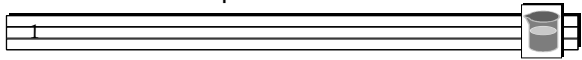




Chapter 11 "Chemical Reactions"

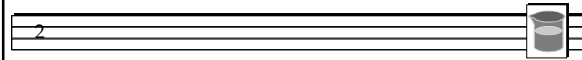
Pre-AP Chemistry
Charles Page High School
Stephen L. Cotton



Section 11.1 Describing Chemical Reactions

● **OBJECTIVES:**

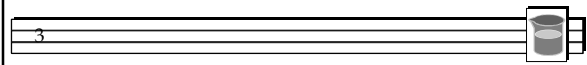
- Describe how to write a *word equation*.



Section 11.1 Describing Chemical Reactions

● **OBJECTIVES:**

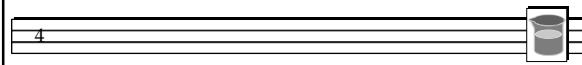
- Describe how to write a *skeleton equation*.



Section 11.1 Describing Chemical Reactions

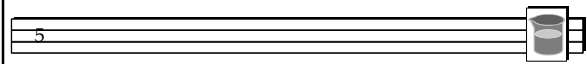
● **OBJECTIVES:**

- Describe the steps for writing a *balanced chemical equation*.



All chemical reactions...

- have two parts:
 1. **Reactants** = the substances you start with
 2. **Products** = the substances you end up with
- The reactants will turn into the products.
- Reactants → Products



114 Figure 11.1 Chemical Changes

- Page 321



In a chemical reaction

- Atoms aren't created or destroyed (according to the Law of Conservation of Mass)
- A reaction can be described several ways:
 - #1. In a **sentence** every item is a word
Copper reacts with chlorine to form copper (II) chloride.
 - #2. In a **word equation** some symbols used
Copper + chlorine → copper (II) chloride

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Symbols in equations? – Text page 323

- the arrow (→) separates the reactants from the products (arrow points to products)
–Read as: “reacts to form” or yields
- The plus sign = “and”
- (s) after the formula = solid: Fe_(s)
- (g) after the formula = gas: CO_{2(g)}
- (l) after the formula = liquid: H₂O_(l)

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Symbols used in equations

- (aq) after the formula = dissolved in water, an aqueous solution:
NaCl_(aq) is a salt water solution
- ↑ used after a product indicates a gas has been *produced*: H₂↑
- ↓ used after a product indicates a solid has been *produced*: PbI₂↓

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Symbols used in equations

- \rightleftharpoons double arrow indicates a reversible reaction (more later)
- Δ shows that heat is supplied to the reaction
- Pt is used to indicate a catalyst is supplied (in this case, platinum is the catalyst)

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What is a catalyst?

- A substance that speeds up a reaction, without being changed or used up by the reaction.
- Enzymes are biological or protein catalysts in your body.

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#3. The Skeleton Equation

- Uses formulas and symbols to describe a reaction
–but doesn't indicate how many; this means they are NOT balanced
- All chemical equations are a *description* of the reaction.

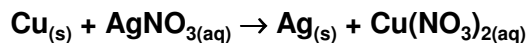
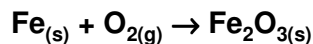
12

Write a skeleton equation for:

1. Solid iron (III) sulfide reacts with gaseous hydrogen chloride to form iron (III) chloride and hydrogen sulfide gas.
2. Nitric acid dissolved in water reacts with solid sodium carbonate to form liquid water and carbon dioxide gas and sodium nitrate dissolved in water.

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Now, read these equations:



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#4. Balanced Chemical Equations

- Atoms can't be created or destroyed in an ordinary reaction:
 - All the atoms we start with we must end up with (meaning: *balanced!*)
- A balanced equation has the same number of each element on both sides of the equation.

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Rules for balancing:

- 1) Assemble the correct formulas for all the reactants and products, using "+" and "→"
- 2) Count the number of atoms of each type appearing on both sides
- 3) Balance the elements *one at a time* by adding *coefficients* (the numbers in front) where you need more - save balancing the H and O until LAST!
(hint: I prefer to save O until the *very* last)
- 4) Double-Check to make sure it is balanced.

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Never

- **Never** change a subscript to balance an equation (You can only change coefficients)
 - If you change the subscript (formula) you are describing a different chemical.
 - H₂O is a different compound than H₂O₂
- **Never** put a coefficient in the *middle* of a formula; they must go only in the front
2NaCl is okay, but Na2Cl is not.

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Practice Balancing Examples

- 2AgNO₃ + Cu → Cu(NO₃)₂ + 2Ag
- 3Mg + N₂ → Mg₃N₂
- 4P + 5O₂ → P₄O₁₀
- 2Na + 2H₂O → H₂ + 2NaOH
- CH₄ + 2O₂ → CO₂ + 2H₂O

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Section 11.2
Types of Chemical Reactions

● **OBJECTIVES:**

– **Describe** the *five general types* of reactions.

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Section 11.2
Types of Chemical Reactions

● **OBJECTIVES:**

– **Predict** the *products* of the five general types of reactions.

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Types of Reactions

- There are probably millions of reactions.
- We can't remember them all, but luckily they will fall into several categories.
- We will learn: a) the **5 major types**.
- We will be able to: b) **predict the products**.
- For some, we will be able to: c) predict whether or not they will **happen at all**.
- **How?** We recognize them by their **reactants**

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#1 - Combination Reactions

- Combine = put together
- 2 substances combine to make one compound (also called "synthesis")
- $\text{Ca} + \text{O}_2 \rightarrow \text{CaO}$
- $\text{SO}_3 + \text{H}_2\text{O} \rightarrow \text{H}_2\text{SO}_4$
- We can predict the products, especially if the *reactants are two elements*.
- $\text{Mg} + \text{N}_2 \rightarrow \underline{\text{Mg}_3\text{N}_2}$ (symbols, charges, cross)

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Complete and balance:

- $\text{Ca} + \text{Cl}_2 \rightarrow$
- $\text{Fe} + \text{O}_2 \rightarrow$ (assume iron (II) oxide is the product)
- $\text{Al} + \text{O}_2 \rightarrow$
- Remember that the **first step** is to write the *correct formulas* – you can still change the *subscripts* at this point, but not later while balancing!
- **Then balance** by changing just the coefficients only

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#1 – Combination Reactions

- **Additional Important Notes:**
 - a) Some nonmetal oxides **react with water** to produce an acid:
 $\text{SO}_2 + \text{H}_2\text{O} \rightarrow \text{H}_2\text{SO}_3$
(This is what happens to make "acid rain")
 - b) Some metallic oxides **react with water** to produce a base:
 $\text{CaO} + \text{H}_2\text{O} \rightarrow \text{Ca(OH)}_2$

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#2 - Decomposition Reactions

- decompose = fall apart
- one reactant breaks apart into two or more elements or compounds.
- NaCl Na + Cl₂
- CaCO₃ CaO + CO₂
- Note that energy (heat, sunlight, electricity, etc.) is usually required

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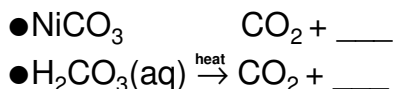
#2 - Decomposition Reactions

- We can predict the products if it is a binary compound (which means it is made up of only two elements)
 - It breaks apart into the elements:
- H₂O
- HgO

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#2 - Decomposition Reactions

- If the compound has more than two elements you must be given one of the products
 - The other product will be from the missing pieces



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#3 - Single Replacement Reactions

- One element replaces another
- Reactants must be *an element and a compound*.
- Products will be a different element and a different compound.
- Na + KCl → K + NaCl (Cations switched)
- F₂ + LiCl → LiF + Cl₂ (Anions switched)

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#3 Single Replacement Reactions

- Metals will replace other metals (and they can also replace hydrogen)
- K + AlN →
- Zn + HCl →
- Think of water as: HOH
 - Metals replace the first H, and then combines with the hydroxide (OH).
- Na + HOH →

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#3 Single Replacement Reactions

- We can even tell *whether or not* a single replacement reaction will happen:
 - Because some chemicals are more “active” than others
 - More active *replaces* less active
- There is a list on page 333 - called the Activity Series of Metals
- Higher on the list replaces those lower.

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The "Activity Series" of Metals

Higher activity	█ Lithium
	█ Potassium
	█ Calcium
	█ Sodium
	█ Magnesium
	█ Aluminum
	█ Zinc
	█ Chromium
	█ Iron
	█ Nickel
	█ Lead
	█ Hydrogen
	█ Bismuth
	█ Copper
	█ Mercury
	█ Silver
	█ Platinum
Lower activity	█ Gold

1) Metals can replace other metals, provided they are above the metal they are trying to replace (for example, zinc will replace lead)

2) Metals above hydrogen can replace hydrogen in acids.

3) Metals from sodium upward can replace hydrogen in water.

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The "Activity Series" of Halogens

Higher Activity	█ Fluorine
	█ Chlorine
	█ Bromine
	█ Iodine
Lower Activity	

Halogens can replace other halogens in compounds, provided they are above the halogen they are trying to replace.

$$2\text{NaCl}_{(s)} + \text{F}_{2(g)} \rightarrow 2\text{NaF}_{(s)} + \text{Cl}_{2(g)}$$

$$\text{MgCl}_{2(s)} + \text{Br}_{2(g)} \rightarrow \text{No Reaction!}$$

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#3 Single Replacement Reactions Practice:

- $\text{Fe} + \text{CuSO}_4 \rightarrow$
- $\text{Pb} + \text{KCl} \rightarrow$
- $\text{Al} + \text{HCl} \rightarrow$

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#4 - Double Replacement Reactions

- Two things replace each other.
 - Reactants must be two ionic *compounds*, in aqueous solution
- $\text{NaOH} + \text{FeCl}_3 \rightarrow$
 - The positive ions change place.
- $\text{NaOH} + \text{FeCl}_3 \rightarrow \text{Fe}^{+3} \text{OH}^- + \text{Na}^{+1} \text{Cl}^{-1}$
 $= \text{NaOH} + \text{FeCl}_3 \rightarrow \text{Fe}(\text{OH})_3 + \text{NaCl}$

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#4 - Double Replacement Reactions

- Have certain "driving forces", or reasons
 - Will only happen if one of the products:
 - doesn't dissolve in water and forms a solid (a "precipitate"), or
 - is a gas that bubbles out, or
 - is a molecular compound (which will usually be water).

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Complete and balance:

- assume all of the following reactions actually take place:
 - $\text{CaCl}_2 + \text{NaOH} \rightarrow$
 - $\text{CuCl}_2 + \text{K}_2\text{S} \rightarrow$
 - $\text{KOH} + \text{Fe}(\text{NO}_3)_3 \rightarrow$
 - $(\text{NH}_4)_2\text{SO}_4 + \text{BaF}_2 \rightarrow$

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How to recognize which type?

● **Look at the reactants:**

E + E = Combination

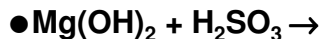
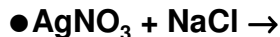
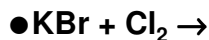
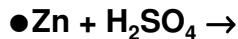
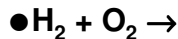
C = Decomposition

E + C = Single replacement

C + C = Double replacement

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Practice Examples:



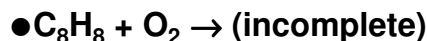
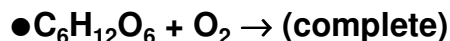
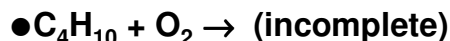
38

#5 – Combustion Reactions

- Combustion means **“add oxygen”**
- Normally, a compound composed of only C, H, (and maybe O) is reacted with oxygen – usually called **“burning”**
- If the combustion is complete, the products will be CO_2 and H_2O .
- If the combustion is incomplete, the products will be CO (or possibly just C) and H_2O .

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Combustion Reaction Examples:



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SUMMARY: An equation...

- Describes a reaction
- Must be balanced in order to follow the Law of Conservation of Mass
- Can only be balanced by changing the coefficients.
- Has *special symbols* to indicate the physical state, if a catalyst or energy is required, etc.

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Reactions

- Come in 5 major types.
- We can tell what type they are by looking at the reactants.
- Single Replacement happens based on the *Activity Series*
- Double Replacement happens if one product is: 1) a precipitate (an insoluble solid), 2) water (a molecular compound), or 3) a gas.

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- There are some more practice problems of balancing equations located from:
 - my website
 - Interesting Links
 - Balancing Equations

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Section 11.3 Reactions in Aqueous Solution

- **OBJECTIVES:**
 - Describe the information found in a *net ionic equation*.

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Section 11.3 Reactions in Aqueous Solution

- **OBJECTIVES:**
 - Predict the formation of a *precipitate* in a double replacement reaction.

45



Net Ionic Equations

- Many reactions occur in water- that is, in aqueous solution
- When dissolved in water, many ionic compounds “dissociate”, or separate, into cations and anions
- Now we are ready to write an ionic equation

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Net Ionic Equations

- Example (needs to be a double replacement reaction)

$$\text{AgNO}_3 + \text{NaCl} \rightarrow \text{AgCl} + \text{NaNO}_3$$
 1. this is the full balanced equation
 2. next, write it as an ionic equation by splitting the compounds into their ions:

$$\text{Ag}^{1+} + \text{NO}_3^{1-} + \text{Na}^{1+} + \text{Cl}^{1-} \rightarrow$$

$$\text{AgCl} + \text{Na}^{1+} + \text{NO}_3^{1-}$$

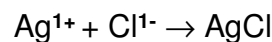
Note that the AgCl did not ionize, because it is a “precipitate”

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Net Ionic Equations

3. *simplify* by crossing out ions not directly involved (called spectator ions)



This is called the net ionic equation

Let’s talk about precipitates before we do some other examples

48



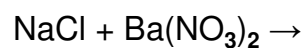
Predicting the Precipitate

- Insoluble salt = a precipitate [note Figure 11.11, p.342 (AgCl)]
- General solubility rules are found:
 - a) Table 11.3, p. 344 in textbook
 - b) Reference section - page R54 (back of textbook)
 - c) Lab manual Table A.3, page 332
 - d) Your **periodic table handout**

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Let's do some examples together of net ionic equations, starting with these reactants:



End of Chapter 11

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