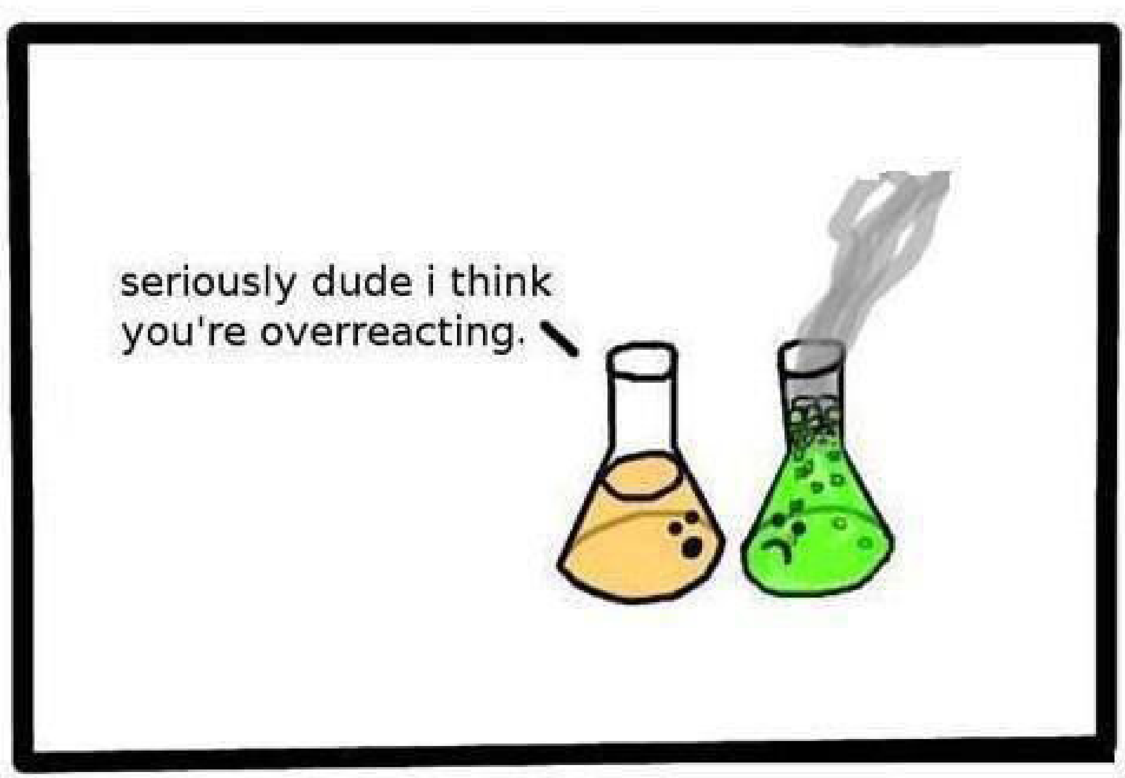
CHEMISTRY 11

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

Cupric- Cu^{2+}

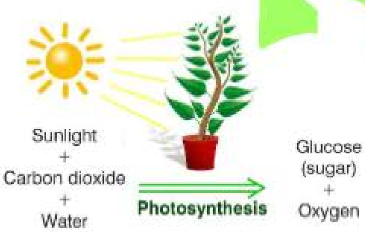
$+O_2$

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.

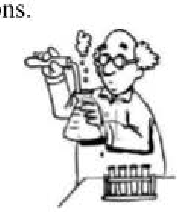
Chemical reactions are continuously occurring all around and within you. In fact, it is a series of chemical reactions within your brain that are allowing you to see and understand the words you are reading on this page. In this unit, we will study the fundamentals of chemical change, develop a method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

Chemical Reactions Run Our World



Chemistry is the study of matter and its changes. Earlier this year, you learned how physical and chemical changes differ. Chemical changes 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 *(with new and different properties)* called products.
2. A change in energy. occurs *(e.g. emission or absorption of heat, light, electricity, etc.)*

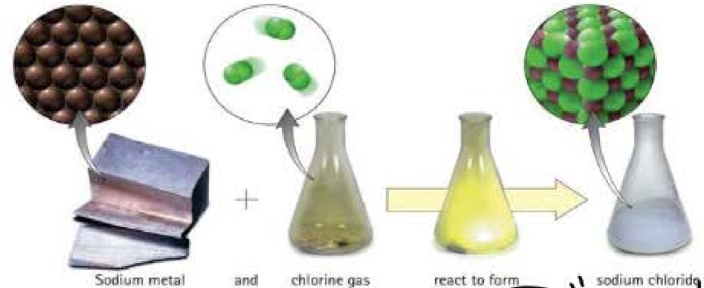
Other **evidence of a chemical reaction**:

- different colored 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: **Reactant(s)** $\xrightarrow{\text{rxn occurs}}$ **Product(s)** "reaction"

NOT Balanced.

Chemical Word Equation: "Sodium metal reacts with chlorine gas to produce sodium chloride crystals." (s) + ionic compound.

Skeleton chemical reaction equation: $Na_{(s)} + Cl_{2(g)} \rightarrow NaCl_{(s)}$

Balanced chemical reaction equation: $2Na_{(s)} + Cl_{2(g)} \rightarrow 2NaCl_{(s)}$

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 ___aqueous___, meaning the substance is dissolved in ___water (H_2O)___ (solid - ionic comp.)

Word equation: mercury(II) oxide → mercury + oxygen (Hg^{2+} O^{2-})

Formula equation: $2HgO_{(s)} \rightarrow 2Hg_{(l)} + O_{2(g)}$

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 ___colour change___
- evolution of a ___gas___ (bubbles may be visible)
- formation of a solid from two solutions. Such a solid is called a ___precipitate___. (aq) + (aq) → (s) + (aq) . . . .

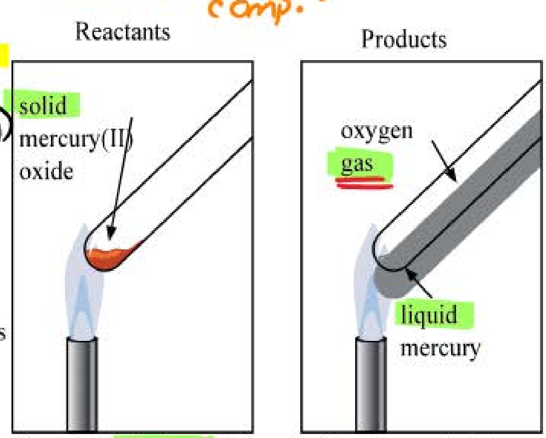


Figure 4.1.1 Heating solid mercury(II) oxide results in the formation of liquid mercury and oxygen gas.

PRACTICE

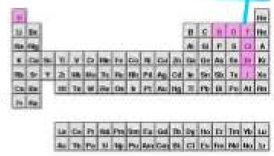
1. Convert the following from a *word equation* to a **formula equation**:
   Calcium oxide powder is combined with water to form calcium hydroxide solid.
   $CaO_{(s)} + H_2O_{(l)} \rightarrow Ca(OH)_{2(s)}$

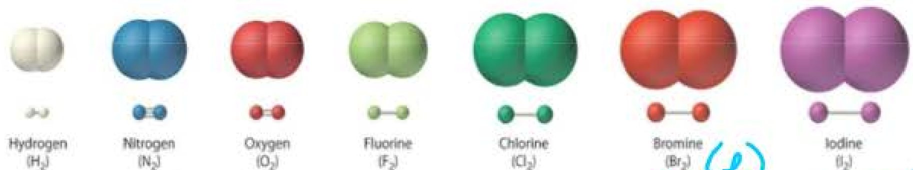
2. Convert the following from a *formula equation* to a **word equation**:
   $H_2CO_3(aq) \rightarrow CO_2(g) + H_2O(l)$
   carbonic acid → carbon dioxide gas + water

REMEMBER! The HOFBrINCl Elements:

+1 ← "7"

- There are 7 naturally occurring elements that, when found in nature, exist as diatomic molecules.
- The atoms will not exist alone, they **bond** to each other.
- This means they must ***always*** be written with a subscript of 2.





(l) [at room temp]

C. Chemical Systems

- A ***system*** is the part of the universe you are observing.
- It is a closed if nothing can enter or escape.
- It is an open is something can enter or escape.

with respect to mass, volume + energy.

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 $H_2O_{(l)} \longrightarrow H_{2(g)} + O_{2(g)}$

– 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 closed with respect to mass but open with respect to heat. Why?

heat energy will be transferred and "lost" through the bottle; however all H_2O molecules remain within.

D. The Conservation Laws

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

A Conservation Law

– is an experimentally observed and determined law

– tells you what is conserved 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***: total mass before = total mass after
2. ***Law of Conservation of Atoms***: number + type of atoms are conserved. *
3. ***Law of Conservation of Charge***: electric charge is conserved.
4. ***Law of Conservation of Energy***: total energy in a closed system does not change (the type of energy may change; total is conserved)

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 assignmentwith 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.

skeleton equation

A brief examination of the equation for the decomposition of mercury(II) oxide, $HgO_{(s)} \rightarrow Hg_{(l)} + O_{2(g)}$, shows that it **does not obey** the law of conservation of mass.

The reactant, HgO contains LESS oxygen atom than the products, Hg and O_2.

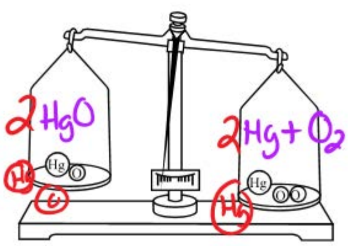


Figure 4.1.2 The reactant in this reaction weighs less than the products.
(unbalanced)

* 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.

"big numbers" in front

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

eg. $2Mg(NO_3)_2$

6×2
12=O
N=4

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.** (little numbers)
> Altering the subscripts will give an **incorrect formula** for a substance.

eg. $O_2 \rightarrow O_3$

completely different molecule

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

Example Balance: $(NH_4)_3PO_4 + NaOH \longrightarrow Na_3PO_4 + NH_3 + H_2O$

The Method

Note:

- **Until you are finished balancing, missing coefficients are treatedas 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!
  single metal atoms
- **Balance preserved groups (those that don't come apart) whenever possible.**
  polyatomic ions/groups.
  balance as a whole, never separate.
  eg. (NH_4), (PO_4), (OH)

single metal element

Balance: $(NH_4)_3PO_4 + NaOH \longrightarrow Na_3PO_4 + NH_3 + H_2O$

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.

Na, N, or PO_4

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

$\_\_(NH_4)_3PO_4 + 3NaOH \longrightarrow 1Na_3PO_4 + \_\_NH_3 + \_\_H_2O$

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.

∴ needs to be $1 \cdot PO_4$ There is $1 \times PO_4$

$1(NH_4)_3PO_4 + 3NaOH \longrightarrow 1Na_3PO_4 + \_\_NH_3 + \_\_H_2O$

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

$3 \times N$

$1(NH_4)_3PO_4 + 3NaOH \longrightarrow 1Na_3PO_4 + 3NH_3 + \_\_H_2O$

$3 \cdot N$ need $3 \times N$

Step ④ $1(NH_4)_3PO_4 + 3NaOH \longrightarrow 1Na_3PO_4 + 3NH_3 + 3H_2O$

$3 \cdot 4 = 12$ $3 \cdot H$ 15 Hydrogen

$3 \times 3 = 9$ $2 \cdot H$ $3 \cdot 2 = 6$ 11·H total. * need to be 6 H

Step ⑤ Omit coefficients of 1 in your final answer. ****Always do a check to make sure that all atoms are balanced.****

Check don't write "1"

$(NH_4)_3PO_4 + 3NaOH \longrightarrow Na_3PO_4 + 3NH_3 + 3H_2O$

O=4 O=3 O=7 O=4 O=3 O=7

✓ N=3 H=15 P=1 Na=3 N=3 H=15 P=1 Na=3

Advanced Tips:

- If you want "**n**" atoms of a polyatomic element, then multiply it by "**n/x**", where *x* is the multiplicity of that element in the molecule. + diatomics

  need subscript coefficient is a fraction

  Examples 7 oxygen atoms from a O_2 molecule? $n = 7$ and $x = 2$ Therefore use $\frac{7}{2} \times O_2$ $\frac{7}{2} \times 2 = 7$

  15 phosphorus atoms from P_4 molecule? $n = 15$ and $x = 4$ Therefore use $\frac{15}{4} \times P_4$ $\frac{15}{4} \times 4 = 15$

- 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 (ie: all coefficents will change)

multiply by the denominator to cancel the fraction.

PRACTICE

Balancing Formula Equations

$Ba(s) + Al(NO_3)_3(aq) \rightarrow Ba(NO_3)_2(aq) + Al(s)$

* DO NO_3 first, because it is the only species with subscripts... only 1 that is unbalanced!

What to Think about	How to Do It
1. Barium and aluminum must be left until last as both appear in elemental form.	$\underline{3}Ba + \underline{2}Al(NO_3)_3 \rightarrow \underline{3}Ba(NO_3)_2 + \underline{2}Al$ $2\times3=6$ $3\times2=6$ * Find a multiple of 3 and 2... 6 2×Al need 2×Al need 3×Ba 3×Ba
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.	$\Rightarrow 3Ba_{(s)} + 2Al(NO_3)_{3(aq)} \rightarrow 3Ba(NO_3)_{2(aq)} + 2Al_{(s)}$
4. Complete the balance with the barium and replace the phase indicators if required.	Ba = 3, Al = 2, NO_3 = 6 | Ba = 3, Al = 2, NO_3 = 6 ✓
5. Be sure to complete the process by checking the total number of barium, aluminum, nitrogen, and oxygen atoms on each side.	

Example 1 Balance $\underline{1}K + \underline{1}H_2O \longrightarrow \underline{1}KOH + \underline{\tfrac{1}{2}}H_2$

① metals 1st ⇒ K

② bal. O. only — need 1×O

③ 1×O — 2×H — need 2 Hydrogens Total on product side. $1 + \underline{1} = 2$

④ Fraction need subscript.

⑤ multiply All coefficients by 2 to cancel Fraction ×2

$$2K + 2H_2O \rightarrow 2KOH + H_2$$

metal.

Example 2 Balance $\underline{1}MoCl_3 + \underline{\quad}O_2 + \underline{1}AgCl \longrightarrow \underline{1}MoCl_4 + \underline{\quad}Ag_2O$

Mo, O, Al are all in just 1 place

① balance metal. 3·Cl — need 4×Cl total — 4×Cl

② balance Cl

$1MoCl_3 + \underline{\quad}O_2 + \underline{1}AgCl \longrightarrow 1MoCl_4 + \underline{\tfrac{1}{2}}Ag_2O$

treat as ∅   1×Ag   ③   only 1 Ag is needed... but there are 2.

④ ×2 to clear fraction immediately.

$2MoCl_3 + \underline{\tfrac{1}{2}}O_2 + 2AgCl \rightarrow 2MoCl_4 + \boxed{1}Ag_2O$

only need 1·O

⑤ 1×O on "products side"

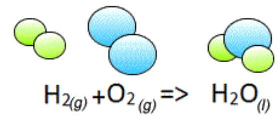
⑥ ×2 to clear the fraction

$$\boxed{\Rightarrow 4MoCl_3 + O_2 + 4AgCl \rightarrow 4MoCl_4 + 2Ag_2O}$$

chemistry homework

ASSIGNMENT #2: Balancing Equations
Worksheet #7-27 *You may complete THIS ASSIGNMENT in your booklet on the attached pages.*

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. __ Sn + __ $O_2 \longrightarrow$ __ SnO

9. __ N_2 + __ $H_2 \longrightarrow$ __ NH_3

11. __ NH_3 + __ $O_2 \longrightarrow$ __ N_2 + __ H_2O

13. __ $KNO_3 \longrightarrow$ __ KNO_2 + __ O_2

15. __ C_5H_{12} + __ $O_2 \longrightarrow$ __ CO_2 + __ H_2O

2×H 2×H

17. 2 KOH + 1 $H_2SO_4 \longrightarrow$ 1 K_2SO_4 + 2 H_2O

2×O 4×H ✓

4H

19. __ C + __ $SO_2 \longrightarrow$ __ CS_2 + __ CO

ANSWERS:

7. $2\,Sn + O_2 \longrightarrow 2\,SnO$
8. $H_2 + Cl_2 \longrightarrow 2\,HCl$
9. $N_2 + 3\,H_2 \longrightarrow 2\,NH_3$
10. $2\,Na + 2\,H_2O \longrightarrow 2\,NaOH + H_2$
11. $4\,NH_3 + 3\,O_2 \longrightarrow 2\,N_2 + 6\,H_2O$
12. $2\,C_6H_{14} + 19\,O_2 \longrightarrow 12\,CO_2 + 14\,H_2O$
13. $2\,KNO_3 \longrightarrow 2\,KNO_2 + O_2$
14. $CaC_2 + 2\,O_2 \longrightarrow Ca + 2\,CO_2$
15. $C_5H_{12} + 8\,O_2 \longrightarrow 5\,CO_2 + 6\,H_2O$
16. $K_2SO_4 + BaCl_2 \longrightarrow 2\,KCl + BaSO_4$
17. $2\,KOH + H_2SO_4 \longrightarrow K_2SO_4 + 2\,H_2O$
18. $Ca(OH)_2 + 2\,NH_4Cl \longrightarrow 2\,NH_3 + CaCl_2 + 2\,H_2O$
19. $5\,C + 2\,SO_2 \longrightarrow CS_2 + 4\,CO$

21. $\_\_\ V_2O_5 + \_\_\ Ca \longrightarrow \_\_\ CaO + \_\_\ V$

23. $\_\_\ Fe_3O_4 + \_\_\ H_2 \longrightarrow \_\_\ Fe + \_\_\ H_2O$

25. $\_\_\ Al + \_\_\ H_2SO_4 \longrightarrow \_\_\ H_2 + \_\_\ Al_2(SO_4)_3$

27. $\_\_\ NH_3 + \_\_\ O_2 \longrightarrow \_\_\ N_2H_4 + \_\_\ H_2O$

29. $\_\_\ BN + \_\_\ F_2 \longrightarrow \_\_\ BF_3 + \_\_\ N_2$

31. $\_\_\ C_3H_7N_2O_7 + \_\_\ O_2 \longrightarrow \_\_\ CO_2 + \_\_\ H_2O + \_\_\ N_2$

33. $\_\_\ Na + \_\_\ ZnI_2 \longrightarrow \_\_\ NaI + \_\_\ NaZn_4$

35. $\_\_\ Al_4C_3 + \_\_\ H_2O \longrightarrow \_\_\ Al(OH)_3 + \_\_\ CH_4$

ANSWERS

20. $Mg_3N_2 + 6\,H_2O \longrightarrow 3\,Mg(OH)_2 + 2\,NH_3$

21. $V_2O_5 + 5\,Ca \longrightarrow 5\,CaO + 2\,V$

22. $2\,Na_2O_2 + 2\,H_2O \longrightarrow 4\,NaOH + O_2$

23. $Fe_3O_4 + 4\,H_2 \longrightarrow 3\,Fe + 4\,H_2O$

24. $Cu + 2\,H_2SO_4 \longrightarrow CuSO_4 + 2\,H_2O + SO_2$

25. $2\,Al + 3\,H_2SO_4 \longrightarrow 3\,H_2 + Al_2(SO_4)_3$

26. $2\,Si_4H_{10} + 13\,O_2 \longrightarrow 8\,SiO_2 + 10\,H_2O$

27. $4\,NH_3 + O_2 \longrightarrow 2\,N_2H_4 + 2\,H_2O$

28. $2\,C_{15}H_{30} + 45\,O_2 \longrightarrow 30\,CO_2 + 30\,H_2O$

29. $2\,BN + 3\,F_2 \longrightarrow 2\,BF_3 + N_2$

30. $CaSO_4{\cdot}2H_2O + 2\,SO_3 \longrightarrow CaSO_4 + 2\,H_2SO_4$

31. $4\,C_3H_7N_2O_7 + 5\,O_2 \longrightarrow 12\,CO_2 + 14\,H_2O + 4\,N_2$

32. $C_7H_{16}O_4S_2 + 11\,O_2 \longrightarrow 7\,CO_2 + 8\,H_2O + 2\,SO_2$

33. $9\,Na + 4\,ZnI_2 \longrightarrow 8\,NaI + NaZn_4$

34. $HBrO_3 + 5\,HBr \longrightarrow 3\,H_2O + 3\,Br_2$

35. $Al_4C_3 + 12\,H_2O \longrightarrow 4\,Al(OH)_3 + 3\,CH_4$

*no metals present… start with carbon

Example 3 Balance $\boxed{\left(\underline{1}\,C_{19}H_{17}NO_3 + \underline{\emptyset}\,O_2 \longrightarrow \underline{19}\,CO_2 + \underline{\emptyset}\,H_2O + \underline{\cancel{2}\ \tfrac{1}{2}}\,N_2\right)^{③} \times 2}$

① ×2 ×2

1×N ② need only 1×N

④ ×2 ⑥ ×2

$$2C_{19}H_{17}NO_3 + \cancel{2}\ 87\,O_2 \longrightarrow 38CO_2 + \underline{17}\,H_2O + 1N_2$$

34 H total

⑤ $38(2) + 17 = 93$ Ox. total

need Oxy: $93 - 6 = 87$ more

$$4C_{19}H_{17}NO_3 + 87O_2 \longrightarrow 76CO_2 + 34H_2O + 2N_2$$

$12 + 174 = 186$ $152 + 34 = 186$ ✓

Example 4 Balance $\underline{1}\,Cr_2(SO_4)_3 + \underline{\quad}\,KI + \underline{1}\,KIO_3 + \underline{3}\,H_2O \longrightarrow \underline{2}\,Cr(OH)_3 + \underline{3}\,K_2SO_4 + \underline{\quad}\,I_2$

3×SO_4

12 O ✓

need 18 total

$18 - 12 - 3 = ③$

6×H

Total Oxy: $6 + 12 = 18$

$$1Cr_2(SO_4)_3 + \underline{5}\,KI + 1KIO_3 + 3H_2O \longrightarrow 2Cr(OH)_3 + 3K_2SO_4 + \underline{3}\,I_2$$

need 6 Total, $6 - 1 = 5$

6·K total

6 I in total

"The Algebra Method"

Alternative Method: (useful if you are having trouble, or if the reaction is very complicated… not for everyday)

Example 4 Balance $a\,Cr_2(SO_4)_3 + b\,KI + c\,KIO_3 + d\,H_2O \longrightarrow e\,Cr(OH)_3 + f\,K_2SO_4 + g\,I_2$

*let a=1

Cr: $2a = e$
S: $3a = f$
O: $12a + 3c + d = 3e + 4f$
K: $b + c = 2f$
I: $b + c = 2g$ *
H: $2d = 3e$

Reactants → products.

$a = 1$
Cr: $2(1) = e = 2$
S: $3(1) = f = 3$

H: $2d = 3(2)$
$d = \frac{3(2)}{2}$ $d = 3$

O: $12(1) + 3c + (3) = 3(2) + 4(3)$
$12 + 3c + 3 = 18$
$3c = 18 - 12 - 3$
$\frac{3c}{3} = \frac{3}{3}$ $c = 1$

K: $b + 1 = 2(3)$
$b = 6 - 1$
$b = 5$

I: $(5) + (1) = 2g$
$\frac{6}{2} = g$
$g = 3$

$a=1$, $b=5$, $c=1$, $d=3$, $e=2$, $f=3$, $g=3$

$$\underline{1}\,Cr_2(SO_4)_3 + \underline{5}\,KI + \underline{1}\,KIO_3 + \underline{3}\,H_2O \longrightarrow \underline{2}\,Cr(OH)_3 + \underline{3}\,K_2SO_4 + \underline{3}\,I_2$$

chemistry homework

ASSIGNMENT #3: Balancing Equations
Worksheet #29-55 (odd) *You may complete THIS ASSIGNMENT in your booklet on the attached pages.*

Balancing Equations

Use the Algebra Method
...continued

37. __ CH_3NO_2 + __ Cl_2 ⟶ __ CCl_3NO_2 + __ HCl

39. __ Al_2C_6 + __ H_2O ⟶ __ $Al(OH)_3$ + __ C_2H_2

41. __ LiH + __ $AlCl_3$ ⟶ __ $LiAlH_4$ + __ LiCl

43. __ $CaSi_2$ + __ $SbCl_3$ ⟶ __ Si + __ Sb + __ $CaCl_2$

45. __ NH_3 + __ O_2 ⟶ __ NO + __ H_2O

47. __ NH_4Cl + __ CaO ⟶ __ NH_3 + __ $CaCl_2$ + __ H_2O

49. __ Be_2C + __ H_2O ⟶ __ $Be(OH)_2$ + __ CH_4

51. __ NO_2 + __ H_2O ⟶ __ HNO_3 + __ NO

ANSWERS

37. $CH_3NO_2 + 3\,Cl_2 \longrightarrow CCl_3NO_2 + 3\,HCl$
38. $Ca_3(PO_4)_2 + 3\,SiO_2 + 5\,C \longrightarrow 3\,CaSiO_3 + 5\,CO + 2\,P$
39. $Al_2C_6 + 6\,H_2O \longrightarrow 2\,Al(OH)_3 + 3\,C_2H_2$
40. $2\,NaF + CaO + H_2O \longrightarrow CaF_2 + 2\,NaOH$
41. $4\,LiH + AlCl_3 \longrightarrow LiAlH_4 + 3\,LiCl$
42. $2\,CaF_2 + 2\,H_2SO_4 + SiO_2 \longrightarrow 2\,CaSO_4 + SiF_4 + 2\,H_2O$
43. $3\,CaSi_2 + 2\,SbCl_3 \longrightarrow 6\,Si + 2\,Sb + 3\,CaCl_2$
44. $2\,TiO_2 + B_4C + 3\,C \longrightarrow 2\,TiB_2 + 4\,CO$
45. $4\,NH_3 + 5\,O_2 \longrightarrow 4\,NO + 6\,H_2O$
46. $SiF_4 + 8\,NaOH \longrightarrow Na_4SiO_4 + 4\,NaF + 4\,H_2O$
47. $2\,NH_4Cl + CaO \longrightarrow 2\,NH_3 + CaCl_2 + H_2O$
48. $4\,NaPb + 4\,C_2H_5Cl \longrightarrow Pb(C_2H_5)_4 + 3\,Pb + 4\,NaCl$
49. $Be_2C + 4\,H_2O \longrightarrow 2\,Be(OH)_2 + CH_4$
50. $4\,NpF_3 + O_2 + 4\,HF \longrightarrow 4\,NpF_4 + 2\,H_2O$

51. $3\,NO_2 + H_2O \longrightarrow 2\,HNO_3 + NO$

5 ==> $3Cu + 8HNO_3 \rightarrow 3Cu(NO_3)_2 + 2NO + 4H_2O$

The last two are tough ("Mind Benders") and require you to use trial and error. Good luck!

53. $\underline{1}^{\times 3}$ Cu + $\underline{\tfrac{8}{3}}^{\times 3}$ $HNO_3 \longrightarrow$ $\underline{1}^{\times 3}$ $Cu(NO_3)_2$ + $\underline{\tfrac{2}{3}}^{\times 3}$ NO + $\underline{\tfrac{4}{3}}^{\times 3}$ H_2O) ×3 5

Cu: $a = c$

H: $b = 2e$ $\quad e = \frac{1}{2}b$

N: $b = 2c + d$ $\quad d = b - 2c$

$\boxed{O: 3b = 6c + d + e}$

1 let a = 1

∴ c = 1

2 $3b = 6c + d + e$ $\quad \boxed{c = 1}$

$3b = 6c + (b - 2c) + (\frac{1}{2}b)$

$3b - b - \frac{1}{2}b = 6(1) - 2(1)$

$\frac{6}{2}b - \frac{2}{2}b - \frac{1}{2}b = 6 - 2$

$\times 2\left(\frac{3}{2}b = 4\right)$

$\frac{3b}{3} = \frac{8}{3}$

$\boxed{b = \frac{8}{3}}$

3 $b = 2e$

(÷ 2 same a × ½)

× multiply by reciprocal

$\frac{8}{3} = \frac{2e}{2}$

$\frac{1}{2} \times \frac{8}{3} = e$

$\boxed{\frac{8}{6} = e} = \frac{4}{3}$

4 $d = b - 2c$

$d = (\frac{8}{3}) - 2(1)$

$d = \frac{8}{3} - \frac{6}{3}$

$\boxed{d = \frac{2}{3}}$

55. $\underset{a}{\underline{1}}$ $KMnO_4$ + $\underset{b}{\underline{8}}$ HBr $\longrightarrow$ $\underset{c}{\underline{1}}$ $MnBr_2$ + $\underset{d}{\underline{5/2}}$ Br_2 + $\underset{e}{\underline{1}}$ KBr + $\underset{f}{\underline{4}}$ H_2O

1 let a = 1

∴ c = 1, e = 1

K: $a = e = 1$

Mn: $a = c = 1$

$\boxed{O: 4a = f}$

$\boxed{H: b = 2f}$

Br: $b = 2c + 2d + e$

2 $4a = f$

$4(1) = f$

$4 = f$

3 $b = 2f$

$b = 2(4)$

$b = 8$

4 The only variable we don't know is "d", so we must use this formula to solve for "d"

$b = 2c + 2d + e$

$(8) = 2(1) + 2d + (1)$

$8 = 2 + 2d + 1$

$-2 \; -1 \quad -2 \quad\quad -1$

$\frac{5}{2} = \frac{2d}{2}$ $\quad$ ∴ $d = \frac{5}{2}$

* we will have to multiply the <u>entire</u> equation ×2 to cancel the fraction

5 $\left(1KMnO_4^{\times 2 = 2} + 8HBr^{\times 2 = 16} \rightarrow 1MnBr_2^{\times 2 = 2} + \frac{5}{2}Br_2^{\times 2 = 5} + 4H_2O^{\times 2 = 8}\right) \times 2$

$= 2KMnO_4 + 16HBr \rightarrow 2MnBr_2 + 5Br_2 + 8H_2O$

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)*. (s)

☝ Most elements and compounds containing metals (ionic compounds) are solids at room temperature. When dissolved in water, however, they are in the *aqueous state*.

(aq) 'solution

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

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

$Cu^{2+}S^{2-}$ $\quad O_{2(g)}$ $\quad Cu^{+}O^{2-}$ $\quad SO_{2(g)}$

2:2 $\quad$ 2:1

$$\left(\underline{2}\,CuS_{(s)} + \underline{\tfrac{5}{2}}\,O_{2(g)} \rightarrow \underline{1}\,Cu_2O_{(s)} + \underline{2}\,SO_{2(g)}\right)\times 2$$

5 Oxyg. atoms

$$\boxed{4CuS_{(s)} + 5O_{2(g)} \rightarrow 2Cu_2O_{(s)} + 4SO_{2(g)}}$$

13

PRACTICE

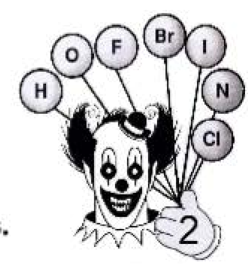
2. Zinc bromide and silver nitrate solutions react to form a zinc nitrate solution containing silver bromide as a precipitate.

$$1\,ZnBr_{2(aq)} + 2\,AgNO_{3(aq)} \rightarrow 1\,Zn(NO_3)_{2(aq)} + 2\,AgBr_{(s)}$$

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

$$2\,HCl_{(aq)} + 1\,CaCO_{3(s)} \rightarrow 1\,CaCl_{2(aq)} + 1\,H_2O_{(l)} + 1\,CO_{2(g)}$$

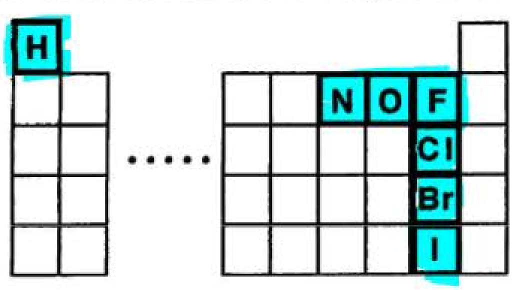
DO NOT FORGET ABOUT "HOFBrINCl THE CLOWN" WHEN YOU ARE WRITING YOUR OWN EQUATIONS!



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.**

gases → (g)

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

PRACTICE

Reaction Type	Reactants	Products
Synthesis (combination)	two substances →	one substance
Decomposition	one substance	two substances
Single replacement	element + compound	new element + compound
Double replacement	two compounds	two new compounds
Neutralization	acid + base	salt + water
Combustion	organic compound + oxygen	carbon dioxide + water

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

	Equation	Type
1.	$Na(s) + H_2O(l) \rightarrow NaOH(aq) + H_2(g)$	single replacement
2.	$Li_2O(s) + H_2O(l) \rightarrow LiOH(aq)$	synthesis
3.	$C_6H_{14}(l) + O_2(g) \rightarrow CO_2(g) + H_2O(g)$	combustion
4.	$HCl(aq) + Sr(OH)_2(aq) \rightarrow SrCl_2(aq) + H_2O(l)$	neutralization
5.	$AlBr_3(s) \rightarrow Al(s) + Br_2(l)$	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 that involves 2 or more simple substances (elements or comp.) combining to form 1 more complex substance

Synthesis (A + B → AB)

- Two elements (or simple compounds) combine to form a more complex compound.
- Use valence (e^-) (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.

bond formation releases stored energy.

Synthesis reactions sometimes require a small amount of "start-up" energy to begin. This start-up energy is known as "activation energy".

activation energy < energy released } EXO

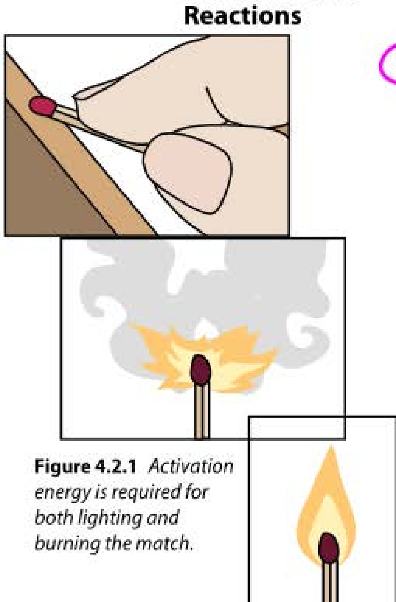
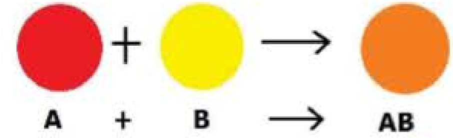


Figure 4.2.1 *Activation energy is required for both lighting and burning the match.*

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:

$$P_4(s) + 5O_2(g) \rightarrow 2P_2O_5(g) + \text{energy}$$



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*.

Ca^{2+} P^{3-}
3:2

Examples $6\,Ca + P_4 \rightarrow 2\,Ca_3P_2$

$2\,H_2 + O_2 \rightarrow 2\,H_2O$

non-metal oxide

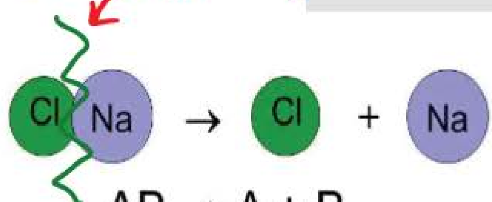
$\boxed{SO_{3(g)}} + \underbrace{H_2O_{(l)}}_{\text{water}} \rightarrow \underbrace{H_2SO_4}_{\text{acid.}}$ .... formation of sulphuric acid.

H_2 O^{2-}

2) Decomposition Reactions

A **decomposition** reaction involves a complex compound being broken down into 2 or more simpler substances (elements or simple compounds)

energy



AB → A + 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. (compound)
Reactions that absorb energy to break bonds are called to be 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:

Add energy (electric)

$2\,Al_2O_3(l) \rightarrow 4\,Al(s) + 3\,O_2(g)$ ← in most cases decomp. reactions split into elements

Examples $2\,H_2O \rightarrow 2\,H_{2(g)} + O_{2(g)}$ (water is decomposed by electricity)

$2\,NI_3 \rightarrow N_{2(g)} + 3\,I_{2(g)}$

$\underbrace{H_2CO_{3(aq)}}_{\text{carbonic acid}} \rightarrow \underbrace{H_2O_{(l)}}_{\text{water}} + \underbrace{CO_{2(g)}}_{\text{non-metal oxide}}$

synthesis.

chemistry homework

Assignment #5: Types of Chemical Reactions Worksheet Part I & II

Complete this assignment in this booklet! Show all working out!

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.

Type of Reaction	Definition	★ Equation
Synthesis	Two or more elements or compounds combine to make a more complex substance	$A + B \rightarrow AB$
Decomposition	Compounds break down into simpler substances	$AB \rightarrow A + B$
Single Replacement	Occurs when one element replaces another one in a compound	$AB + C \rightarrow AC + B$
Double Replacement	Occurs when different atoms in two different compounds trade places	$AB + CD \rightarrow AC + BD$

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

Reaction	Reaction type
1. $\underline{2}$ $NH_3 \rightarrow$ ___ N_2 + $\underline{3}$ H_2	D
2. $\underline{2}$ K + ___ $Br_2 \rightarrow$ $\underline{2}$ KBr	S
3. $\underline{2}$ $H_2O \rightarrow$ $\underline{2}$ H_2 + ___ O_2	D
4. $\underline{2}$ Al + $\underline{3}$ $Cl_2 \rightarrow$ $\underline{2}$ $AlCl_3$	S
5. ___ O_2 + $\underline{2}$ Be $\rightarrow$ $\underline{2}$ BeO	S
6. ___ P_4 + $\underline{6}$ $F_2 \rightarrow$ $\underline{4}$ PF_3	S
7. $\underline{2}$ H_2 + ___ $O_2 \rightarrow$ $\underline{2}$ H_2O	S
8. $\underline{2}$ $KClO_3 \rightarrow$ $\underline{2}$ KCl + $\underline{3}$ O_2	D
9. ___ S_8 + $\underline{12}$ $O_2 \rightarrow$ $\underline{8}$ SO_3	S
10. $\underline{2}$ Ti + $\underline{3}$ $Cl_2 \rightarrow$ $\underline{2}$ $TiCl_3$	S
11. ___ $CO_2 \rightarrow$ ___ C + ___ O_2	D
12. $\underline{2}$ $NaClO_3 \rightarrow$ $\underline{2}$ NaCl + $\underline{3}$ O_2	D

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

13.	$\underline{3}$ $Mg_{(s)}$ + ___ $N_{2(g)} \rightarrow Mg_3N_{2(s)}$
14.	___ $MgF_{2(s)} \rightarrow Mg_{(s)} + F_{2(g)}$
15.	___ $Ca_3N_{2(s)} \rightarrow 3Ca_{(s)} + N_{2(g)}$
16.	$\underline{2}$ $Al_{(s)}$ + $\underline{3}$ $F_{2(g)} \rightarrow 2\,AlF_{3(s)}$
17.	$\underline{4}$ $K_{(s)}$ + ___ $O_{2(g)} \rightarrow 2K_2O_{(s)}$
18.	___ $Cd_{(s)}$ + ___ $I_{2(s)} \rightarrow CdI_{2(s)}$ (you would have to know Cd is +2 → wouldn't be a test question)
19.	$\underline{2}$ $K_2O_{(s)} \rightarrow 4K_{(s)} + O_{2(g)}$
20.	$\underline{2}$ $AuCl_{3(s)} \rightarrow 2Au_{(s)} + 3\,Cl_{2(g)}$

compound Element

3) Single Replacement Reactions

Kicked out (displaced)

$\oplus \ominus$ + metal → metal + $\oplus \ominus$ ionic compound

AB + C → A + BC

A **single replacement** involves a compound + an element. The element will replace an atom in the compound of the "same type".
→ 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

H^+ ion behaves like a metal

- metal replacement: $2\,K(s) + Cu(NO_3)_2(aq) \rightarrow 2\,KNO_3(aq) + Cu(s)$
- hydrogen replacement: $2\,Na(s) + 2\,HOH(l) \rightarrow 2\,NaOH(aq) + H_2(g)$
- non-metal replacement: $Br_2(l) + 2\,NaI(aq) \rightarrow 2\,NaBr(aq) + I_2(s)$

Examples

$CuCl_2\,(aq) + Al\,(s) \rightarrow Cu_{(s)} + AlCl_{3(aq)}$ $Al^{3+}\ Cl^-$

$F_2\,(g) + MoCl_5\,(s) \rightarrow Cl_{2(g)} + MoF_{5(s)}$

$Cu\,(s) + AlCl_3\,(aq) \rightarrow$ no reaction (Cu < Al)

$I_2\,(s) + SnBr_4\,(aq) \rightarrow$ no reaction (I < 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. (displace)

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

1. $3\,Na(s) + AlCl_3(aq) \rightarrow 3\,NaCl_{(aq)} + Al_{(s)}$
2. $Cu(s) + KBr(aq) \rightarrow$ no rxn (Cu<K)
3. $F_2(g) + 2\,LiI(aq) \rightarrow 2\,LiF_{(aq)} + I_{2(g)}$
4. $Ca(s) + 2\,HOH(l) \rightarrow Ca(OH)_{2(aq)} + H_{2(g)}$ $H_2O_{(l)}$

Table 4.2.2 *Series of Chemical Reactivity*

Two Activity Series		
Metals	Decreasing Activity	Halogens
lithium	most reactive	flourine
potassium		chlorine
calcium		bromine
sodium		iodine
magnesium	Hydrogen => part of H_2O HOH	
aluminum		
zinc		
chromium		
iron		
nickel		
tin		
lead		
HYDROGEN*	=> part of an acid eg HCl, H_2SO_4	
copper		
mercury		
silver		
platinum		
gold	Least reactive	

* 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.

icnic compound that has been dissolved in H_2O (aq)

4) Double Replacement Reactions

A **double replacement** reaction involves 2 soluble salts (aq) or complex compounds. The ions will "trade positions" and 2 NEW compounds are formed (at least 1 must be a solid ... otherwise it's no rxn)

REMEMBER: an ionic compound is made up of a positively (+) charged, c ation 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)

one of the products must form a "precipitate"

- occur when the anions of two ionic compounds switch places
- technically, one of the two products ***must form a solid*** for the reaction to occur (otherwise, no evidence of chemical change)
- anything with alkali ions (group 1 metals (+)), ammonium ions (NH_4^+) or nitrate ions(NO_3^-) will **NOT** form a precipitate. "
- If no solid forms, just write ***aqueous solution, NR = no reaction*** = "no rxn"

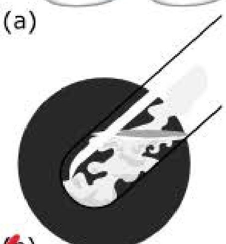
There are three categories of double replacement reactions:

1. precipitation → solid formation
2. neutralization → acid + base
3 a. gas formation → acid + base; product is gas

(a)

1. Precipitation Reaction

- ions "trade partners" ← icnic compounds (aq) + (s)
- two new salts are formed
- at least one of which is a low solubility salt/compound.
- the low solubility salt gives immediate evidence of chemical change as it forms a solid suspended in solution
- this is called a "precipitate" → separate into ions



(b)

Figure 4.2.3 (a) *Solutions of sodium chloride and silver nitrate; (b) Precipitate of silver chloride suspended in a solution of sodium nitrate*

Remember that EVERY salt dissociate 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. (low concentration)

Table 7.1.1 Solubility of Common Compounds in Water

Note: in this table, soluble means > 0.1 mol/L at 25°C.

Negative Ions (Anions)	Positive Ions (Cations)	Solubility of Compounds
All	Alkali ions: Li^+, Na^+, K^+, Rb^+, Cs^+, Fr^+	Soluble
All	Hydrogen ion: H^+	Soluble
All	Ammonium ion: NH_4^+	Soluble
Nitrate: NO_3^-	All	Soluble
Chloride: Cl^- Or Bromide: Br^- Or Iodide: I^-	All others	Soluble
	Ag^+, Pb^{2+}, Cu^+	Low solubility
Sulphate: SO_4^{2-}	All others	Soluble
	Ag^+, Ca^{2+}, Sr^{2+}, Ba^{2+}, Pb^{2+}	Low solubility
Sulphide: S^{2-}	Alkali ions, H^+, NH_4^+, Be^{2+}, Mg^{2+}, Ca^{2+}, Sr^{2+}, Ba^{2+}	Soluble
	All others	Low solubility
Hydroxide: OH^-	Alkali ions, H^+, NH_4^+, Sr^{2+}	Soluble
	All others	Low solubility
Phosphate: PO_4^{3-} Or Carbonate: CO_3^{2-} Or Sulphite: SO_3^{2-}	Alkali ions, H^+, NH_4^+	Soluble
	All others	Low solubility

SOLUBILITY OF COMMON COMPOUNDS IN WATER

The term soluble here means > 0.1 mol/L at 25°C.

low [conc]

Example:

$NaCl(aq) + AgNO_3(aq) \rightarrow$ **$AgCl(s)$** $+ NaNO_3(aq)$

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

use table to find (s) or (aq)

what is your precipitate? or is there one?

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 soid-s)** and whether they form a precipitate (ppt) in water.

1. Sodium hydroxide $NaOH_{(aq)}$
2. Ammonium acetate $NH_4CH_3COO_{(aq)}$
3. Calcium sulphate $CaSO_{4(s)}$
4. Lead (II) chloride $PbCl_{2(s)}$
5. Potassium chloride $KCl_{(aq)}$
6. Calcium bromide $CaBr_{2(aq)}$
7. Potassium carbonate $K_2CO_{3(aq)}$
8. Aluminum sulphate $Al_2(SO_4)_{3(aq)}$
9. Copper (II) chloride $CuCl_{2(aq)}$
10. Copper (I) chloride $CuCl_{(s)}$

SOLUBILITY OF COMMON COMPOUNDS IN WATER

The term soluble here means > 0.1 mol/L at 25°C.

Negative Ions (Anions)	Positive Ions (Cations)	Solubility of Compounds
All	Alkali ions: Li^+, Na^+, K^+, Rb^+, Cs^+, Fr^+	Soluble (aq)
All	Hydrogen ion: H^+	Soluble
All CH_3COO^-	Ammonium ion: NH_4^+	Soluble (aq)
Nitrate, NO_3^-	All	Soluble
Chloride, Cl^- or Bromide, Br^- or Iodide, I^-	All others Ca^{2+}, Cu^{2+} Copper(II)	Soluble (aq)
	Ag^+, Pb^{2+}, Cu^+ Copper(1)	Low Solubility
Sulphate, SO_4^{2-}	All others, Al^{3+}	Soluble
	(Ag^+, Ca^{2+}, Sr^{2+}, Ba^{2+}, Pb^{2+})	Low Solubility (s) precipitate
Sulphide, S^{2-}	Alkali ions, H^+, NH_4^+, Be^{2+}, Mg^{2+}, Ca^{2+}, Sr^{2+}, Ba^{2+}	Soluble
	All others	Low Solubility
Hydroxide, OH^-	$\boxed{\text{Alkali ions}}$ ← Na^+, H^+, NH_4^+, Sr^{2+}	Soluble (aq)
	All others	Low Solubility (s)
Phosphate, PO_4^{3-} or Carbonate, CO_3^{2-} or Sulphite, SO_3^{2-}	Alkali ions, H^+, NH_4^+	Soluble
	All others	Low Solubility

chemistry homework

Assignment #6: Types of Chemical Reactions Worksheet Part 3& 4 *Complete this assignment in this booklet! Show all working out!*

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.

	Reaction	Reaction type
21.	3 Li + ____ $AlCl_3$ → ____ Al + 3 LiCl	SR
22.	2 Zn + ____ SnF_4 → ____ Sn + 2 ZnF_2	SR
23.	____ $FeBr_2$ + ____ $ZnSO_4$ → ____ $ZnBr_2$ + ____ $FeSO_4$	DR
24.	2 NH_4OH + ____ H_2CO_3 → 2 H_2O + ____ $(NH_4)_2CO_3$	DR
25.	2 $Au(CN)_3$ + 3 Zn → 2 Au + 3 $Zn(CN)_2$	SR
26.	2 $FeBr_3$ + 3 Zn → 3 $ZnBr_2$ + 2 Fe	SR
27.	____ Ni + 2 HCl → ____ $NiCl_2$ + ____ H_2	SR
28.	2 $FeCl_3$ + 3 Na_2SO_3 → 6 NaCl + ____ $Fe_2(SO_3)_3$	DR
29.	____ $Al_2(SO_4)_3$ + 2 Na_3PO_4 → 3 Na_2SO_4 + 2 $AlPO_4$	DR
30.	2 Al + ____ Fe_2O_3 → 2 Fe + ____ Al_2O_3	SR
31.	____ $(NH_4)_2S$ + ____ $Mn(NO_3)_2$ → 2 NH_4NO_3 + ____ MnS	DR
32.	2 H_3PO_4 + 3 $Cu(OH)_2$ → 6 H_2O + ____ $Cu_3(PO_4)_2$	DR

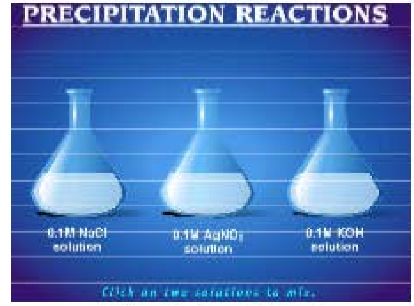
Part 4 – Complete the following single and double replacement reactions. Balance!

33.	3 $PbCl_4$ (aq) + 4 Al (s) → 4 $AlCl_3$ (aq) + 3 Pb (s)
34.	2 Na (s) + ____ Cu_2O (aq) → Na_2O (aq) + 2 Cu (s)
35.	____ CaS (aq) + 2 NaOH (aq) → $Ca(OH)_2$ (s) + Na_2S (aq)
36.	____ CuF_2 (aq) + ____ Mg (s) → MgF_2 (aq) + Cu (s)
37.	2 K_3PO_4 (aq) + 3 MgI_2 (aq) → 6 KI (aq) + $Mg_3(PO_4)_2$ (s)
38.	____ $SrCl_2$ (aq) + ____ $Pb(NO_3)_2$ (aq) → $Sr(NO_3)_2$ (aq) + $PbCl_2$ (s)
39.	____ Cl_2 (g) + 2 CsBr (aq) → 2 CsCl (aq) + Br_2 (l)
40.	____ $AlCl_3$ (aq) + 3 $CuNO_3$ (aq) (Cu^{+1}) → $Al(NO_3)_3$ (aq) + 3 CuCl (s)

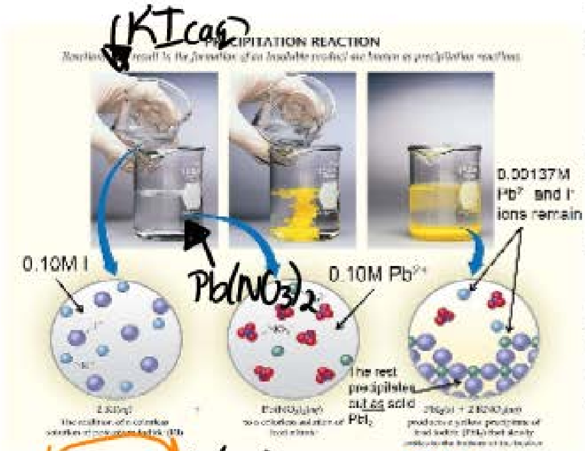
PRACTICE

Suppose you wanted to make a saturated solution of PbI_2. One way you could do this is to dissolve (dissociate) PbI_2 in water (making $Pb^{2+}_{(aq)}$ and $I_{(aq)}$) until no more will dissolve and you have excess $PbI_{2(s)}$ on the bottom.

Another way is to mix one solution that has $Pb^{2+}_{(aq)}$ ions to another solution that has $I^-_{(aq)}$ ions….



WATCH THIS:
http://www.mhhe.com/physsci/chemistry/animations/chang_7e_esp/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 $Pb(NO_3)_{2(aq)}$

$KI_{(aq)}$ is actually $K^+_{(aq)}$ and $I^-_{(aq)}$

$Pb(NO_3)_{2\,(aq)}$ is actually $Pb^{2+}_{(aq)}$ and $NO_3^-{}_{(aq)}$

By mixing, you've introduced Pb^{2+} to I^- and also K^+ to NO_3^-.

If either of these combinations are 'low solubility' together, they will be 'oversaturated' and precipitate out of solution (form a solid).

The precipitate will be: → no more can dissolve

$PbI_{2(s)}$ creating a saturated solution of PbI2.

In this case, K^+ and NO_3^- are 'spectator ions', meaning they do not participate directly in the reaction. ⟹ (aq) in solution

$$KI_{(aq)} + Pb(NO_3)_{2(aq)} \rightarrow PbI_{2(s)} + KNO_{3(aq)}$$

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

[+ $CO_{2(g)}$ when there is a non-hydroxide base eg. K_2CO_3 base)
+ $NH_{3(g)}$ when there is NH_4^+ in the acid. eg. NH_4Br]

- 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 (H_2O)
- Evidence of chemical change is *less obvious* during (when an indicator is used) a neutralization reaction, although heat is often release.d + bubbles (gas)

Example: metal hydroxide salt

$$H_2SO_4(aq) + 2\,KOH(aq) \rightarrow K_2SO_4(aq) + 2\,H_2O(l) + heat$$

H_2O

Gas Formation ③

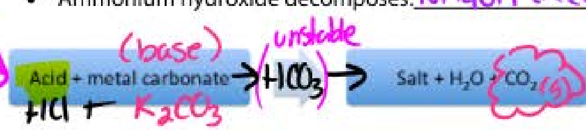
Some products of double replacement reactions are *not very stable* and spontaneously decompose to form water and a gas.

- Carbonic acid decomposes : HCO_3 decomposes (immediately) to form $H_2O_{(l)} + CO_{2(g)}$
- Ammonium hydroxide decomposes: NH_4OH decompose (immediately) to form $H_2O_{(l)} + NH_{3(g)}$

Base Acid Salt water carbon dioxide

$$K_2CO_3(aq) + 2\,HCl(aq) \rightarrow 2\,KCl(aq) + H_2O(l) + CO_2(g)$$

$$NH_4Br(aq) + NaOH(aq) \rightarrow NaBr(aq) + H_2O(l) + NH_3(g)$$



If CO_2 **gas** is a product, after an **acid** is added, it is likely that the reactant compound contained base carbonate CO_3^{2-}

If NH_3 **gas** is a product, following addition of a **base** to a compound it is likely that the reactant compound contained________

$$NH_4Br_{(aq)} + NaOH_{(aq)} \rightarrow (NH_4OH) \rightarrow NaBr_{(aq)} + H_2O_{(l)} + NH_{3(g)}$$

acid base decomposes immediately. salt water gas (ammonia)

Acid + metal carbonate → Salt + H_2O + CO_2

HCl + K_2CO_3

base acid salt water additional product.

$$K_2CO_3(aq) + 2\,HCl(aq) \rightarrow 2\,KCl(aq) + H_2O(l) + CO_2(g)$$

$$NH_4Br(aq) + NaOH(aq) \rightarrow NaBr(aq) + H_2O(l) + NH_3(g)$$

Acid base salt

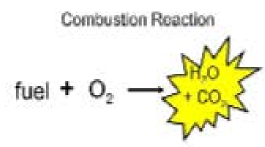
*If CO_2 **gas** is a product, after an **acid** is added, it is likely that the reactant compound contained carbonic acid (HCO_3)

If NH_3 **gas** is a product, following addition of a **base** to a compound it is likely that the reactant compound contained ammonium ion (NH_4^+)

$$K_2CO_3 + HCl \rightarrow \boxed{HCO_3} + KCl \longrightarrow KCl + \boxed{H_2O + CO_2}$$

carbonic acid decompose.

5) Combustion Reactions



A reaction of a fuel with oxygen, producing energy in the form of heat and/or light

A **combustion reaction** involve the reaction of a hydrocarbon (C_nH_n) or a carbohydrate ($C_nH_nO_n$) with oxygen gas (O_2) to always produce $CO_{2(g)}$ and H_2O.

HydroCarbon Carbon Dioxide

$$CH + O_2 \rightarrow CO_2 + H_2O$$

Oxygen Water

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. ($O_{2(g)}$)

Water will be released as vapour. ($H_2O_{(g)}$) (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:

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

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:

$$C_6H_{12}O_6(s) + 6\ O_2(g) \rightarrow 6\ CO_2(g) + 6\ H_2O(l)$$

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**. ***Example:*** $C_5H_{12}O + 7O_2 \rightarrow 5CO_{2(g)} + 6H_2O_{(g)}$

"long-chain" hydrocarbons.

b) If **S** is also present, then sulphur dioxide SO_2, is produced along with **CO2** & **H2O**. additional product.

Example: $\boxed{C_6H_{15}S} + O_2 \rightarrow CO_{2(g)} + H_2O_{(g)} + SO_{2(g)}$

sulphur dioxide ($SO_{2(g)}$) reacts with water to form sulphurous acid. $H_2O + SO_2 \rightarrow H_2SO_{3(aq)}$

☝ Sulfur-containing hydrocarbons are said to be "dirty hydrocarbons" because their combustion releases SO_2; one of the major chemical species that produces acid rain, which is harmful to the environment and damages manmade structures, too.

chemistry homework

Assignment #7: Types of Chemical Reactions
Worksheet Part 5-8
Complete this assignment in this booklet! Show all working out!

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.

Reaction	Reaction type
41. ____ S_8 + 12 O_2 → 8 SO_3	S
42. ____ $(NH_4)_2CO_3$ + ____ $Ca(NO_3)_2$ → 2 NH_4NO_3 + ____ $CaCO_3$	DR
43. ____ N_2 + 3 Zn → ____ Zn_3N_2	S
44. ____ C_4H_8 + 6 O_2 → 4 CO_2 + 4 H_2O	C
45. ____ $Pb(NO_3)_2$ + 2 KI → ____ PbI_2 + 2 KNO_3	DR
46. ____ Zn + 2 HCl → ____ $ZnCl_2$ + ____ H_2	SR
47. ____ H_2SO_4 + 2 NaOH → ____ Na_2SO_4 + 2 H_2O	N
48. 2 HF → ____ H_2 + ____ F_2	D
49. 2 $Au(NO_3)_3$ + 3 Cu → 2 Au + 3 $Cu(NO_3)_2$	SR

Part 6 – Complete and balance the following reactions.

50. 6 $Na_{(s)}$ + ____ $N_{2(g)}$ → 2 $Na_3N_{(s)}$

51. 2 $AlF_{3(s)}$ → 2 $Al_{(s)}$ + 3 $F_{2(g)}$

52. 3 $CuSO_{4(aq)}$ + 2 $Al_{(s)}$ → $Al_2(SO_4)_{3(aq)}$ + 3 $Cu_{(s)}$

53. ____ $CaI_{2(aq)}$ + ____ $Pb(NO_3)_{2(aq)}$ → $Ca(NO_3)_{2(aq)}$ + $PbI_{2(s)}$

54. 2 $C_4H_{10(s)}$ + 13 $O_{2(g)}$ → 8 $CO_{2(g)}$ + 10 $H_2O_{(g)}$

55. ____ $HCl_{(aq)}$ + ____ $NaOH_{(aq)}$ → $H_2O_{(l)}$ + $NaCl_{(aq)}$

Part 7 – Identify which reaction type or types match the following descriptions:

56. There is only one reactant. Decomposition
57. There is only one product. Synthesis
58. The reactants are an acid and a base. Neutralization
59. The products are an element and a compound. Single Replacement
60. The products are carbon dioxide and water. Combustion
61. Both reactants are compounds. Double Replacement OR Neutralization
62. One reactant is an element. The other is a compound. Single Replacement OR Combustion

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

63. sodium + oxygen → ? $4Na_{(s)} + O_{2(g)} \rightarrow 2Na_2O_{(s)}$
64. sodium sulfate + calcium chloride → ? $Na_2SO_{4(aq)} + CaCl_{2(aq)} \rightarrow 2NaCl_{(aq)} + CaSO_{4(s)}$
65. propane (C_3H_8) + oxygen → ? $C_3H_{8(s)} + 5O_{2(g)} \rightarrow 3CO_{2(g)} + 4H_2O_{(g)}$
66. sulfuric acid + potassium hydroxide → ? $H_2SO_{4(aq)} + 2KOH_{(aq)} \rightarrow 2H_2O_{(l)} + K_2SO_{4(aq)}$
67. ? → aluminum + chlorine $2AlCl_{3(s)} \rightarrow 2Al_{(s)} + 3Cl_{2(g)}$
68. ? → cadmium nitrate + rubidium $2RbNO_{3(aq)} + Cd_{(s)} \rightarrow Cd(NO_3)_{2(aq)} + 2Rb_{(s)}$
69. ? → potassium chloride $2K_{(s)} + Cl_{2(g)} \rightarrow 2KCl_{(s)}$

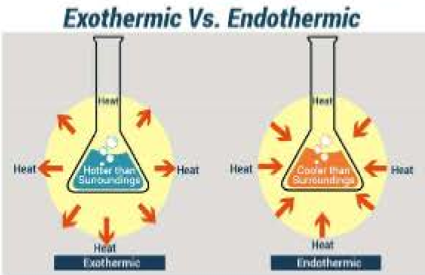
The Energy of Chemical Bonds

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.



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) required
- In an ***exothermic reaction***, energy is released from the system to the surroundings. (feel warm) produced

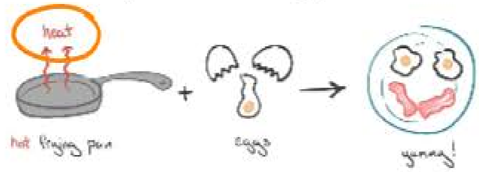
Endothermic reactions: Heat is absorbed.

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.

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

energy is a reactant



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,*

$$CH_4 + 2O_2 \rightarrow CO_2 + 2H_2O + heat$$

energy as a product

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).

(our chemical reaction)

BOND ENERGY

"stored energy"

All molecules and compounds posses bond energy. This chemical potential energy is the energy of Intra-molecular 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 inter-molecular forces.

"forces" (attractions) ≠ bonds

- 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 Enthalpy change or ΔH value.

symbol

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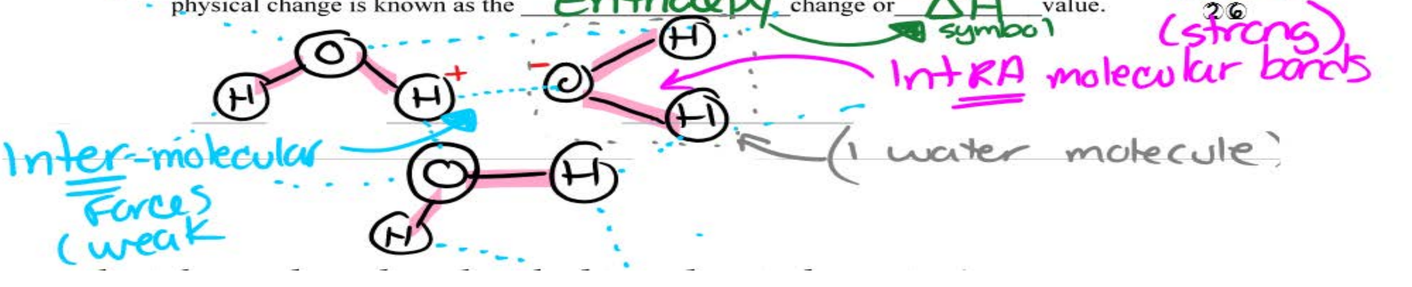


energy is required to break bonds

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.

add heat/energy

heat absorbed, chemical bonds ...

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.

heat released, chemical bonds formed

lose energy (release)

Summarizing:

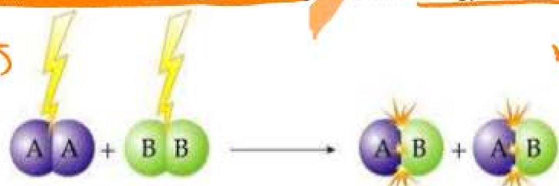
Bond breaking is always endothermic.

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.

more energy is required to break reactant bonds.

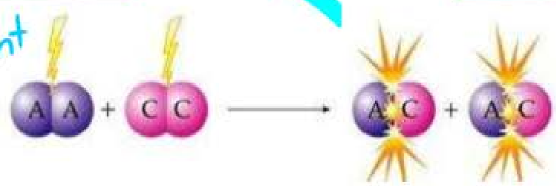


very little energy released when product bonds form.

new bonds form + energy is released... BUT much less energy released compared to the amount required at start of rxn.

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

Little energy required to break reactant bonds.



MORE energy is released when product bonds form.

LOTS of energy released as excess. (exothermic)

new bonds form (requires little energy)

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.**

- $+\Delta H$ In an **endothermic** reaction, more energy will be stored in the products than in the reactants: energy has entered the system (required) $H_{reactants} < H_{products}$
- $-\Delta H$ In an **exothermic** reaction, less energy will be stored in the products than in the reactants: energy has exited the system (released) $H_{reactants} > H_{products}$

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:

$$\boxed{\Delta H = H_{\text{products}} - H_{\text{reactants}}}$$

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

ΔH value positive --> energy absorbed --> endothermic reaction

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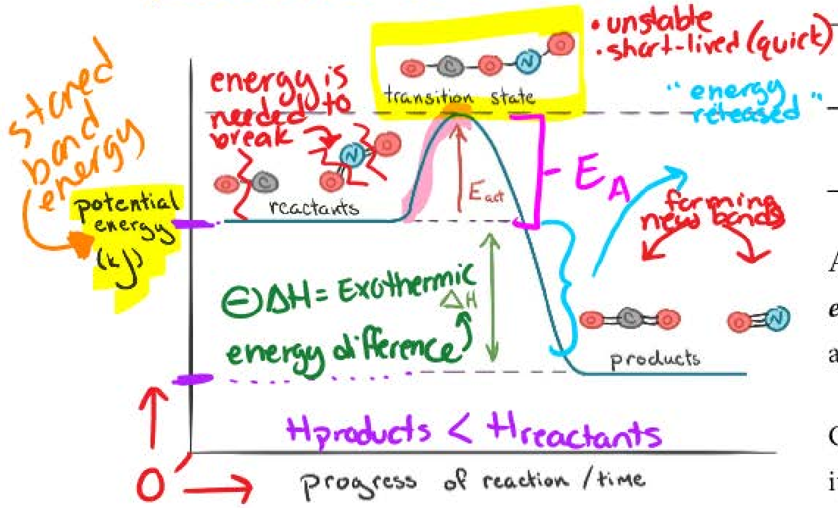
Potential Energy Diagrams

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 (E_A)

$CO_{(g)} + NO_{2(g)} \rightarrow CO_{2(g)} + NO_{(g)}$



We can define *activation energy* as: the difference in the energy between the transition state, and the reactanst.

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

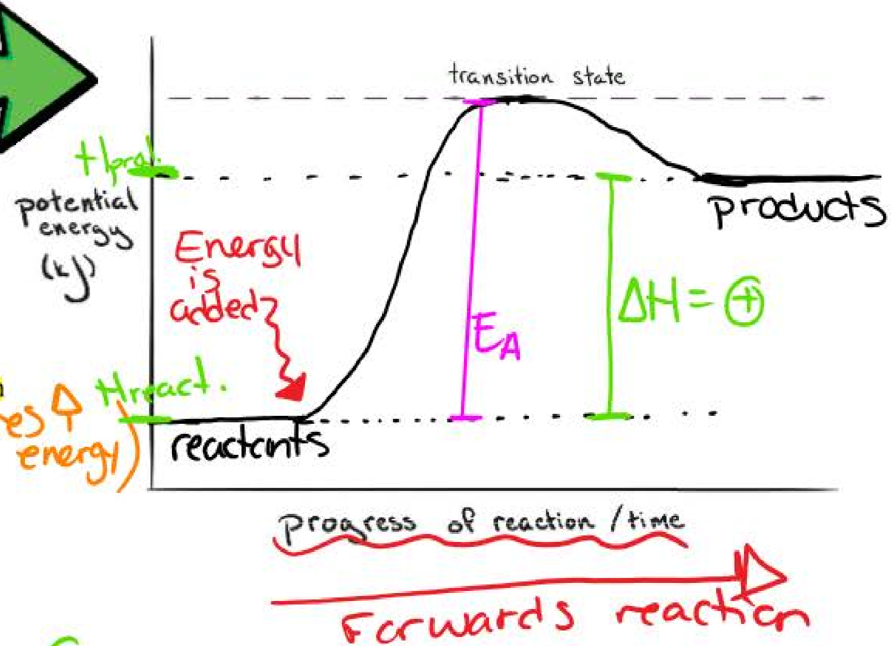
Can calculate the E_A and ΔH for any reaction from its potential energy diagram.

The activation energy (E_A) 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.*** + or −

Endothermic Reaction

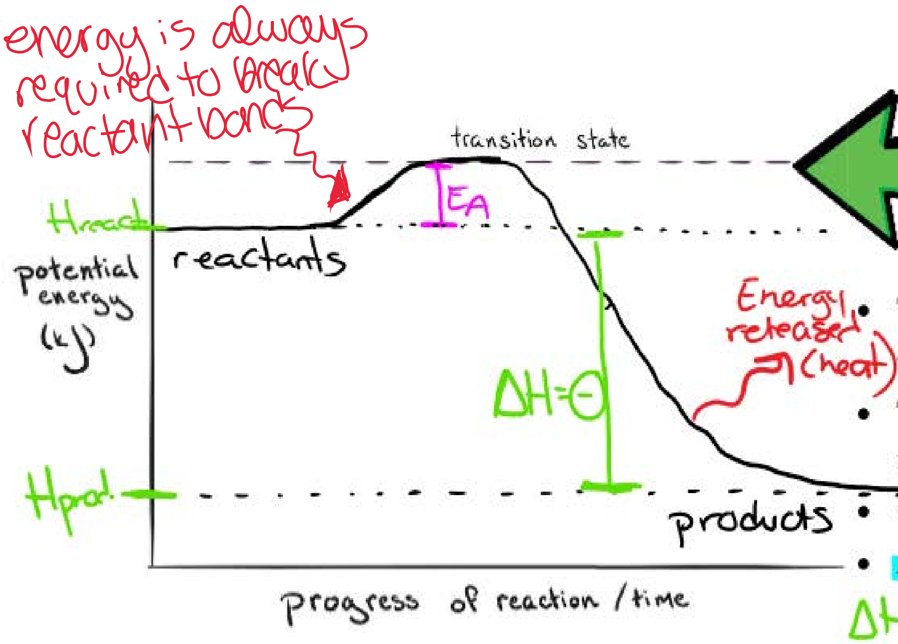
- The ***reactants are at a lower energy*** level compared to the products
- The ***products are less stable*** than the reactants.
- forcing the reaction in the forward direction towards more unstable species (requires ↑ energy)
- overall ***ΔH*** for the reaction is ***positive***,
- energy is absorbed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



$\Delta H = H_{prod.} - H_{react.}$

$\Delta H = BIG\# - small\# = \oplus$

28



Exothermic Reaction

- The ***reactants are at a higher energy*** level compared to the products
- The ***products are more stable*** than the reactants. (small E_A)
- Overall ***ΔH*** for the reaction is ***negative***
- ***Energy is released*** in the form of ***heat.***

$\Delta H = H_{prod.} - H_{reac.}$

$\Delta H = \text{small\#} - \text{BIG\#} = \ominus$

Representing Energy Changes within Chemical Reaction Equations

- Enthalpy has units of Joules (*J*) => or kJ
- Balanced reaction equations that include the enthalpy change are known as **thermochemical equations.** "Heat"
- 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:*

This form distinguishes exothermic from endothermic by heat term sign

ΔH ⊕/⊖

1. Writing the enthalpy change ***immediately after*** the equation - ***using the sign of ΔH*** to indicate whether the change is endothermic or exothermic. written after RXN

Exothermic Example: $2\,C_8H_{18} + 25\,O_2 \rightarrow 16\,CO_2 + 18\,H_2O$; $\Delta H =$ −10 992 kJ

Endothermic Example: $6\,CO_2 + 6\,H_2O \rightarrow C_6H_{12}O_6 + 6\,O_2$; $\Delta H =$ +393 kJ

This form distinguishes exothermic from endothermic by the side the heat term is written on.

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

Exothermic Example: $2C_8H_{18} + 25O_2 \rightarrow 16\,CO_2 + 18\,H_2O +$ 10 992 kJ

Endothermic Example: $6\,CO_2 + 6\,H_2O +$ 393 kJ $\rightarrow C_6H_{12}O_6 + 6\,O_2$

written on reactant side.

chemistry homework

Assignment #8: Hebden 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.