Writing Chemical Formulas and Chemical Reactions

Chemical Formula

Chemical formulas are a useful way to convey information about a compound such as:

- What elements make up the compound
- The ratio or number of atoms in the compound

The chemical formula has different meanings depending on the type of **intramolecular** forces holding together the compound:

Covalent Compounds – the chemical formula represents how many of each type of atom are in each individual molecule.

Example: H_2O_2 is a molecule with exactly 2 hydrogen atoms and 2 oxygen atoms per molecule.

Ionic Compounds – the formula represents a ratio rather than a discrete particle (recall that ionic compounds form lattice structure).

Example: MgO is an ionic compound that has one magnesium atom attached to every one oxygen atom in the crystal lattice structure.

When writing chemical formula, they must be ordered such that:

- Least electronegative element/ion is first
- **Most electronegative element/ion is second**

Name: ______________________________ Date: ____________________

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Nomenclature

1. BONDING CAPACITY (VALENCE)

The number of bonds an atom can make. For a **Cation**, the bonding capacity is the number of electrons lost to become stable. For an **Anion**, the bonding capacity is the number of electrons gained to fill the outer energy level to eight.

single e

Ex. For sodium the bonding capacity is **ONE**. For oxygen the bonding capacity is **TWO**.

Na

O

2. OXIDATION NUMBERS

The charge of the ion of an atom.

Ex. For sodium the oxidation number is **+1**. For oxygen the oxidation number is **-2**.

3. NAMING ELEMENTS

<u>All elements</u> (substances composed of only one type of atom) are named as on the periodic table.

Ex. Mg → Magnesium Fe → Iron

4. NAMING DIATOMIC GASES

The following gases exist in nature in a diatomic form having the general chemical formula **"X_2"**. The names of these binary compounds are found by just using the element's name from the Periodic Table.

Name of diatomic gas	**Formula for diatomic gas**
hydrogen	$H_2(g)$
oxygen	$O_2(g)$
fluorine	$F_2(g)$
bromine	$Br_2(l)$
iodine	$I_2(s)$
nitrogen	$N_2(g)$
chlorine	$Cl_2(g)$

REMEMBER: H_2, O_2, F_2, Br_2, I_2, N_2, Cl_2

5. NAMING MONATOMIC GASES

The elements of Group 8A (Noble gases) exist in nature as monatomic gases. These gases are considered "inert" or non-reactive under most conditions. Some may react under extreme pressures or temperatures. These elements are **NOT** binary chemical compounds, but you should know the names and formulas of these elements. Use your Periodic Table to determine the noble gas that is at the end of each period.

period	name of gas	formula or symbol
period 1	helium gas	$He(g)$
period 2	neon gas	$Ne(g)$
period 3	argon gas	$Ar(g)$
period 4	krypton gas	$Kr(g)$
period 5	xenon gas	$Xe(g)$
period 6	radon gas	$Rn(g)$

6. NAMING BINARY IONIC COMPOUNDS

Binary ionic compounds are compounds containing only two elements (a metal and a non-metal)

RULES:

1. Write the cation (metal) first, using the name of the element as on the periodic table.

2. Write the anion second, dropping the usual ending (-ine, -ium, -ogen etc.).and replace it with -**ide**.

Ex. $NaCl$ sodium chloride
KBr potassium bromide
K_2O potassium oxide

chemical formula	chemical name
NaI	sodium iodide
BeF_2	beryllium fluoride
MgO	magnesium oxide
Na_2O	sodium oxide

Li_2S	
BCl_3	
Al_2O_3	
K_2S	

7. WRITING FORMULA FOR BINARY IONIC COMPOUNDS RULES:

1. Write the symbol for the cation (metal) then write the symbol for the anion (nonmetal) beside.
2. Write the charge of the ion (oxidation number) for each element above the element.
3. Cross over the charges for each ion and leave out the charge sign.
4. Write the crossed over charges as subscripts behind the ion to which it refers.
5. Reduce the whole number ratio to lowest terms

Ex. # 1 potassium oxide

1. K O
2. $\overset{+1}{K}$ $\overset{-2}{O}$
3. K_2O

Ex # 2 calcium nitride

1. Ca N
2. $\overset{+2}{Ca}$ $\overset{-3}{N}$
3. Ca_3N_2

chemical name	chemical formula
sodium fluoride	
* lithium chloride	$\overset{+1}{Li}$ $\overset{-1}{Cl}$
beryllium bromide	
* magnesium oxide	$\overset{+2}{Mg_2}\overset{-2}{O_2} \rightarrow MgO$
boron iodide	
* aluminum sulfide	$\overset{+3}{Al_2}\overset{-2}{S_3}$
potassium oxide	
calcium fluoride	

8. WRITING NAMES AND FORMULA FOR MULTI-VALENT CATIONS

A "**multi-valent cation**" is an element that can form more than one stable **POSITIVE** ion. The term "multi-valent" means the same as "multi-oxidation state". Different positive ions of the same element are formed when reacting under different conditions. Use your "Oxidation States" sheet and your Periodic Table to identify the "multi-valent cations".

plumbum = Pb
cuprum = Cu
ferrum = Fe
aurum = Au
stannum = Sn

A) "Ous-ic" Method

RULES:

1. Find the latin name of the cation in the chemical formula. Usually, the latin name for **Hg** and **Sb** are not used. **<u>If the cation does not have a latin name, ignore this step.</u>**
2. Remove the last syllable (usually "um" for the latin name) and add the suffix ("ous" or "ic") in its place. Arsenic's name remains unchanged when the higher oxidation state is used. For some elements, the last syllable is not removed (i.e., Co, Ni).

The suffix "ous" indicates the lower oxidation state was used for the cation.

The suffix "ic" indicates the higher oxidation state was used for the cation.

3. The anion name is written as you have done previously (ending with "ide").

chemical name	chemical formula	chemical name	chemical formula
ferrous oxide	$\overset{+2}{Fe}\ \overset{-2}{O}$	ferric oxide	Fe_2O_3
stannous chloride			$SnCl_4$
plumbous sulfide	$\overset{+2}{Pb}\ \overset{-2}{S}$	plumbic sulphide	$\overset{+4}{Pb}\overset{-2}{S_2}$
cuprous bromide			$CuBr_2$
aurous iodide			AuI_3
mercurous fluoride			HgF_2

B) Roman Numeral Method

The "Roman Numeral" or "Stock System" method is the **most widely used** and the **preferred** method for naming chemical compounds containing a **multi-valent metal cation**. This method is **NOT** used if the cation has only a single valence or oxidation state.

I II III IV V VI VII
1 2 3 4 5 6 7

RULES:

i) Naming:

1. The English name of the multi-valent metal cation is written first.
2. A Roman numeral indicating the positive charge on the cation is written in brackets after the cation's name. **No space** is left between the cation name and the Roman numeral in brackets.
3. The anion name is written as you have done previously (ending with "ide").

chemical formula	chemical name
CuF_2	copper(II) fluoride
+,6,+4,2 MnO_2 -2	manganese(IV) oxide
$NiCl_3$	
SnS_2	
HgI_2	mercury(II) iodide

ii) Writing Formula

1. Follow the same rules used when writing the formula of regular binary compounds; however use the oxidation number indicated in brackets after the cation.

chemical name	chemical formula
copper(I) fluoride	
manganese(II) oxide	MnO
nickel(II) chloride	$NiCl_2$
tin(II) sulfide	SnS
mercury(I) iodide	

9. NAMING BINARY MOLECULAR (COVALENT) COMPOUNDS

Binary molecular (covalent) compounds are compounds containing only two elements (**a non-metal and a non-metal**)

RULES:

1. Attach a prefix that indicates the number of atoms in the chemical formula to the front the name of element that is more to the left on the periodic table. **The prefix "mono-" is omitted from the name of the first element if it is the prefix required.** The Greek prefixes are as follows:

mono - 1	tri - 3	penta - 5	hepta - 7	nona - 9
di - 2	tetra - 4	hexa - 6	octa - 8	deca - 10

2. A second prefix indicating the number of atoms in the chemical formula is attached to the second name of the element. This element is usually found on the right of the periodic table. The name of the second element ends in **-ide**.

3. The "o" or "a" ending of the prefix is omitted if the cation or anion name starts with an "o" or "a".

Ex. P_2O_3 is called diphosphorous trioxide
CO_2 is called carbon dioxide
N_2O_5 is called dinitrogen pentoxide
H_2O is called dihydrogen monoxide

chemical formula	chemical name
CO_2	
As_2O_3	diarsenic trioxide
NO_2	
P_2O_5	diphosphorus pentoxide
CBr_4	

10. WRITING FORMULA FOR MOLECULAR (COVALENT) COMPOUNDS

RULES:

1. Write the elemental symbol for each of the elements named.
2. Use the prefixes to determine the number of elements in each molecule.

Ex. dinitrogen tetroxide N_2O_4
diphosphorous trisulfide P_2S_3
carbon tetrachloride CCl_4
sulfur hexafluoride SF_6

chemical name	chemical formula
nitrogen monoxide	NO
silicon dioxide	SiO_2
carbon monoxide	
sulfur trioxide	
phosphorus pentabromide	

11. WRITING NAMES FOR POLYATOMIC IONS

Polyatomic ions are groups of atoms which act as a unit. They consist of two or more different non-metal atoms joined by a covalent bond. The bonding capacity of the polyatomic ion is the same as the charge.

PO_4^{-3}	phosphate	ClO_3^{-1}	chlorate	CO_3^{-2}	carbonate
SO_4^{-2}	sulfate	NO_3^{-1}	nitrate	OH^{-1}	hydroxide
NH_4^{+1}	ammonium	HCO_3^{-1}	hydrogen carbonate		

RULES:

1. Write the cation first, using the name of the element as on the periodic table.
2. Write the name of the polyatomic ion with the -<u>ate</u> ending.

Ex. $Na_2(SO_4)$ sodium sulf<u>ate</u> $K_3(PO_4)$ potassium phosph<u>ate</u>

chemical formula	chemical name
H_3PO_4	hydrogen phosphate
K_2CO_3	potassium carbonate
H_2SO_4	hydrogen sulphate
$(NH_4)_2S$	ammonium sulphide

12. WRITING FORMULA FOR POLYATOMIC IONS

RULES:

1. Write the symbol for the cation or polyatomic ion named first then write the symbol for the polyatomic ion or anion named second.
2. Write the charge of the ion (oxidation number) for each element or polyatomic ion above each element or polyatomic ion.
3. Cross over the charges for each ion and leave out the charge sign.
4. Write the crossed over charges as subscripts behind the ion to which it refers.
5. Reduce the whole number ratio to lowest terms.

Ex. # 1 Magnesium sulfate

1. Mg SO_4
2. $\overset{+2}{Mg}$ $\overset{-2}{SO_4}$
3. Mg SO_4

Ex # 2 ammonium chloride

1. NH_4 Cl
2. $\overset{+1}{NH_4}$ $\overset{-1}{Cl}$
3. NH_4 Cl

chemical name	chemical formula
ammonium sulfate	$(NH_4)_2SO_4$
sodium hydroxide	NaOH
magnesium sulfate	$MgSO_4$
hydrogen carbonate	H_2CO_3

13. NAMING AND WRITING FORMULAS FOR POLYATOMIC ION DERIVATIVES (SALTS):

A **derivative** may be formed during a chemical reaction when atom(s) or ions are added to or removed from a polyatomic ion. Information in the box below shows how the name and formula are changed when adding or removing atoms or ions from the original formula.

- if **ONE OXYGEN ATOM IS ADDED**, add prefix "per" to the name.

e.g. ClO_4^{-1} = per chlor ate SO_5^{-2} persulphate

- if **ONE OXYGEN ATOM IS REMOVED**, remove "ate" and add "ite" to name.

e.g. SO_3^{-2} = sulph ite

- if **TWO OXYGEN ATOMS ARE REMOVED**, add prefix "hypo" and "ite" to name.

e.g. NO^{-1} = hypo nitr ite

- if **REPLACING AN OXYGEN ATOM WITH AN S,** add the prefix "thio" to name.

e.g. SO_4^{-2} **to** $S_2O_3^{-2}$ = thio sulph ate

- if **ADDING A HYDROGEN ION**, add prefix "hydrogen" or "bi" to name.

e.g. CO_3^{-2} **to** HCO_3^{-1} = bi carbon ate

- if ADDING **TWO HYDROGEN IONS**, add prefix "dihydrogen" to name.

e.g. CO_3^{-2} **to** H_2CO_3 = dihydrogen carbon ate

14. WRITING NAMES AND FORMULA FOR BINARY ACIDS

A **binary acid** is a binary chemical compound containing hydrogen and a nonmetal from Group 6 or 7. These compounds can be named using the regular naming system for binary molecular compounds if they are gases. But, binary acids are usually found as clear, viscous liquids at room temperature and a different naming system is used when they are in this state. If the binary acid is in aqueous state, the prefix "hydro" and ending "ic" is added to the first syllable of the nonmetal and this becomes the first part of the name. The word "acid" is included as the second part of the name.

Use the above instructions and examples included in the below to name the binary acids in their liquid and gas form.

binary acid formula	name of binary acid (liquid form) (aq)	name of gas (vapour form) (g)
HF	hydrofluoric acid	hydrogen Fluoride
HCl	hydrochloric acid	hydrogen chloride
HBr	hydrobromic acid	hydrogen bromide

HI	hydroiodic acid	hydrogen iodide
* H_2S	hydrosulphuric acid	
H_2Se		

15. WRITING FORMULA FOR OXY ACIDS

An **oxy acid** is a polyatomic compound containing hydrogen, oxygen and an electronegative element (i.e., Cl, N. S, P, etc.). These acids are sometimes referred to as "mother acids" because many names and formulae of other oxy acids and polyatomic ions are derived from these. The top 5 oxy acids are the main oxy acids used in industry. Use the names already given to help fill in the chart below.

oxy acid chemical name	oxy acid chemical formula
phosphoric acid	H_3PO_4
sulphuric acid	H_2SO_4
carbonic acid	H_2CO_3
nitric acid	HNO_3
chloric acid	$HClO_3$
fluoric acid	HFO_3
bromic acid	$HBrO_3$
iodic acid	HIO_3

16. NAMING AND WRITING FORMULAS for OXY ACID DERIVATIVES:

An **oxy acid derivative** may be formed during a chemical reaction when oxygen atom(s) are added to or removed from an **oxy acid**. (**Note:** Although you can write the chemical formulas and names of all oxy acid derivatives using the system described below, some may not be able to be produced naturally or synthetically.) Information in the box below shows how the name and formula of an oxy acid are changed when adding or removing oxygen atoms from the original **oxy acid** formula.

- if **ONE OXYGEN ATOM IS ADDED**, add prefix "per" to the "ic" acid name = per sulphuric acid (called "per-ic" acid) $H_2SO_5(aq)$

- **OXY ACID** = sulfuric acid (called "ic" acid or AKA "mother acid")

- if **ONE OXYGEN ATOM IS REMOVED**, remove "ic" and add "ous" to acid name = sulphur ous acid (called "ous" acid) $H_2SO_3(aq)$

- if **TWO OXYGEN ATOMS ARE REMOVED**, add prefix "hypo" to "ous" acid name = hypo sulphur ous acid (called "hypo-ous" acid) $H_2SO_2(aq)$

Complete the following table to show how you would name and write the formulas for the oxy acid derivatives of sulfuric acid (use information from the box above).

change in oxy acid formula	oxy acid name	oxy acid formula
add one oxygen atom(*) ("per-ic" acid)	persulfuric acid	$H_2SO_5(aq)$
no change ("ic" acid)	sulfuric acid	H_2SO_4
remove one oxygen atom(*) ("ous" acid)	sulfurous acid	$H_2SO_3(aq)$
remove two oxygen atoms(*) ("hypo-ous" acid)	hyposulfurous acid	$H_2SO_2(aq)$

Chemical Reactions

A chemical reaction can be written in a number of different forms:

Chemical Equation

A description of a chemical reaction using **symbols**, not **words**, where:

- The **reactants** are written first
- The **products** are written second
- The state for each element or compound is indicated in brackets – **solid** (s), **liquid** (l), **gas** (g), **aqueous** (aq)
- Reactants and products are separated by an arrow (→) – read as "**yields**"

Example:

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

Word Equation

The elements and compounds that are reacting are written first followed by the products. States are included in the description.

Example:

Hydrogen gas reacts with chlorine gas to produce hydrogen chloride gas

Skeleton Equation

The Law of Conservation of Mass states that matter cannot be **created** or **destroyed**; it can only be **changed** from one form to another. Therefore the **number of atoms** in the reactants must **equal** the number of atoms in the products.

A skeleton equation is an unbalanced equation that **does not** follow the Conservation of Mass. The number of atoms on the left side (reactants) of the chemical equation **does not** equal the number of atoms on the right side (products).

Example:

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

On the reactant side there is a total of **4** atoms (**2** hydrogen and **2** chlorine)
On the product side there is a total of **2** atoms (**1** hydrogen and **1** chlorine)

Balanced Chemical Equation

An equation that follows the Law of Conservation of Mass. The number of atoms on the reactant side equals the atoms on the product side. In most chemical equations, numbers placed in front of the elements or compounds (**coefficients**) are required to balance the equation.

Example:

$$1H_2\ (g) + 1Cl_2\ (g) \rightarrow 2HCl\ (g)$$

On the reactant side there is a total of **4** atoms (**2** hydrogen and **2** chlorine)
On the product side there is a total of **4** atoms (**2** hydrogen and **2** chlorine)

When there is a coefficient of "**1**", it is typically not written:

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

Balancing Equations

All chemical equations must be balanced so that they are consistent with the Law of Conservation of Mass.

Here are some suggestions for balancing equations:

1. When balancing equations, always start with the "ugliest" molecule first (polyatomics).

2. To balance, place the desired number (coefficient) in from of the element or compound. Never split-up a compound and never change the subscripts in the chemical formula.

3. It is often useful to balance the diatomic molecules, if they are present, last.

4. Creating a chart to keep track of the type and number of each atoms on the reactant and product side of the equation can make balancing easier.

5. Make sure to always recheck the final balanced equation.

leave H to the end AND O to the very, very end

Examples:

I can do it!

__2__ Mg (s) + _____ O_2 (g) → __2__ MgO (s)

Atoms	Reactants	Products
Mg		
O		

__2__ H_2 (g) + _____ O_2 (g) → __2__ H_2O (g)

Atoms	Reactants	Products
H		
O		

4 Fe (s) + 3 O_2 (g) → 2 Fe_2O_3 (s)

Atoms	Reactants	Products
Fe		
O		

_____ Al_2O_3 (s) + 3 H_2 (g) → 3 H_2O (l) + 2 Al (s)

Atoms	Reactants	Products
Al		
O		
H		

3 $Pb(NO_3)_2$ (aq) + 2 BF_3 (s) → 2 $B(NO_3)_3$ (aq) + 3 PbF_2 (s)

Atoms	Reactants	Products
Pb		
NO_3		
B		
F		

Sometimes to balance an equation, fractions must be used. Fractions are not to be left in the final balanced equation, as it impossible to have part of an atom. To get rid of the fraction, multiply every element or compound in the equation by the denominator of the fraction (i.e. If you use $\frac{1}{2}$ as a coefficient, then multiply by 2).

~~2~~ 4 NH_3 (l) + 7 ~~3~~ O_2 (g) → ~~2~~ 4 NO_2 (g) + 3 6 H_2O(s)

Atoms	Reactants	Products
N		
H		
O		

~~2~~ 4 FeS_2 (s) + 11 ~~M~~ O_2 (g) → 2 Fe_2O_3 (s) + ~~4~~ 8 SO_2 (g)

Atoms	Reactants	Products
Fe		
S		
O		

Balancing chemical equations becomes increasing more difficult when you are given the reaction as a word equation. To balance the equation, you must first convert the elements and/or compounds into their correct chemical formula. Even the slightest mistake will make you equation incorrect and could possibly create an equation that is impossible to balance. Be careful, and make sure to always check your work.

Write out a balanced chemical equation for the following:

Oxygen gas reacts with solid aluminum sulfide to produce solid aluminum oxide and sulfur dioxide gas.

Types of Chemical Reactions

It is important to be able to classify chemical reactions as it enables scientists to predict possible products or outcomes. For example, think of appropriate storage of chemicals…

Why are some chemicals stored in dark glass jars?

Why is it inappropriate to store propane tanks in areas that are not air-conditioned?

Below are 4 major categories of chemical reactions:

1. Synthesis

A synthesis reaction occurs when 2 or more **elements** combine to form a new **molecule** or **compound**.

The general equation for a synthesis reaction is: **A + B → C**

Specific types of synthesis reactions:

a) **Metals** react with **oxygen** to produce a **metal oxide** !

$2Mg_{(s)} + O_{2(g)} \rightarrow 2MgO_{(s)}$

b) A **non-metal** reacts with **oxygen** to produce a **non-metal oxide**

$C_{(s)} + O_{2(g)} \rightarrow CO_{2(g)}$

c) A **metal** and **non-metal** combine to form a **binary ionic compound**

$2K_{(s)} + Cl_{2(g)} \rightarrow 2K^{+1}Cl^{-1}_{(s)}$

d) **Non-metallic oxides** react with **water** to produce an **acid** **

$SO_{3(g)} + H_2O_{(l)} \rightarrow H_2SO_{4(aq)}$

e) **Metallic oxides** react with **water** to produce a **base**

$CaO_{(s)} + H_2O_{(l)} \rightarrow Ca^{+2}(OH)^{-1}_{2(aq)}$

2. Decomposition

A decomposition reaction is the reverse to a synthesis reaction, a compound **breaks down** into **elements** or other **compounds**

The general equation for a decomposition reaction is: **C → A + B**

Example: $2H_2O_2(l) \rightarrow 2H_2O(l) + O_2(g)$

Typically, some form of **energy** or type of **catalyst** is needed to initiate a decomposition reaction.

A catalyst is a substance that controls the **rate** of a reaction, without being **used-up** during the reaction or affecting the overall **products**.

3. Combustion Reactions

Combustion reactions typically involve a **metal** or **non-metal** reacting with **oxygen** to form the most common **oxide** of the elements that make up that compound. There are 2 types of combustion reactions; **complete** and **incomplete.**

Complete combustion occurs in the presence of **surplus oxygen**. The reaction is very **"clean"** with **carbon dioxide** and **water** being the only products.

Example: ____CH_4 + __2__O_2 → $CO_2(g) + 2H_2O(g)$

____C_8H_{16} + __12__O_2 → $8CO_2(g) + 8H_2O(g)$

Incomplete combustion occurs in the presence of an **insufficient amount** of **oxygen**. The reaction is very "**dirty**" with various products, which can include **carbon dioxide**, **water**, **carbon monoxide** and **soot** (solid carbon).

INSUFFICIENT

Example: ____CH_4 + ____O_2 → $CO_2(g) + H_2O(g) + CO(g) + C(s)$

Note: CO_2 and H_2O are also produced in both of these types of reactions

3. Single Displacement Reaction

A single Displacement reaction occurs when one **element** in a compound is **displaced/replaced** by another **element**. This can occur in 2 ways, a **metal** can replace a **metal** or a **non-metal** can replace a **non-metal**

The general equation for a single displacement reaction is: **A + BC → AC + B**

1.31 1.65

Examples: $\mathbf{Mg + ZnCl_2 \rightarrow MgCl_2 + Zn}$

$\mathbf{Cu + AgNO_3 \rightarrow CuNO_3 + Ag}$

How do you know that a single displacement reaction can occur or do they always occur?

For example, explain why the two above reactions occur but the following reaction does not?

$\mathbf{Ni + CdCl_2 \rightarrow}$ **No Reaction**

In order to determine if an element will displace another element in a single displacement reaction you must refer to an **activity series chart**.

If one element is **above** another element in the compound, it can be **bumped out** and a single displacement reaction will occur.

EN

Non-metals, typically **halogens** are involved in Single Displacement Reactions. To determine who can bump out whom, you must refer to the **Activity Series for Halogens**.

Predict if the following reactions will occur and what the products are:

Fluorine
Chlorine
Bromine
Iodine

2.66 3.16

I_2 + NaCl → NR

3.98 2.96

F_2 + 2KBr → 2KF + Br_2

4. Double Displacement Reactions

A double displacement reaction occurs when there is an **exchange** of **cations** between two **ionic** compounds.

The general equation for a double displacement reaction is:

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

In the general equation above, A and C are **cations** (written first) and B and D are **anions.**

How do you know that a double displacement reaction can occur or will they always occur?

Evidence that a double displacement reaction will/has occurred:

A) A solid precipitate (ppt) forms
B) A gas is produced, bubbles form
C) Water (H_2O) is formed

Example: $NaCl + AgNO_3 \rightarrow NaNO_{3(aq)} + AgCl_{(s)}$

Example: $Na_2CO_3 + HCl \rightarrow NaCl_{(aq)} + H_2CO_3 \rightarrow H_2O + CO_{2(g)}$

once on assign

Example: $2H_3PO_4 + 3Ca(OH)_2 \rightarrow 6H_2O_{(l)} + Ca_3(PO_4)_2$

+2 -3

Water is evidence of an **acid/base** reaction (**neutralization**), which is a type of double displacement reaction. Since water is a clear, colourless, liquid, it typically cannot be seen by looking at the reaction. To determine if water is present, it has to be tested using **indicators** or **pH values**.