strong acids include HCl, HNO₃, H₂SO₄
  – All three of these in water ionize virtually completely
• Weak acids do not ionize to any great extent in solution
  – At equilibrium, only a small fraction of the molecules disassociate to form H₃O⁺
• Acetic acid, HC₂H₃O₂ (vinegar) is a common weak acid
Strong Acids Ionize more Completely

- *Conduct Electricity Better*
- *Bulb Glows Brightly*
Dynamic Equilibrium

- $\text{HC}_2\text{H}_3\text{O}_2 + \text{H}_2\text{O} \rightleftharpoons \text{H}_3\text{O}^+ + \text{C}_2\text{H}_3\text{O}_2$

- A double arrow may be used if the reverse reaction is significant

- When two ‘competing’ reactions are occurring at the same time, the system is in dynamic equilibrium

- In the case above, only a small fraction of the acetic acid molecules disassociate
Acids are Very Useful Compounds

• Sulfuric acid has a number of different industrial uses
• Refining petroleum, processing steel, fertilizers
• *Dilute HCl helps digest food in the stomach*
• Citric acid ($H_3C_6H_5O_7$) in citrus fruits
• Carbonic ($H_2CO_3$) and Phosphoric ($H_3PO_4$) acids in soft drinks
• Acetic acid ($HC_2H_3O_2$) in vinegar
An Arrhenius Base

• When pure sodium hydroxide (NaOH) is added to water, it dissolves releasing $\text{Na}^+ \text{ and OH}^-$

• The basic properties of NaOH are due to the hydroxide ions (OH$^-$)

• Therefore an Arrhnius base is a substance that produces hydroxide ions (OH$^-$)
An Arrhenius Base

- Ammonia (NH$_3$) is also considered a base even though it does not contain (OH$^-$)
- In water ammonia reacts with H$_2$O to form (OH$^-$)
- $\text{NH}_3 + \text{H}_2\text{O} \rightleftharpoons \text{NH}_4^+ + \text{OH}^-$
- Common household bases include Drano (NaOH) and Windex (NH$_3$)
Water

- Water will slightly ionize by itself

\[ \text{H}_2\text{O} + \text{H}_2\text{O} \leftrightarrow \text{H}_3\text{O}^+ + \text{OH}^- \]

- Therefore all aqueous solutions have both the hydronium ion (H$_3$O$^+$) and the hydroxide ion (OH$^-$)

- For pure water the concentrations of H$_3$O$^+$ and OH$^-$ are equal, and therefore *neutral*

- An acidic solution has more H$_3$O$^+$

- A basic solution has more OH$^-$
Acidic and Basic Solutions

Acidic solution – higher concentration of $\text{H}_3\text{O}^+$
Basic solution – higher concentration of $\text{OH}^-$
The relative acidity or basicity of a solution is commonly designated by citing the pH.

The pH of a solution is a logarithmic measure of the concentration of the hydrogen ion.

Since the pH scale is logarithmic, each drop/rise of 1 on the scale is actually a tenfold increase/decrease in acidity.
The pH Scale

A solution with a pH of 7 is neutral, a solution with a pH less than 7 is acidic, and a solution with a pH greater than 7 is basic.
Body Fluids and pH

• Most body fluids of a healthy person must remain in a very narrow range on the pH scale.

• Thus the pH of different body fluids may be used as a diagnostic measure.
  – *The pH of blood should be between 7.35-7.45*
Acid-Base Reaction

• When an acid is brought in contact with a base its characteristic properties disappear
  – And vice versa
• This is known as an acid-base reaction
• An acid and a hydroxide base react to produce water and a salt
• A salt is an ionic compound composed of any cation except H\(^+\) and any anion except OH\(^-\)
  – Examples of salt include, KCl, Ca\(_3\)(PO\(_4\))\(_2\), CuSO\(_4\).5H\(_2\)O, CaSO\(_4\).2H\(_2\)O
CuSO$_4$·5H$_2$O is blue – anhydrous CuSO$_4$ is white

Adding water will once again hydrate the substance
White Sand National Monument, NM

*The Sand is composed of the salt CaSO$_4$·2H$_2$O*
Acid-Base Reaction – An Example

- Write a balanced equation for stomach acid HCl and milk of magnesia Mg(OH)$_2$
- HCl is an acid and Mg(OH)$_2$ is a base
- Therefore, we know the water and a salt is produced
- HCl + Mg(OH)$_2$ $\rightarrow$ H$_2$O + a salt
- Determine the salt from the given reactants
- HCl + Mg(OH)$_2$ $\rightarrow$ H$_2$O + MgCl$_2$
- Balance the equation, using Mg$^{2+}$
- 2 HCl + Mg(OH)$_2$ $\rightarrow$ 2 H$_2$O + MgCl$_2$
Acid-Base Reaction – Confidence Exercise

• Write a balanced equation for stomach acid $HCl$ and aluminum hydroxide (Di-Gel) $Al(OH)_3$
• $HCl$ is an acid and $Al(OH)_3$ is a base
• $\therefore$ we know the water and a salt is produced
• $HCl + Al(OH)_3 \rightarrow H_2O + a \text{ salt}$
• Determine the salt from the given reactants
• $HCl + Al(OH)_3 \rightarrow H_2O + AlCl_3$
• Balance the equation, using $Al^{3+}$
• $3 \ HCl + Al(OH)_3 \rightarrow 3 \ H_2O + AlCl_3$
Acid-Carbonate Reaction

• An acid and a carbonate or hydrogen carbonate react to give CO$_2$, H$_2$O, and a salt

• For example the antacid Tums contains CaCO$_3$ and can also be used to neutralize too much stomach acid
Acid-Carbonate Reaction – An Example

- Write the balance equation for the reaction between stomach acid (HCl) and Tums (CaCO₃)
- This is an acid-carbonate reaction, therefore the products are CO₂, H₂O, and a salt
- HCl + CaCO₃ → CO₂ + H₂O + a salt
- Determine the salt from the given reactants
- HCl + CaCO₃ → CO₂ + H₂O + CaCl₂
- Balance the equation, using Ca²⁺
- 2 HCl + CaCO₃ → CO₂ + H₂O + CaCl₂
Acid-Carbonate Reaction

- \(2\text{HCl} + \text{CaCO}_3 \rightarrow \text{CO}_2 + \text{H}_2\text{O} + \text{CaCl}_2\)
Acid-Carbonate Reaction – Confidence Exercise

• Write the balance equation for the reaction between nitric acid (HNO₃) and Na₂CO₃
• This is an acid-carbonate reaction, ∴ the products are CO₂, H₂O, and a salt
• HNO₃ + Na₂CO₃ → CO₂ + H₂O + a salt
• Determine the salt from the given reactants
• HNO₃ + Na₂CO₃ → CO₂ + H₂O + NaNO₃
• Balance the equation
• 2 HNO₃ + Na₂CO₃ → CO₂ + H₂O + 2 NaNO₃
Double-Replacement Reactions

• Both acid-base and acid-carbonate are types of double-replacement reactions
• In this type of reaction the positive and negative components “change partners”
• The general format is as follows …
• $AB + CD \rightarrow AD + CB$
Double-Replacement Reactions

Have the format $AB + CD \rightarrow AD + CB$, basically all four components ($K$, $Cl$, $Ag$, $NO_3$) “change partners”
Double-Replacement Precipitation Reaction

• When two aqueous solutions of salts are mixed, a precipitate often results.

• One pair of ions form a precipitate, the other pair remains in dissolved.
Double-Replacement Precipitation Reaction - An Example

• Write the equation for the double-replacement reaction between KI (aq) and Pb(NO₃)₂ (aq)

  • KI (aq) + Pb(NO₃)₂ (aq) → AD + CB
  • 2 KI (aq) + Pb(NO₃)₂ (aq) → 2 K(NO₃) (aq) + PbI₂ (s)
Double-Replacement Precipitation Reaction - *Confidence Exercise*

- **Write the equation for the double-replacement reaction between Na$_2$SO$_4$ (aq) and BaCl$_2$ (aq)**
  
- Na$_2$SO$_4$ (aq) + BaCl$_2$ (aq) $\rightarrow$ AD + CB

- Complete the equation by exchanging partners and balancing the equation

- Na$_2$SO$_4$ (aq) + BaCl$_2$ (aq) $\rightarrow$ 2 NaCl (aq) + BaSO$_4$ (s)
Oxidation & Reduction

- **Oxidation** - oxygen combines with a substance
  - *Or when an atom or ion loses electrons*
- **Reduction** – oxygen removed from a substance
  - *Or when an atom or ion gains electrons*
- All of the electrons lost by an atom or ion must be gained by other atoms or ions
  - *therefore oxidation and reduction occur at the same time.*
- These type of reactions are called **oxidation-reduction reactions** or **redox reactions** for short.
Oxidation-Reduction Reaction
Steel-Making $\rightarrow$ Reduction of Hematite Ore

- Hematite ($\text{Fe}_2\text{O}_3$) must be reduced with coke (C) in a furnace to form pure iron.
- $2 \text{Fe}_2\text{O}_3 + 3 \text{C} \rightarrow 4 \text{Fe} + 3 \text{CO}_2$
- $\text{Fe}^{3+} \rightarrow \text{Fe}$
  - Each Fe gains 3 electrons – reduction
- $\text{C} \rightarrow \text{C}^{4+}$
  - Each C loses 4 electrons - oxidation
- Therefore hematite (iron oxide) has been reduced
Activity

- The relative activity of any metal is its tendency to lose electrons to ions of another metal or to hydrogen ions.
- A metal’s activity can be determined by placing the metal in a solution that contains ions of another metal.
- If the test metal replaces the metal in solution, then the test metal is said to be more active.
Activity Series

- *Shows the relative activity of a number of common metals*

<table>
<thead>
<tr>
<th>Metals</th>
<th>Ion Found</th>
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<tbody>
<tr>
<td>Lithium</td>
<td>Li⁺</td>
</tr>
<tr>
<td>Potassium</td>
<td>K⁺</td>
</tr>
<tr>
<td>Calcium</td>
<td>Ca²⁺</td>
</tr>
<tr>
<td>Sodium</td>
<td>Na⁺</td>
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<tr>
<td>Magnesium</td>
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</tbody>
</table>

*Hydrogen is in capital letters because the activities of the metals are often determined in relation to the activity of hydrogen.*
Single-Replacement Reaction

• If an element $a$ is listed above element $B$ in the activity series ...
  – $A$ is more active than $B$.
  – $A$ will replace $B$ in a aqueous solution.
• This is called a single-replacement reaction.
• The general format for this type of reaction is
• $A + BC \rightarrow B + AC$
• For example $Zn$ and $CuSO_4$
Single-Replacement Reactions

- \( A + BC \rightarrow B + AC \)
  - \( \text{Cu} \) precipitates, but the \( \text{ZnSO}_4 \) is soluble.
  - \( \text{Zn} \) loses electrons \( \rightarrow \) \( \therefore \) has been oxidized
  - \( \text{Cu} \) gains electrons \( \rightarrow \) \( \therefore \) has been reduced
Zinc replaces copper ions therefore Zn is more active than Cu.

\[ \text{Zn} + \text{CuSO}_4 \rightarrow \text{Cu} + \text{ZnSO}_4 \]
Single-Replacement Reaction
An Example

• Refer to the activity series table and predict whether placing copper metal in a solution of silver nitrate will lead to a reaction. If so, complete and balance the equation.

• Cu is above Ag, therefore a reaction will occur.

• Cu will be oxidized to Cu^{2+}.

• Ag^{+} will be reduced to Ag.

• NO_{3}^{-} is a polyatomic ion and will stay intact.
Single-Replacement Reaction

An Example (cont.)

- The unbalanced equation for the reaction is:
- \( \text{Cu} + \text{AgNO}_3(aq) \rightarrow \text{Ag} + \text{Cu(NO}_3)_2(aq) \)
- Balance the equation.
  - Remember that the \( \text{NO}_3^- \) stays intact and is balanced as a unit.
- \( \text{Cu} + 2 \text{AgNO}_3(aq) \rightarrow 2 \text{Ag} + \text{Cu(NO}_3)_2(aq) \)
- Cu loses electrons. (oxidation)
- Ag gains electrons. (reduction)
Single-Replacement Reaction Between Copper Wire and Silver Nitrate

\[ Cu + 2 \text{AgNO}_3(aq) \rightarrow 2 \text{Ag} + \text{Cu(NO}_3)_2(aq) \]
Single-Replacement Reaction

Confidence Exercise

• Suppose you place Al metal in a solution of CuSO$_4$ (aq), and also put a strip of Cu metal in a solution of Al$_2$(SO$_4$)$_3$(aq).

• Refer to the activity series and predict which single-replacement reaction will take place.

• Complete and balance the equation for the reaction.
Activity Series

- **Note that Al is more active than Cu, therefore Al will replace the Cu in CuSO₄.**

<table>
<thead>
<tr>
<th>Metals</th>
<th>Ion Found</th>
</tr>
</thead>
<tbody>
<tr>
<td>Lithium</td>
<td>Li⁺</td>
</tr>
<tr>
<td>Potassium</td>
<td>K⁺</td>
</tr>
<tr>
<td>Calcium</td>
<td>Ca²⁺</td>
</tr>
<tr>
<td>Sodium</td>
<td>Na⁺</td>
</tr>
<tr>
<td>Magnesium</td>
<td>Mg²⁺</td>
</tr>
<tr>
<td>Aluminum</td>
<td>Al³⁺</td>
</tr>
<tr>
<td>Zinc</td>
<td>Zn²⁺</td>
</tr>
<tr>
<td>Chromium</td>
<td>Cr³⁺</td>
</tr>
<tr>
<td>Iron</td>
<td>Fe²⁺</td>
</tr>
<tr>
<td>Nickel</td>
<td>Ni²⁺</td>
</tr>
<tr>
<td>Tin</td>
<td>Sn²⁺</td>
</tr>
<tr>
<td>Lead</td>
<td>Pb²⁺</td>
</tr>
<tr>
<td>HYDROGEN*</td>
<td>H⁺</td>
</tr>
<tr>
<td>Copper</td>
<td>Cu²⁺</td>
</tr>
<tr>
<td>Silver</td>
<td>Ag⁺</td>
</tr>
<tr>
<td>Platinum</td>
<td>Pt²⁺</td>
</tr>
<tr>
<td>Gold</td>
<td>Au³⁺</td>
</tr>
</tbody>
</table>

*Hydrogen is in capital letters because the activities of the metals are often determined in relation to the activity of hydrogen.*
Single-Replacement Reaction
Confidence Exercise (cont.)

• $\text{Al} + \text{CuSO}_4(aq) \rightarrow \text{Cu} + \text{Al}_2(\text{SO}_4)_3(aq)$

• Balance the equation.
  – *Remember that $\text{SO}_4^{2-}$ stays intact and is balanced as a unit.*

• $2 \text{Al} + 3 \text{CuSO}_4(aq) \rightarrow 3 \text{Cu} + \text{Al}_2(\text{SO}_4)_3(aq)$

• Al loses electrons. (oxidation)

• Cu gains electrons. (reduction)
Single-Replacement Reaction with Acids

• Metals above hydrogen in the activity series will undergo a single-replacement reaction with acids, giving off hydrogen gas and a salt of the metal.
Single-Replacement Reaction with Acid

- $Fe$ reacts with $H_2SO_4$ to give off $H$ and an Fe salt
Single-Replacement Reactions with Water

- Metals above magnesium in the activity series will undergo a vigorous single-replacement reaction with liquid water, giving off hydrogen gas and the metal hydroxide.

- For example, Ca reacts with water according to the following equation:

  \[ \text{Ca} + 2 \text{H}_2\text{O} \rightarrow \text{H}_2 + \text{Ca(OH)}_2 \]

  - \text{Ca} loses electrons. (oxidation)
  - \text{H} gains electrons. (reduction)
Active Metals and Water

Metals more active than Mg will react with water.
# A Summary of Reaction Types

<table>
<thead>
<tr>
<th>Reaction Type</th>
<th>Example</th>
</tr>
</thead>
<tbody>
<tr>
<td>Combination</td>
<td>$2 \text{Mg} + \text{O}_2 \rightarrow 2 \text{MgO}$</td>
</tr>
<tr>
<td>Decomposition</td>
<td>$2 \text{HgO} \rightarrow 2 \text{Hg} + \text{O}_2$</td>
</tr>
<tr>
<td>Hydrocarbon combustion (complete)</td>
<td>$\text{C}_2\text{H}_4 + 3 \text{O}_2 \rightarrow 2 \text{CO}_2 + 2 \text{H}_2\text{O}$</td>
</tr>
<tr>
<td>Single-replacement</td>
<td></td>
</tr>
<tr>
<td>(a) two metals</td>
<td>$\text{Zn} + \text{CuSO}_4 \rightarrow \text{Cu} + \text{ZnSO}_4$</td>
</tr>
<tr>
<td>(b) metal and acid</td>
<td>$\text{Fe} + 2 \text{HCl} \rightarrow \text{H}_2 + \text{FeCl}_2$</td>
</tr>
<tr>
<td>Double-replacement</td>
<td></td>
</tr>
<tr>
<td>(a) precipitation</td>
<td>$\text{BaCl}_2 + \text{Na}_2\text{SO}_4 \rightarrow \text{BaSO}_4(s) + 2 \text{NaCl}$</td>
</tr>
<tr>
<td>(b) acid–base</td>
<td>$2 \text{HCl} + \text{Ca(OH)}_2 \rightarrow 2 \text{H}_2\text{O} + \text{CaCl}_2$</td>
</tr>
<tr>
<td>(c) acid–carbonate</td>
<td>$\text{H}_2\text{SO}_4 + \text{Na}_2\text{CO}_3 \rightarrow \text{H}_2\text{O} + \text{CO}_2 + \text{Na}_2\text{SO}_4$ (The $\text{H}_2\text{CO}_3$ that is initially formed decomposes.)</td>
</tr>
</tbody>
</table>
**Reaction Types**

**Combination Reaction**

- **Example:** \(2 \text{Mg} + \text{O}_2 \rightarrow 2 \text{MgO}\)

  Magnesium reacts with oxygen to form magnesium oxide.

**Decomposition Reaction**

- **Example:** \(2 \text{H}_2\text{O} \rightarrow 2 \text{H}_2 + \text{O}_2\)

  Water decomposes to form hydrogen and oxygen.

**Hydrocarbon Combustion**

- **Example:** \(\text{CH}_4 + \text{O}_2 \rightarrow \text{CO}_2 + 2 \text{H}_2\text{O}\)

  A hydrocarbon reacts with oxygen to form carbon dioxide and water.

**Single-Replacement Reaction**

- **Example:** \(\text{Cu} + 2 \text{AgNO}_3(\text{aq}) \rightarrow 2 \text{Ag} + \text{Cu(NO}_3)_2(\text{aq})\)

  Copper reacts with silver nitrate to form silver and copper(II) nitrate.

**Double-Replacement Reaction**

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

  Potassium iodide reacts with lead(II) nitrate to form lead(II) iodide and potassium nitrate.
A Mole & Avogadro’s Number

• **Mole** – the quantity of a substance that contains as many formula units as there are atoms in exactly 12 grams of $^{12}\text{C}$
  – *Mole is abbreviated “mol.”*
• In exactly 12 grams of $^{12}\text{C}$ there are exactly $6.023 \times 10^{23}$ $^{12}\text{C}$ atoms.
• This huge number ($6.023 \times 10^{23}$) is referred to as **Avagadro’s Number**.
  – *The Italian physicist Avogadro was the first person to use the term molecule.*
• Therefore a *mole* is $6.023 \times 10^{23}$ units.
A Mole & Formula Mass

• A mole of any substance has a mass equal to the same number of grams as its formula mass. (FM)

• For example: Copper has a FM of 63.5μ
  – Therefore a mole of Cu contains $6.023 \times 10^{23}$ atoms and has a mass of 63.5 grams.

• Water has a FM of 18.0 μ
  – A mole of $H_2O$ contains $6.023 \times 10^{23}$ molecules and has a mass of 18 grams.
One Mole of Six Substances
63.5 g of Cu, 27.0 g of Al, 55.8 g of Fe, 32.1 g of S, 253.8 g of I, and 200.6 g of Hg
Avogadro’s Number

• $6.023 \times 10^{23}$ is a huge number
• If we poured out $6.023 \times 10^{23}$ BB’s on the U.S., they would cover the entire country to a depth of 4 inches!
• But how do we actually know how many particles there are in a mole?
A Mole = 6.023 x 10^{23} units

• We know that it takes 96,485 coulombs (C) to reduce 1 mole of singly charged ions to atoms.
  – For example Na^+ + e^- → Na
• Thus, 96,485 C must be the total charge on one mole of electrons.
• Single electron has a charge of 1.6022 x 10^{-19} C.
• Therefore Avogadro’s number must be
• (96,485 C/mol)/(1.6022 x 10^{-19} C/electron)
• = 6.023 x 10^{23} electrons/mole