UNIT 4: CHEMICAL REACTIONS & STOICHIOMETRY

BOOK 1: BALANCING EQUATIONS, TYPES OF REACTIONS & ENERGY OF REACTIONS

Name: KEY

Block: _____
Part 1: Introduction to Chemical Reactions & Balancing

The beauty, malleability, and corrosion-resistance of copper have long been valued and employed in everything from artwork to architecture. For example, many iconic structures have used copper cladding, sheets, flashing, gutters, downspouts, etcetera because of their beauty and durability. And yet, with time, the copper will develop a characteristic patina - a green copper(II) oxide skin - due to a chemical reaction between the copper and atmospheric oxygen. While the cupric oxide layer hides the lustre of the underlying metal, it also protects it from further reaction. For this reason, a properly installed copper roof can be expected to outlast the structure it adorns.

The oxidation of a copper roof, the rusting of a nail and the tarnishing of silver earrings are all examples of chemical change. While such chemical changes occur over months or years, other chemical changes play out on much shorter timescales. Cooking an egg, for example, takes several minutes. Explosions, on the other hand, occur so quickly that we think of them as being instantaneous.

Chemical reactions are continuously occurring all around and within you. In fact, it is a series of chemical reactions within your brain that are allowing you to see and understand the words you are reading on this page. In this unit, we will study the fundamentals of chemical change, develop a method to balance reaction equations, and learn how to predict the products that will form during characteristic forms of chemical reaction.

Chemistry is the study of matter and its changes. Earlier this year, you learned how physical and chemical changes differ. Chemical change always produce new substances with new properties and their own unique chemical formulas. Such changes involve the breaking and formation of chemical bonds. They are referred to as chemical reactions. The processes of photosynthesis and aerobic cellular respiration, for example, involve a series of chemical reactions that produce and use oxygen. These reactions are taking place right now in our bodies and in most of the living things in our world.

Another series of oxygen-requiring chemical reactions are necessary to heat our homes and move our vehicles from place to place. Chemical reactions involving oxygen can also be a problem when the metal in many human-made objects spontaneously breaks down in a chemical process called corrosion. Most people take chemical reactions for granted as if they were magic. It is important to appreciate that every waking moment of our lives, the matter of our world is continually undergoing an endless series of chemical reactions.

A. What Characterizes All Chemical Reactions?

All chemical reactions have two key features in common:

1. One or more substances called reactants are converted into one or more new substances, called products.

   energy

2. A change in (with new and different properties) occurs (e.g. emission or absorption of heat, light, electricity, etc.)

Other evidence of a chemical reaction:
- different coloured materials may be formed
- new phases may be formed (be careful, it's not only a phase change)
B. Chemical Reaction Equations

Chemical word equations are descriptive but chemical reaction equations are much more efficient.

Using chemical symbols and numbers creates a "chemical shorthand" for all chemists to understand... much like txt language saves our thumbs and shortens convs.

The general form of a chemical equation is:

\[
\text{Reactants} \rightarrow \text{Products}
\]

Chemical Word Equation: "Sodium metal reacts with chlorine gas to produce sodium chloride crystals."

Skeleton chemical reaction equation:

\[
\text{Na} + \text{Cl}_2 \rightarrow \text{NaCl}
\]

Balanced chemical reaction equation:

\[
2\text{Na} + \text{Cl}_2 \rightarrow 2\text{NaCl}
\]

Formula equations may sometimes be referred to as _molecular_ equations. Both of these equations indicate a starting substance or _reactant(s)_ forming new substances or _product(s)_.

Formula equations may include italicized subscript letters to indicate the physical state of the reactants and products: solid (\(s\)), liquid (\(l\)), or gas (\(g\)).

*The symbol (\(aq\)) stands for _aq_, meaning the substance is dissolved in water (\(H_2O\)).

Word equation: mercury(II) oxide → mercury + oxygen

Formula equation: \(2\text{HgO} \rightarrow \text{Hg} + \text{O}_2\)

You can tell that a chemical reaction has occurred by looking for evidence of chemical change.

Some examples of such evidence include:
- noticeable absorption or release of ENERGY (heat or sometimes light)
- a color change
- evolution of a gas (bubbles may be visible)
- formation of a solid from two solutions. Such a solid is called a precipitate.

**Practice**

1. Convert the following from a word equation to a formula equation:
   Calcium oxide powder is combined with water to form calcium hydroxide solid.

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

2. Convert the following from a formula equation to a word equation:

   \[
   \text{H}_2\text{CO}_3 \rightarrow \text{CO}_2 + \text{H}_2\text{O}
   \]
C. Chemical Systems

A system is the part of the universe you are observing.

It is a _______________ if nothing can enter or escape.
It is an _______________ is something can enter or escape.

A system can be open with respect to one thing but closed with respect to another thing.

Examples:

a) A pot of boiling water without a lid
   - is an open system because water molecules and heat energy can escape

b) Boiling water is poured into a vacuum bottle and then the lid is securely sealed.
   - is a system (bottle and water) that is ___________ with respect to mass
     but ___________ with respect to heat. Why?

D. The Conservation Laws

Definition: Any quantity that does not change during a reaction is said to be conserved.

A Conservation Law

- is an _______________ law
- tells you what _______________ under a special set of circumstances.

Conservation Laws YOU Must Know:

During any chemical reaction occurring in a closed system:

1. Law of Conservation of Mass: _______________
2. Law of Conservation of Atoms: _______________
3. Law of Conservation of Charge: _______________
ASSIGNMENT #1: Hebden pg 106-107 Questions #1-6
Complete ALL assignments on a separate piece of paper and attach to your booklet when handing in at the end of the unit. Be sure to clearly number each assignment with a heading.

E. Balancing Chemical Reaction Equations

Famous chemist, Antoine Lavoisier’s experiments were more quantitative than those of others in his time. That is, Lavoisier liked to measure the volumes and masses of the chemicals he studied. Lavoisier is generally credited with formulating the law of conservation of mass.

A brief examination of the equation for the decomposition of mercury(II) oxide, \( \text{HgO(s)} \rightarrow \text{Hg(l)} + \text{O}_2(g) \), shows that it does not obey the law of conservation of mass.

The reactant, \( \text{HgO} \) contains **LESS** oxygen atom than the products, \( \text{Hg} \) and \( \text{O}_2 \).

![Unbalanced chemical equation](image)

To show that the mass before and after a chemical reaction occurs remains constant, the formula equation has to be **BALANCED**.

Balancing a chemical equation requires the placement of coefficients in front of reactant and/or product species.

Coefficients are numbers that **multiply** the entire chemical species that follows them.

These numbers ensure that the number of **atoms of each kind** on the reactant side is equal to those on the product side of the equation.

It is critical to remember balancing must always involve the placement of coefficients and NEVER the changing of subscripts. Altering the subscripts will give an incorrect formula for a substance.

In Chemistry 11, “trial and error” won’t cut it anymore! Here is a method that always ;-) works.

**Example**: Balance: \( (\text{NH}_4)_2\text{PO}_4 + \text{NaOH} \rightarrow \text{Na}_3\text{PO}_4 + \text{NH}_3 + \text{H}_2\text{O} \)

**The Method**

**Note**: ✅ Until you are finished balancing, missing coefficients are treated as zeros!!

✅ Do not start with atoms that are easy or difficult to balance; it's best to balance them last.

Elemental species are easy; elements that occur in more than one species on each side (usually O and H) are tough!

✅ Balance preserved groups (those that don’t come apart) whenever possible.
Balance: \((\text{NH}_4\text{)}_2\text{PO}_4 + \text{NaOH} \rightarrow \text{Na}_3\text{PO}_4 + \text{NH}_3 + \text{H}_2\text{O}\)

**Step 1**
Find an element that only occurs in one species on each side; these are usually metal ions. Put a coefficient in front of the two species so as to balance the element of interest.

(\*\* Very Important: The 1\* step is the only step where you place coefficients on both sides. \*\*)

\[ \text{\underbrace{1\ (\text{NH}_4\text{)}_2\text{PO}_4}} + \underbrace{3\ \text{NaOH}} \rightarrow \underbrace{1\ \text{Na}_3\text{PO}_4} + \underbrace{\_\ \text{NH}_3} + \underbrace{\_\ \text{H}_2\text{O}} \]

**Step 2**
You have successfully balanced one element. Booyah! The problem has now becomes easier to solve. The coefficients you added will also have fixed another element (or group) on one side.

Now place a coefficient to balance the atom (or group) on the opposite side.

\[ \text{\underbrace{2\ (\text{NH}_4\text{)}_2\text{PO}_4}} + \underbrace{3\ \text{NaOH}} \rightarrow \underbrace{2\ \text{Na}_3\text{PO}_4} + \underbrace{3\ \text{NH}_3} + \underbrace{\_\ \text{H}_2\text{O}} \]

**Step 3**
Repeat the process until all of the elements are balanced.

\[ \text{\underbrace{2\ (\text{NH}_4\text{)}_2\text{PO}_4}} + \underbrace{3\ \text{NaOH}} \rightarrow \underbrace{2\ \text{Na}_3\text{PO}_4} + \underbrace{3\ \text{NH}_3} + \underbrace{\_\ \text{H}_2\text{O}} \]

**Step 4**
Omit coefficients of 1 in your final answer. \(\text{Always do a check to make sure that all atoms are balanced.}\)

\[ \text{\underbrace{2\ (\text{NH}_4\text{)}_2\text{PO}_4}} + \underbrace{3\ \text{NaOH}} \rightarrow \underbrace{2\ \text{Na}_3\text{PO}_4} + \underbrace{3\ \text{NH}_3} + \underbrace{\text{H}_2\text{O}} \]

**Advanced Tips:**
- If you want \(n\) atoms of a polyatomic element, then multiply it by \(\text{n/}x\), where \(x\) is the multiplicity of that element in the molecule.

  Examples:
  - 7 oxygen atoms from \(\text{NO}_3\) molecule?
    \[ n = \frac{7}{3} \text{ and } x = 3 \text{ Therefore use } \frac{7}{3} \times \text{O}_2 \]
  - 15 phosphorus atoms from \(\text{P}_4\) molecule?
    \[ n = \frac{15}{1} \text{ and } x = 1 \text{ Therefore use } 15 \times \text{P}_4 \]

- Get rid of fractions as soon as they appear (by multiplying both sides through by the denominator). Remember that species without coefficients really have zeros in front of them and are therefore unchanged by this operation.

  multiply the entire balanced equation... not just the fraction
  (i.e.: any coefficients will change)
Balancing Formula Equations

Ba(s) + Al(NO₃)₃(aq) → Ba(NO₃)₂(aq) + Al(s)

What to Think about

1. Barium and aluminum must be left until last as both appear in elemental form.

2. The nitrate ion may be balanced as a group as it is the only aggregation of nitrogen and oxygen in the equation. The way several atoms are combined in an equation is called their aggregation. Consider the lowest common multiple of nitrate’s subscripts 3 and 2 to determine the first set of coefficients. That number is 6 so coefficients of 2 and 3 are required to make the number of reactant and product nitrate groups equal.

3. Now it is simply a matter of balancing the elements. As the coefficient 2 is first, you could balance aluminum next.

4. Complete the balance with the barium and replace the phase indicators if required.

5. Be sure to complete the process by checking the total number of barium, aluminum, nitrogen, and oxygen atoms on each side.

How to Do It

[Diagram showing the balancing process with coefficients and reactions]

Example 1

Balance: 2K + 2H₂O → 2KOH + H₂

Example 2

Balance: 2MoCl₃ + 3O₂ + 1AgCl → 2MoCl₄ + 1Ag₂O

⇒ 4MoCl₃ + O₂ + 4AgCl → 4MoCl₄ + 2Ag₂O
Balancing Equations

Balance the following chemical reaction equations using the method(s) shown in class. NO CREDIT will be given if you fail to show your steps clearly.

7. \(_{\text{Sn}} + \text{O}_2 \rightarrow \text{SnO}\)

9. \(_{\text{N}_2} + \text{H}_2 \rightarrow \text{NH}_3\)

11. \(_{\text{NH}_3} + \text{O}_2 \rightarrow \text{N}_2 + \text{H}_2\text{O}\)

13. \(_{\text{KNO}_3} \rightarrow \text{KNO}_2 + \text{O}_2\)

15. \(_{\text{C}_3\text{H}_2} + \text{O}_2 \rightarrow \text{CO}_2 + \text{H}_2\text{O}\)

17. \(_{2\text{KOH} + \text{H}_2\text{SO}_4} \rightarrow \text{K}_2\text{SO}_4 + \text{H}_2\text{O}\)

19. \(_{\text{C}} + \text{SO}_2 \rightarrow \text{CS}_2 + \text{CO}\)

ANSWERS:

7. 2 Sn + O\(_2\) → 2 SnO
8. H\(_2\) + Cl\(_2\) → 2 HCl
9. N\(_2\) + 3 H\(_2\) → 2 NH\(_3\)
10. 2 Na + 2 H\(_2\)O → 2 NaOH + H\(_2\)
11. 4 NH\(_3\) + 3 O\(_2\) → 2 N\(_2\) + 6 H\(_2\)O
12. 2 C\(_6\)H\(_4\) + 19 O\(_2\) → 12 CO\(_2\) + 14 H\(_2\)O
13. 2 KNO\(_3\) → 2 KNO\(_2\) + O\(_2\)
14. CaC\(_2\) + 2 O\(_2\) → Ca + 2 CO\(_2\)
15. C\(_5\)H\(_12\) + 8 O\(_2\) → 5 CO\(_2\) + 6 H\(_2\)O
16. K\(_2\)SO\(_4\) + BaCl\(_2\) → 2 KCl + BaSO\(_4\)
17. 2 KOH + H\(_2\)SO\(_4\) → K\(_2\)SO\(_4\) + 2 H\(_2\)O
18. Ca(OH)\(_2\) + 2 NH\(_4\)Cl → 2 NH\(_3\) + CaCl\(_2\) + 2 H\(_2\)O
19. 5 C + 2 SO\(_2\) → CS\(_2\) + 4 CO
20. $\text{Mg}_3\text{N}_2 + 6 \text{H}_2\text{O} \rightarrow 3 \text{Mg(OH)}_2 + 2 \text{NH}_3$

21. $\text{V}_2\text{O}_5 + 5 \text{Ca} \rightarrow 5 \text{CaO} + 2 \text{V}$

22. $2\text{Na}_2\text{O}_2 + 2 \text{H}_2\text{O} \rightarrow 4 \text{NaOH} + \text{O}_2$

23. $\text{Fe}_3\text{O}_4 + 4 \text{H}_2 \rightarrow 3 \text{Fe} + 4 \text{H}_2\text{O}$

24. $\text{Cu} + 2 \text{H}_2\text{SO}_4 \rightarrow \text{CuSO}_4 + 2 \text{H}_2\text{O} + \text{SO}_2$

25. $2\text{Al} + 3 \text{H}_2\text{SO}_4 \rightarrow 3 \text{H}_2 + \text{Al}_2(\text{SO}_4)_3$

26. $2\text{Si}_4\text{H}_{10} + 13 \text{O}_2 \rightarrow 8 \text{SiO}_2 + 10 \text{H}_2\text{O}$

27. $4\text{NH}_3 + \text{O}_2 \rightarrow 2 \text{N}_2\text{H}_4 + 2 \text{H}_2\text{O}$

28. $2\text{C}_15\text{H}_30 + 45 \text{O}_2 \rightarrow 30 \text{CO}_2 + 30 \text{H}_2\text{O}$

29. $2\text{BN} + 3 \text{F}_2 \rightarrow 2 \text{BF}_3 + \text{N}_2$

30. $\text{CaSO}_4 + 2\text{H}_2\text{O} + 2 \text{SO}_3 \rightarrow \text{CaSO}_4 + 2 \text{H}_2\text{SO}_4$

31. $4\text{C}_3\text{H}_7\text{N}_2\text{O}_7 + 5 \text{O}_2 \rightarrow 12 \text{CO}_2 + 14 \text{H}_2\text{O} + 4 \text{N}_2$

32. $\text{C}_7\text{H}_6\text{O}_4\text{S}_2 + 11 \text{O}_2 \rightarrow 7 \text{CO}_2 + 8 \text{H}_2\text{O} + 2 \text{SO}_2$

33. $9\text{Na} + 4 \text{Zn}_2 \rightarrow 8 \text{NaI} + \text{NaZn}_4$

34. $\text{HBrO}_3 + 5 \text{HBr} \rightarrow 3 \text{H}_2\text{O} + 3 \text{Br}_2$

35. $\text{Al}_4\text{C}_3 + 12 \text{H}_2\text{O} \rightarrow 4 \text{Al(OH)}_3 + 3 \text{CH}_4$
Example 3: Balance \( \text{C}_6\text{H}_8\text{N}_3\text{O}_7 \rightarrow \text{N}_2 \)… start with carbon

\[ 3\text{C}_6\text{H}_8\text{N}_3\text{O}_7 + 7\text{O}_2 \rightarrow 6\text{CO}_2 + 12\text{H}_2\text{O} + 3\text{N}_2 \]

need only \( \text{N}_2 \)

\[ \text{O}_2 \] total

Example 4: Balance \( \text{C}_1\text{H}_4\text{N}_3\text{O}_7 \rightarrow \text{CO}_2 + \text{H}_2\text{O} + \text{N}_2 \)

\[ 4\text{C}_1\text{H}_4\text{N}_3\text{O}_7 + 8\text{O}_2 \rightarrow 7\text{CO}_2 + 3\text{H}_2\text{O} + 2\text{N}_2 \]

Total \( \text{C} = 12 \) \( \text{H} = 12 \) \( \text{O} = 46 \) \( \text{N} = 6 \)

Total \( \text{C} = 12 \) \( \text{H} = 12 \) \( \text{O} = 46 \) \( \text{N} = 6 \)

alternative method: using the algebra method

**Chemistry Homework**

Assignment #3: Balancing Equations

Worksheet #29-55 (odd) You may complete THIS ASSIGNMENT in your booklet on the attached pages.
Use the Algebra Method

...continued

37. \(_{\text{CH}_3\text{NO}_2} + \_\text{Cl}_2 \rightarrow \_\text{CCl}_3\text{NO}_2 + \_\text{HCl}\)

39. \(_\text{Al}_2\text{C}_6 + \_\text{H}_2\text{O} \rightarrow \_\text{Al(OH)}_3 + \_\text{C}_2\text{H}_2\)

41. \(_\text{LiH} + \_\text{AlCl}_3 \rightarrow \_\text{LiAlH}_4 + \_\text{LiCl}\)

43. \(_\text{CaSi}_2 + \_\text{SbCl}_3 \rightarrow \_\text{Si} + \_\text{Sb} + \_\text{CaCl}_2\)

45. \(_\text{NH}_3 + \_\text{O}_2 \rightarrow \_\text{NO} + \_\text{H}_2\text{O}\)

47. \(_\text{NH}_4\text{Cl} + \_\text{CaO} \rightarrow \_\text{NH}_3 + \_\text{CaCl}_2 + \_\text{H}_2\text{O}\)

49. \(_\text{Be}_2\text{C} + \_\text{H}_2\text{O} \rightarrow \_\text{Be(OH)}_2 + \_\text{CH}_4\)

51. \(_\text{NO}_2 + \_\text{H}_2\text{O} \rightarrow \_\text{HNO}_3 + \_\text{NO}\)
The last two are tough ("Mind Benders") and require you to use trial and error. Good luck!
F. Identifying and Assigning Phases in Reaction Equations

The phase of each reactant and product in a chemical reaction is indicated by writing using symbols in parentheses immediately after each formula:

What do state symbols show?

State symbols are added to a symbol equation to show whether the reactants and products are:

- **solid** – symbol is \((s)\)
- **liquid** – symbol is \((l)\)
- **gas** – symbol is \((g)\)
- **dissolved in water** – symbol is \((aq)\).

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

With state symbols in place, this symbol equation now shows that the sulfur is a solid, the oxygen is a gas and the sulfur dioxide is also a gas.

- Clues that something is a solid are found in adjectives for such as **crystal powder** and **precipitate**
  (a precipitate is a solid that forms when two liquid or aqueous solutions react with each other).
- Most elements and compounds containing metals (ionic compounds) are solids at room temperature. When dissolved in water, however, they are in the **aqueous state**.

**EXAMPLE:** Translate the following word equations into balanced chemical reaction equations:

1. Cupric sulfide reacts with oxygen to produce cuprous oxide and sulfur dioxide.

\[
2CuS(s) + \frac{5}{2}O_2(g) \rightarrow Cu_2O(s) + 2SO_2(g)
\]

\[
4CuS(s) + 5O_2(g) \rightarrow 2Cu_2O(s) + 4SO_2(g)
\]
2. Zinc bromide and silver nitrate solutions react to form a zinc nitrate solution containing silver bromide as a precipitate.

\[ \text{ZnBr}_2(aq) + 2\text{AgNO}_3(aq) \rightarrow \text{Zn(NO}_3)_2(aq) + 2\text{AgBr(s)} \]

3. Aqueous hydrochloric acid reacts with calcium carbonate crystals, producing aqueous calcium chloride, gaseous carbon dioxide and liquid water.

\[ 2\text{HCl}_{(aq)} + \text{CaCO}_3(s) \rightarrow \text{CaCl}_2(aq) + \text{H}_2\text{O}_{(l)} + \text{CO}_2(g) \]

DO NOT FORGET ABOUT "HOEBRINCI THE CLOWN" WHEN YOU ARE WRITING YOUR OWN EQUATIONS!

example: "oxygen gas" always means \(O_{2(g)}\) NEVER just \(O\)

IMPORTANT: Memorize the fact that seven of the elements form diatomic molecules.

MEMORY AID: There are 7 diatomic elements "in the shape of a seven plus one."
Look at the representation of the periodic table below to see what this phrase means.

See how the elements N, O, F, Cl, Br and I are arranged in the shape of a "7". The element "H" is by itself (this is the "plus one"). Therefore, the seven diatomic elements are \(N_2\), \(O_2\), \(F_2\), \(Cl_2\), \(Br_2\), \(I_2\) and \(H_2\).

chemistry homework

ASSIGNMENT #4 Hebden pg 113-114 Questions #57, 58, 60, 62 & 64 Complete ALL assignments on a separate piece of paper and attach to your booklet when handing in at the end of the unit. Be sure to clearly number each assignment with a heading.
## Part 2: Classifying Chemical Reactions and Predicting Products

<table>
<thead>
<tr>
<th>Reaction Type</th>
<th>Reactants</th>
<th>Products</th>
</tr>
</thead>
<tbody>
<tr>
<td>Synthesis (combination)</td>
<td>two substances</td>
<td>one substance</td>
</tr>
<tr>
<td>Decomposition</td>
<td>one substance</td>
<td>two substances</td>
</tr>
<tr>
<td>Single replacement</td>
<td>element + compound</td>
<td>new element + compound</td>
</tr>
<tr>
<td>Double replacement</td>
<td>two compounds</td>
<td>two new compounds</td>
</tr>
<tr>
<td>Neutralization</td>
<td>acid + base</td>
<td>salt + water</td>
</tr>
<tr>
<td>Combustion</td>
<td>organic compound + oxygen</td>
<td>carbon dioxide + water</td>
</tr>
</tbody>
</table>

Balance the following equations. Then use the table above to classify each as one of the major reaction types listed:

1. \( \text{Na(s)} + \text{H}_2\text{O(l)} \rightarrow \text{NaOH(aq)} + \text{H}_2\text{(g)} \)  
   - **Single replacement**
2. \( \text{Li}_2\text{O(s)} + \text{H}_2\text{O(l)} \rightarrow \text{LiOH(aq)} \)  
   - **Synthesis**
3. \( \text{C}_6\text{H}_{14}(l) + \text{O}_2(g) \rightarrow \text{CO}_2(g) + \text{H}_2\text{O(g)} \)  
   - **Combustion**
4. \( \text{HCl(aq)} + \text{Sr(OH)}_2(aq) \rightarrow \text{SrCl}_2(aq) + \text{H}_2\text{O(l)} \)  
   - **Neutralization**
5. \( \text{AlBr}_3(s) \rightarrow \text{Al(s)} + \text{Br}_2(l) \)  
   - **Decomposition**

Reactions, much like elements and compounds, can be classified according to type.

The ability to recognize and classify reactions can help us predict the products of those chemical changes. Classification can also help us predict whether a reaction is likely to occur or not.

You will be expected to be able to predict the products when given the reactants, classify the type of reaction and balance it.

### 1) Synthesis (Combination) Reactions

A synthesis (or combination) reaction must involves two or more simple substances (elements or compounds) combining to form one more complex substance.

\[
\text{Synthesis} \quad (A + B \rightarrow \text{AB})
\]

- Two elements (or simple compounds) combine to form a more complex compound.
- Use valence (assume most common form if polyvalent) to predict products...

Usually synthesis reactions are accompanied by the release of a significant amount of energy in the form of heat and/or light.

That is they are **Exothermic**. The prefix "exo" means outside, while "thermo" refers to heat.

Formation reactions sometimes require a small amount of "start-up" energy to begin. This start-up energy is known as "activation energy". The friction in striking a match provides activation energy for the exothermic reaction between the red phosphorus on the match head and the oxygen gas in the air.

The reaction is:

\[
\text{P}_4(s) + 5\text{O}_2(g) \rightarrow 2\text{P}_2\text{O}_5(g) + \text{energy}
\]
Tips for Synthesis Reactions:

- Two elements (or simple compounds) combine to form a more complex compound.
- Use valence (assume most common form if polyvalent) to predict products.
- Most common reactions of this type involve oxides of metals or non-metals and water.

**Examples**

\[ 6 \text{ Ca} + \text{ P}_4 \rightarrow 2 \text{ Ca}_3\text{P}_2 \]

\[ 2 \text{H}_2 + \text{O}_2 \rightarrow 2\text{H}_2\text{O} \]

\[ \text{SO}_3(g) + \text{H}_2\text{O}(l) \rightarrow \text{H}_2\text{SO}_4(\text{aq}) \]

Formation of sulphuric acid.

**2) Decomposition Reactions**

- A decomposition reaction involves a complex compound being broken down into 1 or more simpler substances (e.g., elements or simple compounds).

\[ \text{Cl}_2 \rightarrow \text{Cl} + \text{Na} \]

\[ \text{AB} \rightarrow \text{A} + \text{B} \]

Most decomposition reactions require a continuous source of energy (often a catalyst).

- This energy is used to break bonds between the elements of the starting material.
- Reactions that absorb energy to break bonds are called endothermic.

Decomposition reactions are commonly used in the mining industry in British Columbia to separate metals from their ores.

For example, aluminum production occurs when electric current is passed through molten aluminum oxide or bauxite ore:

\[ 2 \text{Al}_2\text{O}_3(l) \rightarrow 4 \text{Al}(s) + 3 \text{O}_2(g) \]

Explain: In most cases, decomposition reactions split into elements by electricity.

Examples:

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

\[ 2 \text{NaI} \rightarrow \text{Na}_2(g) + 3\text{I}_2(g) \]

\[ \text{H}_2\text{CO}_3(aq) \rightarrow \text{H}_2\text{O}(l) + \text{CO}_2(g) \]

(carbonic acid)

(carbonic oxide or non-metal oxide)

\[ \text{H}_2\text{O}(l) \rightarrow \text{H}_2\text{O}(l) + \text{CO}_2(g) \]

(carbonic acid)

(carbonic oxide or non-metal oxide)

\[ \text{H}_2\text{O}(l) + \text{CO}_2(g) \rightarrow \text{H}_2\text{CO}_3(aq) \]

(carbonic acid)

(carbonic oxide or non-metal oxide)
Types of Chemical Reactions Worksheet

Part 1 – Classify each of the following reactions as a synthesis (S) or decomposition (D) reaction and then balance each equation.

<table>
<thead>
<tr>
<th>Reaction</th>
<th>Reaction type</th>
</tr>
</thead>
<tbody>
<tr>
<td>1. (2\text{NH}_3 \rightarrow \text{N}_2 + 3\text{H}_2)</td>
<td>D</td>
</tr>
<tr>
<td>2. (2\text{K} + \text{Br}_2 \rightarrow 2\text{KBr})</td>
<td>S</td>
</tr>
<tr>
<td>3. (2\text{H}_2\text{O} \rightarrow 2\text{H}_2 + \text{O}_2)</td>
<td>D</td>
</tr>
<tr>
<td>4. (2\text{Al} + 3\text{Cl}_2 \rightarrow 2\text{AlCl}_3)</td>
<td>S</td>
</tr>
<tr>
<td>5. (\text{O}_2 + 2\text{Be} \rightarrow 2\text{BeO})</td>
<td>S</td>
</tr>
<tr>
<td>6. (\text{P}_4 + 6\text{F}_2 \rightarrow 4\text{PF}_3)</td>
<td>S</td>
</tr>
<tr>
<td>7. (2\text{H}_2 + \text{O}_2 \rightarrow 2\text{H}_2\text{O})</td>
<td>S</td>
</tr>
<tr>
<td>8. (2\text{KClO}_3 \rightarrow 2\text{KCl} + 3\text{O}_2)</td>
<td>D</td>
</tr>
<tr>
<td>9. (\text{S}_8 + 12\text{O}_2 \rightarrow 8\text{SO}_3)</td>
<td>S</td>
</tr>
<tr>
<td>10. (\text{Ti} + 3\text{Cl}_2 \rightarrow 2\text{TiCl}_3)</td>
<td>S</td>
</tr>
<tr>
<td>11. (\text{CO}_2 \rightarrow \text{C} + 3\text{O}_2)</td>
<td>D</td>
</tr>
<tr>
<td>12. (2\text{NaClO}_3 \rightarrow 2\text{NaCl} + 3\text{O}_2)</td>
<td>D</td>
</tr>
</tbody>
</table>

Part 2 – Complete the following synthesis and decomposition reactions. Balance!

<table>
<thead>
<tr>
<th>Reaction</th>
<th>Reaction type</th>
</tr>
</thead>
<tbody>
<tr>
<td>13. (3\text{Mg} + \text{N}_2 \rightarrow \text{Mg}_3\text{N}_2)</td>
<td>S</td>
</tr>
<tr>
<td>14. (\text{MgF}_2 \rightarrow \text{Mg} + \text{F}_2)</td>
<td>S</td>
</tr>
<tr>
<td>15. (\text{Ca}_3\text{N}_2 \rightarrow 3\text{Ca} + \text{N}_2)</td>
<td>S</td>
</tr>
<tr>
<td>16. (2\text{Al} + 3\text{F}_2 \rightarrow 2\text{AlF}_3)</td>
<td>S</td>
</tr>
<tr>
<td>17. (4\text{K} + \text{O}_2 \rightarrow 2\text{K}_2\text{O})</td>
<td>S</td>
</tr>
<tr>
<td>18. (\text{Cd} + \text{I}_2 \rightarrow \text{CdI}_2)</td>
<td>S</td>
</tr>
<tr>
<td>19. (\text{K}_2\text{O} \rightarrow 4\text{K} + \text{O}_2)</td>
<td>S</td>
</tr>
<tr>
<td>20. (\text{AuCl}_3 \rightarrow 2\text{Au} + 3\text{Cl}_2)</td>
<td>S</td>
</tr>
</tbody>
</table>
Table 4.2.2 Series of Chemical Reactivity

<table>
<thead>
<tr>
<th>Metal</th>
<th>Decreasing Activity</th>
<th>Halogens</th>
</tr>
</thead>
<tbody>
<tr>
<td>lithium</td>
<td>most reactive</td>
<td>flourine</td>
</tr>
<tr>
<td>potassium</td>
<td></td>
<td>chlorine</td>
</tr>
<tr>
<td>calcium</td>
<td></td>
<td>bromine</td>
</tr>
<tr>
<td>magnesium</td>
<td></td>
<td>iodine</td>
</tr>
<tr>
<td>zinc</td>
<td></td>
<td></td>
</tr>
<tr>
<td>aluminum</td>
<td></td>
<td></td>
</tr>
<tr>
<td>tin</td>
<td></td>
<td></td>
</tr>
<tr>
<td>lead</td>
<td></td>
<td></td>
</tr>
<tr>
<td>hydrogen</td>
<td>Hydrogen=&gt;part of H&lt;sub&gt;2&lt;/sub&gt;O (aq)</td>
<td></td>
</tr>
<tr>
<td>copper</td>
<td>least reactive</td>
<td></td>
</tr>
<tr>
<td>mercury</td>
<td></td>
<td></td>
</tr>
<tr>
<td>silver</td>
<td></td>
<td></td>
</tr>
<tr>
<td>platinum</td>
<td></td>
<td></td>
</tr>
<tr>
<td>gold</td>
<td></td>
<td></td>
</tr>
</tbody>
</table>

A single replacement reaction involves a compound and an element. The element will replace an atom in the compound of the “same type.” LP metals replace metals, non-metals replace non-metals.

**IMPORTANT TO REMEMBER:**
- A metal element always bonds with a non-metal element.
- Metals replace metals, while non-metals replace non-metals.
- Metal replacement: \( 2 \text{K(s)} + \text{Cu(NO}_3\text{)}_2(\text{aq}) \rightarrow 2 \text{KNO}_3(\text{aq}) + \text{Cu(s)} \)
- Hydrogen replacement: \( 2 \text{Na(s)} + 2 \text{H}_2\text{O(\text{aq})} \rightarrow 2 \text{NaOH(\text{aq})} + \text{H}_2(\text{g}) \)
- Non-metal replacement: \( \text{Br}_2(\text{g}) + 2 \text{Na(s)}(\text{aq}) \rightarrow 2 \text{NaBr(\text{aq})} + \text{I}_2(\text{s}) \)

**Examples**
- \( \text{CuCl}_2(\text{aq}) + \text{Al(s)} \rightarrow \text{Cu(s)} + \text{AlCl}_3(\text{aq}) \)
- \( \text{F}_2(\text{g}) + \text{MoCl}_3(\text{s}) \rightarrow \text{Cl}_2(\text{g}) + \text{MoF}_3(\text{s}) \)
- \( \text{Cu(s)} + \text{AlCl}_3(\text{aq}) \rightarrow \text{no reaction} (\text{Cu} \leq \text{Al}) \)
- \( \text{I}_2(\text{s}) + \text{SnBr}_4(\text{aq}) \rightarrow \text{no reaction} (\text{I} \leq \text{Br}) \)

Will a single replacement reaction will proceed or not?

Must compare the **chemical reactivity** of the element to that of the element it will replace in the compound Table 4.2.2.

Chemical reactivity is the tendency of a substance to undergo chemical change.

A **more reactive** element will “kick off” a **less reactive** one.

Will the following reactions proceed? Predict the products for those that do proceed and balance the equations.

1. \( 3 \text{Na(s)} + \text{AlCl}_3(\text{aq}) \rightarrow 3 \text{NaCl(\text{aq})} + \text{Al(s)} \)
2. \( \text{Cu(s)} + \text{KBr(\text{aq})} \rightarrow \text{no react (\text{Cu} > \text{K})} \)
3. \( \text{F}_2(\text{g}) + 2 \text{LiI(\text{aq})} \rightarrow 2 \text{LiF(\text{aq})} + \text{I}_2(\text{g}) \)
4. \( \text{Ca(s)} + 2 \text{H}_2\text{O(\text{l})} \rightarrow \text{Ca(OH)}_2(\text{aq}) + \text{H}_2(\text{g}) \)

\( ^* \) Hydrogen may be displaced from most acids by all metals above it in the series. However, it may only be displaced from water (at room temperature) by those above magnesium.
A double replacement reaction involves two soluble salts (aq) or complex compounds. The ions will "trade partners" and new compounds are formed (at least one must be a solid...otherwise it's no reaction).

REMEMBER: an ionic compound is made up of a positively charged, cation, bonded to a negatively charged, anion.

When these ions trade positions in their compounds, a new set of compounds is formed.

A and C are Cations (Positive Ions)  B and D are Anions (Negative Ions)

A (aq) + C (aq) → B (s) + D (aq)

1. Precipitation Reaction
   - ions "trade partners"
   - two ionic compounds (aq) + (s) are formed
   - at least one of which is a low solubility salt/compound, giving immediate evidence of chemical change as it forms a precipitate suspended in solution
   - this is called a "precipitate" and separate into ions.

Remember that EVERY salt dissolves to some extent in water.

Some salts dissociate a great amount and have a high molarity at saturation, while others become saturated at a very low molarity.

A 'soluble' salt has a saturation molarity greater than 0.10M, whereas a "low solubility" salt becomes saturated at a molarity lower than 0.10M.

Example:

NaCl(aq) + AgNO₃(aq) → AgCl(s) + NaNO₃(aq)

The precipitate is always indicated by a symbol (s), which indicates that a solid has been formed.

Use table to find (s) or (aq).

Figure 4.2.3 (a) Solutions of sodium chloride and silver nitrate; (b) Precipitate of silver chloride suspended in a solution of sodium nitrate.
How to use the Solubility Table:
Use your table to predict whether the following salts are soluble (and will be aqueous in solution-aq) or low solubility (will precipitate out of solution as a solid-s) and whether they form a precipitate (ppt) in water.

1. Sodium hydroxide
2. Ammonium acetate
3. Calcium sulphate
4. Lead (II) chloride
5. Potassium chloride
6. Calcium bromide
7. Potassium carbonate
8. Aluminum sulphate
9. Copper (II) chloride
10. Copper (I) chloride

SOLUBILITY OF COMMON COMPOUNDS IN WATER
The term soluble here means > 0.1 mol/L at 25°C.

<table>
<thead>
<tr>
<th>Negative Ions (Anions)</th>
<th>Positive Ions (Cations)</th>
<th>Solubility of Compounds</th>
</tr>
</thead>
<tbody>
<tr>
<td>All</td>
<td>Alkali ions: Li⁺, Na⁺, K⁺, Rb⁺, Cs⁺, Fr⁺</td>
<td>Soluble (aq)</td>
</tr>
<tr>
<td>All</td>
<td>Hydrogen ion: H⁺</td>
<td>Soluble</td>
</tr>
<tr>
<td>All</td>
<td>Ammonium ion: NH₄⁺</td>
<td>Soluble (aq)</td>
</tr>
<tr>
<td>Nitrate, NO₃⁻</td>
<td>All</td>
<td>Soluble</td>
</tr>
<tr>
<td>Chloride, Cl⁻</td>
<td>All others</td>
<td>Soluble (aq)</td>
</tr>
<tr>
<td>Bromide, Br⁻</td>
<td>Ag⁺, Pb⁺², Cu⁺</td>
<td>Copper(II)</td>
</tr>
<tr>
<td>Iodide, I⁻</td>
<td>(Ag⁺, Ca²⁺, Sr²⁺, Ba²⁺, Pb⁺²)</td>
<td>Low Solubility (s)</td>
</tr>
<tr>
<td>Sulphate, SO₄²⁻</td>
<td>All others, Al³⁺</td>
<td>Soluble (aq)</td>
</tr>
<tr>
<td>Sulphide, S²⁻</td>
<td>Alkali ions, H⁺, NH₄⁺, Be²⁺, Mg²⁺, Ca²⁺, Sr²⁺, Ba²⁺</td>
<td>Soluble</td>
</tr>
<tr>
<td>Hydroxide, OH⁻</td>
<td>All others</td>
<td>Low Solubility (s)</td>
</tr>
<tr>
<td>Phosphate, PO₄³⁻</td>
<td>Alkali ions, H⁺, NH₄⁺</td>
<td>Soluble (aq)</td>
</tr>
<tr>
<td>Carbonate, CO₃²⁻</td>
<td>All others</td>
<td>Low Solubility (s)</td>
</tr>
<tr>
<td>Sulphite, SO₃²⁻</td>
<td>All others</td>
<td>Low Solubility (s)</td>
</tr>
</tbody>
</table>
Types of Chemical Reactions Worksheet

Part 3 – Classify each of the following reactions as a single replacement (SR) or double replacement (DR) reaction and then balance each equation.

<table>
<thead>
<tr>
<th>Reaction</th>
<th>Reaction type</th>
</tr>
</thead>
<tbody>
<tr>
<td>21. ( \text{Li} + \text{AlCl}_3 \rightarrow \text{Al} + \text{LiCl} )</td>
<td>SR</td>
</tr>
<tr>
<td>22. ( \text{Zn} + \text{SnF}_4 \rightarrow \text{Sn} + \text{ZnF}_2 )</td>
<td>SR</td>
</tr>
<tr>
<td>23. ( \text{FeBr}_2 + \text{ZnSO}_4 \rightarrow \text{ZnBr}_2 + \text{FeSO}_4 )</td>
<td>DR</td>
</tr>
<tr>
<td>24. ( \text{NH}_4\text{OH} + \text{H}_2\text{CO}_3 \rightarrow \text{H}_2\text{O} + (\text{NH}_4)_2\text{CO}_3 )</td>
<td>DR</td>
</tr>
<tr>
<td>25. ( \text{Au(CN)}_3 + \text{Zn} \rightarrow \text{Au} + \text{Zn(CN)}_2 )</td>
<td>SR</td>
</tr>
<tr>
<td>26. ( \text{FeBr}_3 + \text{Zn} \rightarrow \text{ZnBr}_2 + \text{Fe} )</td>
<td>SR</td>
</tr>
<tr>
<td>27. ( \text{Ni} + \text{HCl} \rightarrow \text{NiCl}_2 + \text{H}_2 )</td>
<td>SR</td>
</tr>
<tr>
<td>28. ( \text{FeCl}_3 + \text{Na}_2\text{SO}_3 \rightarrow \text{NaCl} + \text{Fe}_2(\text{SO}_3)_3 )</td>
<td>DR</td>
</tr>
<tr>
<td>29. ( \text{Al}_2(\text{SO}_4)_3 + \text{Na}_2\text{PO}_4 \rightarrow \text{Na}_2\text{SO}_4 + \text{AlPO}_4 )</td>
<td>DR</td>
</tr>
<tr>
<td>30. ( \text{Al} + \text{Fe}_2\text{O}_3 \rightarrow \text{Fe} + \text{Al}_2\text{O}_3 )</td>
<td>SR</td>
</tr>
<tr>
<td>31. ( (\text{NH}_4)_2\text{S} + \text{Mn(NO}_3)_2 \rightarrow \text{NH}_4\text{NO}_3 + \text{MnS} )</td>
<td>DR</td>
</tr>
<tr>
<td>32. ( \text{H}_3\text{PO}_4 + \text{Cu(OH)}_2 \rightarrow \text{H}_2\text{O} + \text{Cu}_3(\text{PO}_4)_2 )</td>
<td>DR</td>
</tr>
</tbody>
</table>

Part 4 – Complete the following single and double replacement reactions. Balance!

<table>
<thead>
<tr>
<th>Reaction</th>
<th></th>
</tr>
</thead>
<tbody>
<tr>
<td>33. ( \text{PbCl}_4 + \text{Al} \rightarrow 4\text{AlCl}_3 + 3\text{Pb} )</td>
<td></td>
</tr>
<tr>
<td>34. ( \text{Na} + \text{Cu}_2\text{O} \rightarrow \text{Na}_2\text{O} + 2\text{Cu} )</td>
<td></td>
</tr>
<tr>
<td>35. ( \text{CaS} + \text{NaOH} \rightarrow \text{Ca(OH)}_2 + \text{Na}_2\text{S} )</td>
<td></td>
</tr>
<tr>
<td>36. ( \text{CuF}_2 + \text{Mg} \rightarrow \text{MgF}_2 + \text{Cu} )</td>
<td></td>
</tr>
<tr>
<td>37. ( \text{K}_3\text{PO}_4 + \text{MgI}_2 \rightarrow 6\text{KI} + \text{Mg}_3(\text{PO}_4)_2 )</td>
<td></td>
</tr>
<tr>
<td>38. ( \text{SrCl}_2 + \text{Pb(NO}_3)_2 \rightarrow \text{Sr(NO}_3)_2 + \text{PbCl}_2 )</td>
<td></td>
</tr>
<tr>
<td>39. ( \text{Cl}_2 + \text{CsBr} \rightarrow 2\text{CsCl} + \text{Br}_2 )</td>
<td></td>
</tr>
<tr>
<td>40. ( \text{AlCl}_3 + \text{CuNO}_3 \rightarrow \text{Al(NO}_3)_3 + 3\text{CuCl}_2 )</td>
<td></td>
</tr>
</tbody>
</table>
Suppose you wanted to make a saturated solution of PbI₂. One way you could do this is to dissolve (dissociate) PbI₂ in water (making Pb²⁺ (aq) and I⁻(aq)) until no more will dissolve and you have excess PbI₂(aq) on the bottom.

Another way is to mix one solution that has Pb²⁺ (aq) ions to another solution that has I⁻(aq) ions...

WATCH THIS:
http://www.mhhe.com/physci/chemistry/animations/chung_7e_esd/crm3s2_3.swf

Let’s suppose you decided to mix equal volumes of two soluble salt solutions together:

such as 0.20M KI(aq) with 0.20M Pb(NO₃)₂(aq)

KI(aq) is actually K⁺ and I⁻(aq)

Pb(NO₃)₂(aq) is actually Pb²⁺ and NO₃⁻(aq)

By mixing, you’ve introduced Pb²⁺ to I⁻ and also K⁺ to NO₃⁻.

If either of these combinations are ‘unsuitable’ together, they will be ‘oversaturated’ and precipitate out of solution (form a solid).

The precipitate will be:
PbI₂(s)

creating a saturated solution of PbI₂.

In this case, K⁺ and NO₃⁻ are ‘insoluble’ meaning they do not participate directly in the reaction.

---

2. Neutralization

is a reaction between an acid + base that always forms a salt (ionic compound) + water:

+ CO₂(g) when there is a non-nitride base eg. K₂CO₃

+ NH₃(g) when there is NH₄⁺ in the acid. eg. NH₄Br

The salt is composed of the cation from the base and the anion from the acid.

The hydrogen ion from the acid and the hydroxide ion from the base combine to form water (H₂O)

Some products of double replacement reactions are not very stable and spontaneously decompose to form water and a gas.

- Carbonic acid decomposes:

  H₂CO₃ decays (immediately) to form H₂O(g) + CO₂(g)

- Ammonium hydroxide decomposes:

  NH₄OH decomposes (immediately) to form H₂O(l) + NH₃(g)

If CO₂ gas is a product, after an acid is added, it is likely that the reactant compound contained,

If NH₃ gas is a product, following addition of a base to a compound it is likely that the reactant compound contained

---

Gas Formation

---

If CO₂ gas is a product, after an acid is added, it is likely that the reactant compound contained,  

If NH₃ gas is a product, following addition of a base to a compound it is likely that the reactant compound contained
5) Combustion Reactions

Combustion reactions are **exothermic** and release a significant amount of energy in the form of **heat**, **sound**, and even **light** (most commonly heat).

Generally the combustion is rapid and involves the burning of an **organic compound** in atmospheric oxygen.

Water will be released as **vapour** (H₂O(g)) (water vapour)

The combustion of a variety of **hydrocarbons** such as propane, fuel oil, and natural gas provides most of the energy for our homes.

The following combustion of octane in gasoline provides the energy to move most of our vehicles:

\[ 2 \text{C}_8\text{H}_{18(l)} + 25 \text{O}_2(g) \rightarrow 16 \text{CO}_2(g) + 18 \text{H}_2\text{O}(g) \]

Slow combustion, sometimes referred to simply as oxidation, occurs in the cells of our body to produce energy.

One of the most common examples is this reaction of the simple sugar glucose with oxygen:

\[ \text{C}_6\text{H}_{12}\text{O}_6(s) + 6 \text{O}_2(g) \rightarrow 6 \text{CO}_2(g) + 6 \text{H}_2\text{O}(l) \]

When predicting products, there are two possible cases:

- a) If **hydrocarbon** only contains C, H (and possibly O), then products are just CO₂ and H₂O.

  **H₂O. Example:** \[ \text{C}_5\text{H}_2\text{O}_4 + 7 \text{O}_2 \rightarrow 5 \text{CO}_2(g) + 6 \text{H}_2\text{O}(g) \]

- b) If S is also present, then **sulphur dioxide** (SO₂) is produced along with CO₂ & H₂O.

  **Example:** \[ \text{C}_6\text{H}_5\text{S} + \text{O}_2 \rightarrow \text{CO}_2(g) + \text{H}_2\text{O}(g) + \text{SO}_2(g) \]

- c) Sulfur-containing hydrocarbons are said to be "dirty hydrocarbons" because their combustion releases **SO₂**, one of the major chemical species that produces **acid rain**, which is harmful to the environment and damages manmade structures, too.
### Part 5

Identify each of the following reactions as synthesis (S), decomposition (D), single replacement (SR), double replacement (DR), acid-base neutralization (N), or combustion (C), and balance the equation.

<table>
<thead>
<tr>
<th>Reaction</th>
<th>Reaction type</th>
</tr>
</thead>
<tbody>
<tr>
<td>41. ( ___ S_8 + 12 ___ O_2 \rightarrow 8 ___ SO_3 )</td>
<td>S</td>
</tr>
<tr>
<td>42. ( ___ (NH_4)_2CO_3 + ___ Ca(NO_3)_2 \rightarrow ___ NH_4NO_3 + ___ CaCO_3 )</td>
<td>DR</td>
</tr>
<tr>
<td>43. ( ___ N_2 + 3 ___ Zn \rightarrow ___ Zn_3N_2 )</td>
<td>S</td>
</tr>
<tr>
<td>44. ( ___ C_4H_8 + 6 ___ O_2 \rightarrow 4 ___ CO_2 + 4 ___ H_2O )</td>
<td>C</td>
</tr>
<tr>
<td>45. ( ___ Pb(NO_3)_2 + 2 ___ KI \rightarrow ___ PbI_2 + 2 ___ KNO_3 )</td>
<td>DR</td>
</tr>
<tr>
<td>46. ( ___ Zn + 2 ___ HCl \rightarrow ___ ZnCl_2 + ___ H_2 )</td>
<td>SR</td>
</tr>
<tr>
<td>47. ( ___ H_2SO_4 + 2 ___ NaOH \rightarrow ___ Na_2SO_4 + ___ H_2O )</td>
<td>N</td>
</tr>
<tr>
<td>48. ( ___ HF \rightarrow ___ H_2 + ___ F_2 )</td>
<td>D</td>
</tr>
<tr>
<td>49. ( ___ Au(NO_3)_3 + 3 ___ Cu \rightarrow ___ Au + ___ Cu(NO_3)_2 )</td>
<td>SR</td>
</tr>
</tbody>
</table>

### Part 6

Complete and balance the following reactions.

50. \( 6 \_\_\_ Na + \_\_\_ N_2 \rightarrow 3 \_\_\_ Na_3N \) (s)

51. \( 2 \_\_\_ AlF_3 \rightarrow 2 \_\_\_ Al(l) + 3 \_\_\_ F_2(g) \) (s)

52. \( 3 \_\_\_ CuSO_4 + 2 \_\_\_ Al \rightarrow \_\_\_ Al_2(SO_4)_3 + 3 \_\_\_ Cu(l) \) (aq)

53. \( \_\_\_ CaCl_2 + \_\_\_ Pb(NO_3)_2 \rightarrow \_\_\_ Ca(NO_3)_2 + \_\_\_ PbCl_2 \) (s)

54. \( 2 \_\_\_ CaH_10 + 13 \_\_\_ O_2 \rightarrow 8 \_\_\_ CO_2(g) + 10 \_\_\_ H_2O(g) \) (s)

55. \( \_\_\_ HCl + \_\_\_ NaOH \rightarrow \_\_\_ H_2O(l) + \_\_\_ NaCl(aq) \) (aq)
Part 7 – Identify which reaction type or types match the following descriptions:

56. There is only one reactant.

57. There is only one product.

58. The reactants are an acid and a base.

59. The products are an element and a compound.

60. The products are carbon dioxide and water.

61. Both reactants are compounds.

62. One reactant is an element. The other is a compound.

---

**Decomposition**

**Synthesis**

**Neutralization**

**Single Replacement**

**Combustion**

**Double Replacement OR Neutralization**

**Single Replacement OR Combustion**

Part 8 – Write a balanced equation for each of the following reactions. Include phases!

63. sodium + oxygen → ?

\[4 \text{Na}(s) + \text{O}_2(g) \rightarrow 2 \text{Na}_2\text{O}(s)\]

64. sodium sulfate + calcium chloride → ?

\[\text{Na}_2\text{SO}_4(aq) + \text{CaCl}_2(aq) \rightarrow 2 \text{NaCl}(aq) + \text{CaSO}_4(s)\]

65. propane (C\(_3\)H\(_8\)) + oxygen → ?

\[\text{C}_3\text{H}_8(s) + 5 \text{O}_2(g) \rightarrow 3 \text{CO}_2(g) + 4 \text{H}_2\text{O}(g)\]

66. sulfuric acid + potassium hydroxide → ?

\[\text{H}_2\text{SO}_4(aq) + \text{KOH}(aq) \rightarrow 2\text{H}_2\text{O}(l) + \text{K}_2\text{SO}_4(aq)\]

67. ? → aluminum + chlorine

\[2\text{AlCl}_3(s) \rightarrow 2\text{Al}(s) + 3\text{Cl}_2(g)\]

68. ? → cadmium nitrate + rubidium

\[2\text{RbNO}_3(aq) + \text{Cd}(s) \rightarrow \text{Cd(NO}_3)_2(aq) + 2\text{Rb}(s)\]

69. ? → potassium chloride

\[\text{K}(s) + \text{Cl}_2(g) \rightarrow 2\text{KCl}(s)\]
The Energy of Chemical Bonds

Almost all energy on which we rely comes from chemical reactions. Energy is released from our food, from fuels for heating and transportation, and when the chemical reactions in batteries power our portable devices.

In any chemical reaction:
1. reactants change into products
2. a change in energy occurs.

Exothermic Vs. Endothermic

As you know from Science 10, there are two kinds of energy changes in chemical reactions:
- In an endothermic reaction, energy is absorbed by the system from the surroundings. (feel cold)
- In an exothermic reaction, energy is released from the system to the surroundings. (feel warm)  → produced

Endothermic reactions: Heat is absorbed.
1) photosynthesis: Plants absorb heat energy from sunlight to convert carbon dioxide and water into glucose and oxygen.
2) cooking: Heat energy is absorbed from the pan to cook the egg.

Exothermic reactions: Heat is released.
1) combustion: The burning of carbon-containing compounds uses oxygen, from air, and produces carbon dioxide, water, and lots of heat. For example, $\text{CH}_4 + 2\text{O}_2 \rightarrow \text{CO}_2 + 2\text{H}_2\text{O} + \text{heat}$

Chemists experiment on chemical systems containing reactants and products which exchange energy with the surroundings - the container and the rest of the universe.

The First Law of Thermodynamics states that:

Energy can neither be created or destroyed

This simple statement means that any energy lost by a system must simultaneously be gained by the surroundings (or vice versa).

**Bond Energy**

All molecules and compounds possess bond energy. This chemical potential energy is the energy of intramolecular bonds. These are bonds that are formed between atoms within a molecule. Weaker bonds exist between molecules in a sample of solid, liquid, and even gaseous matter. These weak bonds hold the molecules of a solid or liquid together. These weak interactions between molecules are called intermolecular forces. The details of intermolecular forces relate to the polarity or lack of polarity of a molecule. (+ or “side”)

- The difference between the potential bond energy of reactants and products before and after a chemical or physical change is known as the change or enthalpy value.
Why is heat released or absorbed in a chemical reaction?

In any chemical reaction, chemical bonds are either broken or formed.

**Rule of thumb is:**

*When chemical bonds are formed, heat is released, and when chemical bonds are broken, heat is absorbed.*

Molecules want to stay together, so formation of chemical bonds between molecules requires less energy as compared to breaking bonds between molecules, which requires MORE energy and results in heat being absorbed from the surroundings.

1. **Energy is REQUIRED** to break the bonds between the atoms in the reactants. 
   
   *... and immediately afterward...*

2. **Energy is RELEASED** as the new bonds form between the atoms in the products.

**Summarizing:**

- Bond breaking is always endothermic.
- Bond forming is always exothermic.

The reaction is either endothermic or exothermic depending on which of these is greater.

**Endothermic Reaction:** Total energy **absorbed** in bond breaking > Total energy **released** during bond forming.

**Exothermic Reaction:** Total energy **absorbed** in bond breaking < Total energy **released** during bond forming.

**Enthalpy ΔH**

The amount of energy stored in the bonds of the reactants or products in a system is called the **Enthalpy** (H) (from the Greek word enthalpein meaning “to warm”).

Since energy will either be **lost or gained** by the system during a reaction, the value of H will **always be different** between the reactants and the products.

In other words, there is a **change in energy**.

- In an endothermic reaction, more energy will be stored in the products than in the reactants:
  
  \[
  \Delta H \text{ has entered the system (required) } H_{\text{reactants}} < H_{\text{products}}
  \]

- In an exothermic reaction, less energy will be stored in the products than in the reactants:
  
  \[
  \Delta H \text{ has exited the system (released) } H_{\text{reactants}} > H_{\text{products}}
  \]

We can never really know the internal energy in a system but we **can measure the change in this energy**.

This **change in energy** is represented by ΔH where:

\[
\Delta H = H_{\text{products}} - H_{\text{reactants}}
\]

**ΔH value negative** -> energy released -> **exothermic reaction**

**ΔH value positive** -> energy absorbed -> **endothermic reaction**
Potential Energy Diagrams

In a chemical reaction, some bonds are broken and some bonds are formed. During the reaction, there is an intermediate stage, where chemical bonds are partially broken and partially formed. This intermediate exists at a higher energy level than the starting reactants; it is very unstable and is referred to as the "transition state" (activated complex). The energy required to reach this transition state is called activation energy (Eₐ).

We can define activation energy as:

\[ \Delta H = \text{Enthalpy change} \]

An energy diagram shows the relative potential energies of reactants, transition states, and products as a reaction progresses.

Can calculate the Eₐ and ΔH for any reaction from its potential energy diagram.

The activation energy (Eₐ) is the difference in the energy between the transition state and the reactants.

The enthalpy change (ΔH) is the difference in the energy between the reactants and the products. (±)

Endothermic Reaction

- The reactants are at a lower energy level compared to the products.
- The products are less stable than the reactants.
- forcing the reaction in the forward direction towards more unstable species
- overall ΔH for the reaction is positive.
- energy is absorbed from the surroundings.

\[ \Delta H = H_{\text{prod.}} - H_{\text{react.}} \]
\[ \Delta H = \text{Big} # - \text{small} # = + \]
Exothermic Reaction

- The reactants are at a higher energy level compared to the products.
- The products are more stable than the reactants.
- Overall $\Delta H$ for the reaction is negative.
- Energy is released in the form of heat.

$\Delta H = +1_{\text{prod}} - +1_{\text{react}}$

$\text{Overall } \Delta H = \text{small #} - \text{big #} = (-)$

Representing Energy Changes within Chemical Reaction Equations

- Enthalpy has units of $\text{joules} (J) \Rightarrow \text{or KJ}$
- Balanced reaction equations that include the enthalpy change are known as thermochemical equations.
- Enthalpy is an extensive property (the energy lost or gained depends on reactant amounts)
- There are two ways to write them, the first shown being the preferred way:

1. Writing the enthalpy change immediately after the equation - using the sign of $\Delta H$ to indicate whether the change is endothermic or exothermic.

   **Exothermic Example:** $2 \text{C}_8\text{H}_18 + 25 \text{O}_2 \rightarrow 16 \text{CO}_2 + 18 \text{H}_2\text{O}; \Delta H = -10,992 \text{kJ}$

   **Endothermic Example:** $6 \text{CO}_2 + 6 \text{H}_2\text{O} \rightarrow \text{C}_6\text{H}_12\text{O}_6 + 6 \text{O}_2; \Delta H = +393 \text{kJ}$

2. Writing the heat term within the chemical equation - using the side to indicate whether the change is endothermic or exothermic.

   **Exothermic Example:** $2\text{C}_8\text{H}_18 + 25\text{O}_2 \rightarrow 16 \text{CO}_2 + 18 \text{H}_2\text{O} + 10,992 \text{kJ}$

   **Endothermic Example:** $6 \text{CO}_2 + 6 \text{H}_2\text{O} + \frac{393 \text{kJ}}{\text{written on reactants side}} \rightarrow \text{C}_6\text{H}_12\text{O}_6 + 6 \text{O}_2$

Assignment #8: Hebben pg 120-122 Questions # 68-80
Complete ALL assignments on a separate piece of paper and attach to your booklet. Clearly number each assignment with a heading.