

# Chapter 4

# Nomenclature and Chemical Equations

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## Chapter Objectives

1. To learn how to name a compound
  2. To write balanced chemical equations to represent chemical reactions
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## Introduction

A chemical reaction occurs when new substances form. To describe chemical reactions, chemists use **chemical equations** which summarize both the **qualitative** information, i.e. what chemicals react and what products form, and **quantitative** information, i.e. the relative amount of reactants consumed and products produced. In this chapter, we will learn how to describe the qualitative information of a reaction by writing chemical formulas of substances and chemical equations. The quantitative information of a reaction will be discussed in the next chapter.

## Nomenclature of Compounds

From the last chapter, we learnt that compounds contain either ionic bonds or covalent bonds. Depending on the type of bonds present in a compound, different rules are applied to its naming.

### (i) Ionic Compounds

In ionic compounds, there are both cations and anions. The ions can be either monatomic or polyatomic. **Monatomic cations** can be divided into two categories. One category includes cations that are unique to their elements. For example, a hydrogen atom always forms a hydrogen ion,  $H^+$ , while a sodium atom always forms a sodium ion,  $Na^+$ . The ions share the same names with their original atoms. This applies to most representative metals and a few d-block elements such as silver and zinc.

When writing the chemical formula of an ion, its charge must be stated with a superscript. The magnitude of the charge is always expressed before the sign of the charge. Calcium atoms can form calcium ions with a +2 charge. Therefore, the chemical formula for a calcium ion is  $\text{Ca}^{2+}$ .

The charges of the monatomic cations formed by most representative elements can be deduced by their group numbers. Sodium is in group IA so the charge of its ion is +1, while calcium is in group IIA and the charge of its ion is +2.

Some common monatomic cations unique to their elements are listed in Table 4.1. Among them, it should be noted that there are two *d*-block elements, zinc and silver, which form unique ions.

**Table 4.1 Common Monatomic Cations Unique to Their Elements**

Charge	Name	Symbol	Charge	Name	Symbol	Charge	Name	Symbol
+1	Hydrogen	$\text{H}^+$	+2	Magnesium	$\text{Mg}^{2+}$	+3	Aluminum	$\text{Al}^{3+}$
	Lithium	$\text{Li}^+$		Calcium	$\text{Ca}^{2+}$			
	Sodium	$\text{Na}^+$		Strontium	$\text{Sr}^{2+}$			
	Potassium	$\text{K}^+$		Barium	$\text{Ba}^{2+}$			
	Rubidium	$\text{Rb}^+$		Zinc	$\text{Zn}^{2+}$			
	Cesium	$\text{Cs}^+$						
	Silver	$\text{Ag}^+$						

The other category includes cations which are formed by the same kind of metal atom but are capable of having different charges. These ions are mainly produced from transition elements. For examples, iron forms either an iron(II) ion,  $\text{Fe}^{2+}$ , or an iron(III) ion,  $\text{Fe}^{3+}$ ; copper forms either a copper(I) ion,  $\text{Cu}^+$ , or a copper(II) ion,  $\text{Cu}^{2+}$ . The names of the ions include both the name of the original atoms and the charges of the ions written in Roman numerals embraced within parentheses. The common ions of this type are listed in Table 4.2 on the next page.

**Table 4.2 Common Monatomic Cations Not Unique to Their Elements**

Charge	Name	Symbol	Charge	Name	Symbol	Charge	Name	Symbol
+1	Copper(I)	Cu <sup>+</sup>	+2	Manganese(II)	Mn <sup>2+</sup>	+3	Chromium(III)	Cr <sup>3+</sup>
				Iron(II)	Fe <sup>2+</sup>		Iron(III)	Fe <sup>3+</sup>
				Copper(II)	Cu <sup>2+</sup>			
				Tin(II)	Sn <sup>2+</sup>			
				Mercury(II)	Hg <sup>2+</sup>			
				Lead(II)	Pb <sup>2+</sup>			

**Polyatomic cations** usually have names different from that of their constituent atoms. In SAT chemistry, there are two polyatomic cations to memorize. They are ammonium ion, NH<sub>4</sub><sup>+</sup>, and hydronium ion, H<sub>3</sub>O<sup>+</sup>, as list in Table 4.3. NH<sub>4</sub><sup>+</sup> is known to form ionic compounds with some anions while H<sub>3</sub>O<sup>+</sup> is only stable in aqueous solution.

**Table 4.3 Common Polyatomic Cations**

Charge	Name	Symbol
+1	Ammonium	NH <sub>4</sub> <sup>+</sup>
	Hydronium	H <sub>3</sub> O <sup>+</sup>

**Monatomic anions** are named by replacing the end of the element names with “-ide”. For example, an oxygen atom can form an oxide ion, O<sup>2-</sup>, while a fluorine atom can form a fluoride ion, F<sup>-</sup>. The charge of the ions can be deduced by subtracting 8 from its group number, except for the hydride ion. Oxygen is in group VIA and its ion has a 6 – 8 = -2 charge. Fluorine is in group VIIA and its ion has a 7 – 8 = -1 charge. The common monatomic anions are listed in Table 4.4 on the next page.

**Table 4.4 Common Monatomic Anions**

Charge	Name	Symbol	Charge	Name	Symbol	Charge	Name	Symbol
-1	Hydride	H <sup>-</sup>	-2	Oxide	O <sup>2-</sup>	-3	Nitride	N <sup>3-</sup>
	Fluoride	F <sup>-</sup>		Sulfide	S <sup>2-</sup>		Phosphide	P <sup>3-</sup>
	Chloride	Cl <sup>-</sup>						
	Bromide	Br <sup>-</sup>						
	Iodide	I <sup>-</sup>						

Most polyatomic anions are produced by acid dissociation in water. For example, nitric acid, HNO<sub>3</sub>, and sulfuric acid, H<sub>2</sub>SO<sub>4</sub>, dissolve in water to give nitrate ions, NO<sub>3</sub><sup>-</sup>, and sulfate ions, SO<sub>4</sub><sup>2-</sup>, respectively. These ions generally contain a central nonmetal atom surrounded by oxygen atoms. The number of oxygen atoms surrounding a specific nonmetal atom can vary which gives rise to a series of anions as shown in Table 4.5. If a series only contains two members, which is most commonly the case, the member that has more oxygen is named by replacing the end of the name of the central atom with “-ate”, while another one is named by replacing with “-ite”. Examples are sulfate, SO<sub>4</sub><sup>2-</sup> and sulfite, SO<sub>3</sub><sup>2-</sup>. There is a four-member series related to chlorine that is worth memorizing. The members are perchlorate ion, ClO<sub>4</sub><sup>-</sup>, chlorate ion, ClO<sub>3</sub><sup>-</sup>, chlorite ion, ClO<sub>2</sub><sup>-</sup> and hypochlorite ion, ClO<sup>-</sup>.

**Table 4.5 Series of Common Polyatomic Anions**

Charge	Name	Symbol	Charge	Name	Symbol
-1	Nitrite	NO <sub>2</sub> <sup>-</sup>	-2	Sulfite	SO <sub>3</sub> <sup>2-</sup>
	Nitrate	NO <sub>3</sub> <sup>-</sup>		Sulfate	SO <sub>4</sub> <sup>2-</sup>
-1	Hypochlorite	ClO <sup>-</sup>	-2	Carbonate	CO <sub>3</sub> <sup>2-</sup>
	Chlorite	ClO <sub>2</sub> <sup>-</sup>			
	Chlorate	ClO <sub>3</sub> <sup>-</sup>		-3	Phosphite
Perchlorate	ClO <sub>4</sub> <sup>-</sup>	Phosphate	PO <sub>4</sub> <sup>3-</sup>		

The maximum number of oxygen atoms varies with different central atoms. The elements in the second period can have at most three oxygen atoms in their ions, while the elements in the third period can have up to four oxygen atoms in their ions. In addition, the negative charges on these ions decrease across the period. This trend is shown in Table 4.6.

**Table 4.6 Polyatomic Anions with the Greatest Number of Oxygen Atom**

	Group IVA	Group VA	Group VIA	Group VIIA
Second period	$\text{CO}_3^{2-}$	$\text{NO}_3^-$		
Third period		$\text{PO}_4^{3-}$	$\text{SO}_4^{2-}$	$\text{ClO}_4^-$

The anions with multiple charges are able to receive hydrogen ions to form other anions. A carbonate ion,  $\text{CO}_3^{2-}$ , can receive one hydrogen ion to form a hydrogen carbonate ion,  $\text{HCO}_3^-$ . Phosphate ion can receive one or two hydrogen ions to form a hydrogen phosphate ion,  $\text{HPO}_4^{2-}$ , or a dihydrogen phosphate ion,  $\text{H}_2\text{PO}_4^-$  respectively. When there is one ionizable hydrogen atom in the ion, “Hydrogen” is added preceding the ion name. When there are two ionizable hydrogen atoms, “di” is added as a prefix to the name. For each addition of a hydrogen ion, the charge of the ion is lowered by one. Therefore,  $\text{HPO}_4^{2-}$  has one less negative charge than  $\text{PO}_4^{3-}$ . Further examples are shown in Table 4.7.

**Table 4.7 Anions Formed by the Combination of Hydrogen Ions and Polyatomic Anions**

Charge	Name	Symbol	Charge	Name	Symbol
-2	Carbonate	$\text{CO}_3^{2-}$	-2	Sulfate	$\text{SO}_4^{2-}$
-1	Hydrogen carbonate	$\text{HCO}_3^-$	-1	Hydrogen sulfate	$\text{HSO}_4^-$
-3	Phosphate	$\text{PO}_4^{3-}$			
-2	Hydrogen phosphate	$\text{HPO}_4^{2-}$			
-1	Dihydrogen phosphate	$\text{H}_2\text{PO}_4^-$			

Some polyatomic anions have special names and require memorization. They are listed in Table 4.8.

**Table 4.8 Series of Common Polyatomic Anions**

Charge	Name	Symbol	Charge	Name	Symbol
-1	Hydroxide	$\text{OH}^-$	-2	Oxalate	$\text{C}_2\text{O}_4^{2-}$
	Cyanide	$\text{CN}^-$		Acetate	$\text{C}_2\text{H}_3\text{O}_2^-$
	Permanganate	$\text{MnO}_4^-$		Chromate	$\text{CrO}_4^{2-}$
				Dichromate	$\text{Cr}_2\text{O}_7^{2-}$

To name an ionic compound, the name of a cation is always followed by that of an anion. The compound comprising sodium ions and chloride ions is thus named sodium chloride, while that comprising iron(III) ions and oxide ions is called iron(III) oxide.

In writing chemical formulas, the charges of cations must balance that of anions. A simple method to match the charges of cations and anions with a correct ratio is shown in Table 4.9

**Table 4.9 Procedure in Deducing Chemical Formulas for Ionic Compounds**

	Example 1 Calcium carbonate	Example 2 Ammonium sulfate	Example 3 Iron(III) oxide
1. Write chemical formulas for the ions.	$\text{Ca}^{2+}, \text{CO}_3^{2-}$	$\text{NH}_4^+, \text{SO}_4^{2-}$	$\text{Fe}^{3+}, \text{O}^{2-}$
2. Write the superscripts of the ions as the subscripts of the opposite ions. Use parentheses for polyatomic ions.	$\text{Ca}_2^{2+}(\text{CO}_3^{2-})_2$	$(\text{NH}_4^{1+})_2(\text{SO}_4^{2-})_1$	$\text{Fe}_2^{3+}\text{O}_3^{2-}$
3. Write the formulas in the simplest whole-number ratio	$\text{Ca}_1^{2+}(\text{CO}_3^{2-})_1$	$(\text{NH}_4^+)_2(\text{SO}_4^{2-})_1$	$\text{Fe}_2^{3+}\text{O}_3^{2-}$
4. Drop the charges on the ions, the subscript 1, and the parentheses if there is only one polyatomic ion in the formula.	$\text{CaCO}_3$	$(\text{NH}_4)_2\text{SO}_4$	$\text{Fe}_2\text{O}_3$

## (ii) Covalent Compounds

Simple rules for naming and writing formulas of binary covalent compounds, inorganic acids and some other common compounds will be discussed in this section. Naming covalent compounds is usually complicated and it will be covered in Chapter 13.

Binary covalent compounds are generally named according to the following rules:

1. The order of the elements in a chemical name is the same as that in the chemical formula.
2. The first element uses its own name while the second element is named by ending with “ide” similar to the anions in ionic compounds.
3. The number of atoms in the molecular formula is specified with prefixes listed in Table 4.10. For example,  $N_2O$  is named as dinitrogen monoxide.
4. When the first element contains only one atom, its prefix is neglected.  $CO_2$  is named as carbon dioxide but not monocarbon dioxide.
5. If oxygen is the second element, the vowels “a” or “o” at the end of a prefix is dropped. Therefore,  $CO$  is called carbon monoxide, not carbon monoxide;  $N_2O_5$  is called dinitrogen pentoxide, not dinitrogen pentoxide.

**Table 4.10 Prefixes in Naming Covalent Compounds**

Number of atom	Prefix	Number of atom	Prefix
1	Mono-	6	Hexa-
2	Di-	7	Hepta-
3	Tri-	8	Octa-
4	Tetra-	9	Nona-
5	Penta-	10	Deca-

For an acid that does not possess any oxygen atoms, “hydro” is used as the prefix and “ic” is used as the suffix to the name of the anion it produces in water.  $HF$  dissociates in water to produce fluoride ions; therefore it is named hydrofluoric acid.

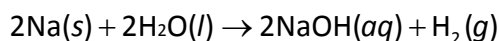
For an acid that contains oxygen atoms, the name of its anion ends in “ic”.  $HNO_3$  dissociates in water to give nitrate ions,  $NO_3^-$ , and so it is called nitric acid. Some common acids are listed in Table 4.11 on the next page.

**Table 4.11 Common Acids**

Chemical name	Chemical formula	Chemical name	Chemical formula
Hydrofluoric acid	HF	Nitric acid	HNO <sub>3</sub>
Hydrochloric acid	HCl	Sulfuric acid	H <sub>2</sub> SO <sub>4</sub>
Hydrobromic acid	HBr	Phosphoric acid	H <sub>3</sub> PO <sub>4</sub>
Hydroiodic acid	HI	Acetic acid	HC <sub>2</sub> H <sub>3</sub> O <sub>2</sub>

## Writing and Balancing Chemical Equations

A chemical equation tells us what reactants and what products are involved in a chemical reaction. An example is shown below:



It is a convention to write the reactants on the left and the products on the right. The letters enclosed in the parenthesis tell us the states of the substances: *s* denotes a solid, *l* denotes a liquid, *g* denotes a gas and *aq* denotes an aqueous solution, i.e. a homogeneous mixture in water. Therefore, the above chemical equation informs us that solid sodium reacts with liquid water to give an aqueous solution of sodium hydroxide and gaseous hydrogen. The numbers in front of the chemical formulas are called the **coefficients**. They represent the relative ratios between every reactant and product so that the numbers of atoms on both sides of the equation are balanced. When the coefficient equals to one, by convention, it is usually ignored in the equation. For the previous example, we can check whether the equation is balanced by the following table:

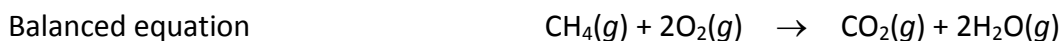
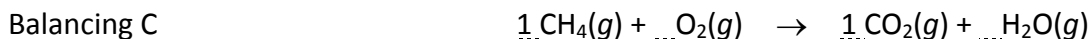
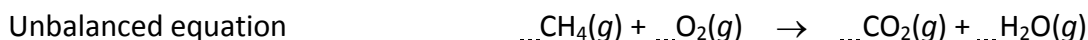
	2Na(s)	+	2H <sub>2</sub> O(l)	→	2NaOH(aq)	+	H <sub>2</sub> (g)
Each of the formulas contains	1Na		2H, 1O		1Na, 1O, 1H		2H
The coefficients of the formulas	2		2		2		1
The total number of atoms in the formulas	2Na		4H, 2O		2Na, 2O, 2H		2H
The total number of atoms on each side	2Na, 4H, 2O				2Na, 4H, 2O		

We can see that on both sides there are two sodium atoms, four hydrogen atoms and two oxygen atoms. Therefore, the equation is balanced.

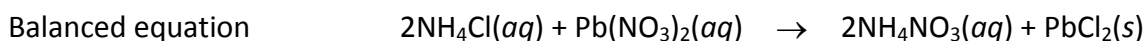
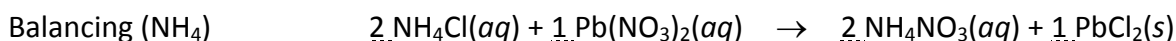
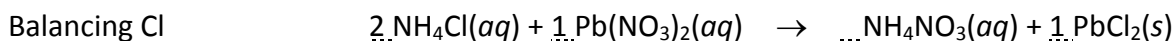
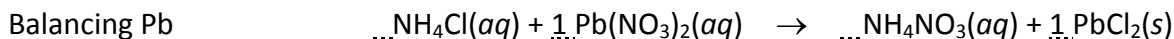
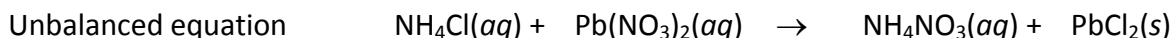


When writing a chemical equation for a reaction, the first thing is to make sure that the chemical formulae of the reactants and the products are correctly written. The next step is to balance the equation. Most chemical reactions can be balanced by simple inspection. Gaining experience in balancing equations is more important than following the sophisticated systematic steps. Nevertheless, some useful reminders for balancing equations are listed below:

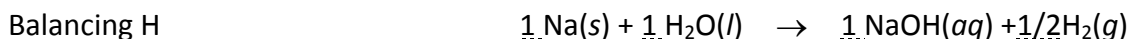
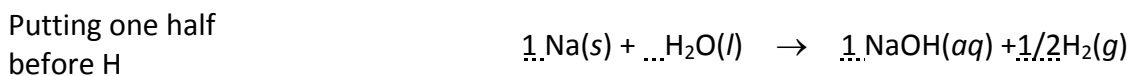
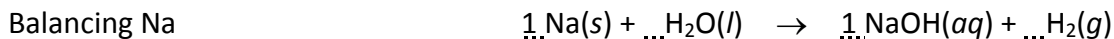
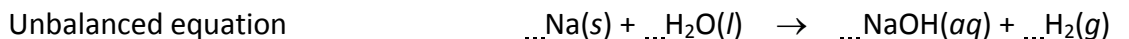
1. Manipulate only the coefficients when balancing equations. Never change the subscripts of the formulas because the subscripts are unique to each compound.
2. Leave the atoms that appear as an element to the very last for balancing. For example, the following equation contains an oxygen gas, which is an element. Oxygen should be chosen as the last element to balance.



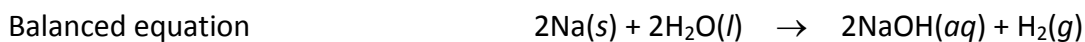
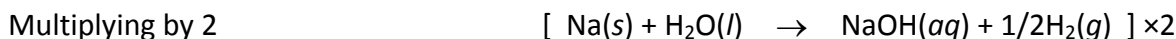
3. Polyatomic ions can be regarded as individual entities if they do not decompose in a reaction. In the following equation, there are polyatomic ions,  $\text{NH}_4^+$  and  $\text{NO}_3^-$ . They do not decompose and each of them is balanced on both sides. There is no need to check whether the numbers of nitrogen, hydrogen and oxygen atoms are balanced on both sides respectively.



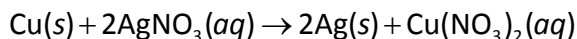
4. Sometimes, it is convenient to use fractions to balance chemical equations involving diatomic molecules. In the reaction of sodium metal and water, hydrogen gas is released as diatomic molecules. We can start guessing the coefficient of  $H_2$  to be one half to make balancing easier.



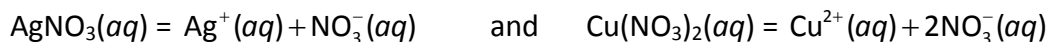
The smallest whole numbers can then be retrieved by multiplying the whole equation by 2. The previous equation thus turns into:



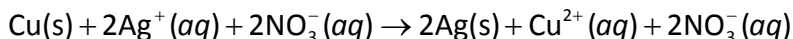
When chemical reactions take place in an aqueous solution, some compounds dissociate to produce ions. Not all ions take part in a chemical reaction though. These ions are called **spectator ions**. For example, in the below equation,



both  $AgNO_3$  and  $Cu(NO_3)_2$  are dissolved in water as stated by the state symbol, aq. They fully dissociate into the corresponding ions in water respectively such that the formulas,  $AgNO_3(aq)$  and  $Cu(NO_3)_2(aq)$ , actually imply the following:

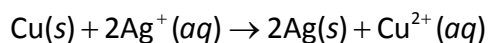


The full equation can be rewritten as the equation below which shows all the ions and thus is called the **complete ionic equation**:



The same amount of nitrate ions appears on both sides of the equation. This shows that nitrate ions are not necessarily involved in the reaction and thus are the spectator ions.

To make the chemical equations concise, the spectator ions are cancelled on both sides as shown below:



The resulting equation is called the *net ionic equation*.

Ionic equations require balancing of both the number of atoms and the number of charges. In the previous example, there are one copper atom and two silver atoms on both sides. For the charges, two  $\text{Ag}^+$  give rise to a total of 2 units of positive charges on the left while one  $\text{Cu}^{2+}$  ion yields the same amount of positive charges on the right. The charges are thus balanced. Not all ionic equations can be balanced by simple inspection. More complicated examples will be discussed in chapter 11.

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## Summary

1. Ionic compounds are named according to their cations and anions.
2. Covalent compounds are named according to the elements present.
3. Chemical equations give information on both the identity and the relative quantities of reactants and products.
4. A net ionic equation can be obtained by crossing out the spectator ions on both sides of a chemical equation.
5. A summary of ions to be memorized is given on the next two pages.

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## Key Words

Chemical equation  
Qualitative  
Quantitative  
Monatomic cations  
Polyatomic cations  
Monatomic anions

Polyatomic anions  
Coefficient  
Spectator ion  
Complete ionic equation  
Net ionic equation

## Tables of Cations

Charges					
+1		+2		+3	
Hydrogen	H <sup>+</sup>	Magnesium	Mg <sup>2+</sup>	Aluminum	Al <sup>3+</sup>
Lithium	Li <sup>+</sup>	Calcium	Ca <sup>2+</sup>	Chromium(III)	Cr <sup>3+</sup>
Sodium	Na <sup>+</sup>	Strontium	Sr <sup>2+</sup>	Iron(III)	Fe <sup>3+</sup>
Potassium	K <sup>+</sup>	Barium	Ba <sup>2+</sup>		
Rubidium	Rb <sup>+</sup>	Zinc	Zn <sup>2+</sup>		
Cesium	Cs <sup>+</sup>	Manganese(II)	Mn <sup>2+</sup>		
Silver	Ag <sup>+</sup>	Iron(II)	Fe <sup>2+</sup>		
Copper(I)	Cu <sup>+</sup>	Copper(II)	Cu <sup>2+</sup>		
Ammonium	NH <sub>4</sub> <sup>+</sup>	Tin(II)	Sn <sup>2+</sup>		
Hydronium	H <sub>3</sub> O <sup>+</sup>	Mercury(II)	Hg <sup>2+</sup>		
		Lead(II)	Pb <sup>2+</sup>		

## Tables of Anions

Charges					
-1		-2		-3	
Hydride	$\text{H}^-$	Oxide	$\text{O}^{2-}$	Nitride	$\text{N}^{3-}$
Fluoride	$\text{F}^-$	Sulfide	$\text{S}^{2-}$	Phosphide	$\text{P}^{3-}$
Chloride	$\text{Cl}^-$	Sulfite	$\text{SO}_4^{2-}$	Phosphite	$\text{PO}_3^{3-}$
Bromide	$\text{Br}^-$	Sulfate	$\text{SO}_4^{2-}$	Phosphate	$\text{PO}_4^{3-}$
Iodide	$\text{I}^-$	Carbonate	$\text{CO}_3^{2-}$		
Nitrite	$\text{NO}_2^-$	Hydrogen phosphate	$\text{HPO}_4^{2-}$		
Nitrate	$\text{NO}_3^-$	Oxalate	$\text{C}_2\text{O}_4^{2-}$		
Hypochlorite	$\text{ClO}^-$	Acetate	$\text{C}_2\text{H}_3\text{O}_2^-$		
Chlorite	$\text{ClO}_2^-$	Chromate	$\text{CrO}_4^{2-}$		
Chlorate	$\text{ClO}_3^-$	Dichromate	$\text{Cr}_2\text{O}_7^{2-}$		
Perchlorate	$\text{ClO}_4^-$				
Hydrogen carbonate	$\text{HCO}_3^-$				
Hydrogen sulfate	$\text{HSO}_4^-$				
Dihydrogen phosphate	$\text{H}_2\text{PO}_4^-$				
Hydroxide	$\text{OH}^-$				
Cyanide	$\text{CN}^-$				
Permanganate	$\text{MnO}_4^-$				

## Exercise on Nomenclature

Fill the table below.

Chemical Name	Formula
Silver oxide	
Iron(III) chloride	
Potassium permanganate	
Lead(II) sulfide	
Ammonium chloride	
Phosphorus pentachloride	
Lithium carbonate	
Hydroiodic acid	
Sodium silicate	
Copper(II) hydroxide	
Lithium phosphide	
Cesium fluoride	
Dioxygen difluoride	
Tin(II) bromide	
Manganese(II) oxide	
Carbon tetrachloride	
Sodium cyanide	
Barium sulfate	
Dinitrogen monoxide	
Magnesium chlorate	
Chromium(III) nitrate	
Calcium sulfite	
Lithium hydride	
Phosphoric acid	
Mercury(II) iodide	
Sodium hydrogencarbonate	
Aluminum nitride	

Chemical Name	Formula
	CaCl <sub>2</sub>
	H <sub>2</sub> SO <sub>4</sub>
	Fe <sub>2</sub> O <sub>3</sub>
	NH <sub>4</sub> NO <sub>3</sub>
	SF <sub>6</sub>
	N <sub>2</sub> O <sub>5</sub>
	Li <sub>3</sub> N
	SO <sub>3</sub>
	Al <sub>2</sub> (SO <sub>4</sub> ) <sub>3</sub>
	CO
	SnCl <sub>2</sub>
	NaClO
	AgI
	K <sub>2</sub> Cr <sub>2</sub> O <sub>7</sub>
	CuBr
	NaH <sub>2</sub> PO <sub>4</sub>
	CaSO <sub>4</sub>
	NCl <sub>3</sub>
	Sr(C <sub>2</sub> H <sub>4</sub> O <sub>2</sub> ) <sub>2</sub>
	Cl <sub>2</sub> O
	LiHCO <sub>3</sub>
	ZnS
	MgSO <sub>4</sub>
	Cr(OH) <sub>3</sub>
	S <sub>2</sub> Cl <sub>2</sub>
	CaCO <sub>3</sub>
	Cu(NO <sub>3</sub> ) <sub>2</sub>

## Review Questions

### Part A

Questions 1-4 refer to the following.

- (A)  $A_2B$
- (B)  $AB$
- (C)  $AB_2$
- (D)  $AB_3$
- (E)  $A_2B_3$

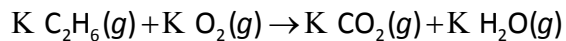
Which of the above represents the formula of the most common compound of elements A and B as shown in the pairs below?

	<u>A</u>	<u>B</u>
1.	Ca	Cl
2.	Al	O
3.	P	H
4.	H	S

### Part B

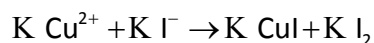
	<u>I</u>		<u>II</u>
101.	Copper(II) oxide is the chemical name of $Cu_2O$	BECAUSE	the formula contains two copper atoms.
102.	In writing a net ionic equation for the reaction of $NaOH(aq)$ and $HCl(aq)$ producing $NaCl(aq)$ and $H_2O(l)$ , $Na^+(aq)$ and $Cl^-(aq)$ can be neglected.	BECAUSE	$Na^+(aq)$ and $Cl^-(aq)$ are spectator ions.

## Part C



5. When the above chemical equation is balanced and all coefficients are reduced to the smallest whole numbers, the coefficient for  $\text{H}_2\text{O}$  is

- (A) 3
- (B) 4
- (C) 6
- (D) 8
- (E) 12



6. When the above chemical equation is balanced and all coefficients are reduced to the smallest whole numbers, which of the following statement(s) is/are correct?

- I. The coefficient for  $\text{Cu}^{2+}$  is 1.
- II. The coefficient for  $\text{I}^-$  is 3.
- III. The coefficient for  $\text{I}_2$  is 1.

- (A) I only
- (B) II only
- (C) III only
- (D) I and II only
- (E) I, II and III



## Answers and Explanation

### Answers to Exercise on Nomenclature

Chemical Name	Formula
Silver oxide	Ag <sub>2</sub> O
Iron(III) chloride	FeCl <sub>3</sub>
Potassium permanganate	KMnO <sub>4</sub>
Lead(II) sulfide	PbS
Ammonium chloride	NH <sub>4</sub> Cl
Phosphorus pentachloride	PCl <sub>5</sub>
Lithium carbonate	Li <sub>2</sub> CO <sub>3</sub>
Hydroiodic acid	HI
Sodium silicate	Na <sub>2</sub> SiO <sub>3</sub>
Copper(II) hydroxide	Cu(OH) <sub>2</sub>
Lithium phosphide	Li <sub>3</sub> P
Cesium fluoride	CsF
Dioxygen difluoride	O <sub>2</sub> F <sub>2</sub>
Tin(II) bromide	SnBr <sub>2</sub>
Manganese(II) oxide	MnO
Carbon tetrachloride	CCl <sub>4</sub>
Sodium cyanide	NaCN
Barium sulfate	BaSO <sub>4</sub>
Dinitrogen monoxide	N <sub>2</sub> O
Magnesium chlorate	Mg(ClO <sub>3</sub> ) <sub>2</sub>
Chromium(III) nitrate	Cr(NO <sub>3</sub> ) <sub>3</sub>
Calcium sulfite	CaSO <sub>3</sub>
Lithium hydride	LiH
Phosphoric acid	H <sub>3</sub> PO <sub>4</sub>
Mercury(II) iodide	HgI <sub>2</sub>
Sodium hydrogencarbonate	NaHCO <sub>3</sub>
Aluminum nitride	AlN

Chemical Name	Formula
Calcium chloride	CaCl <sub>2</sub>
Sulfuric acid	H <sub>2</sub> SO <sub>4</sub>
Iron(III) oxide	Fe <sub>2</sub> O <sub>3</sub>
Ammonium nitrate	NH <sub>4</sub> NO <sub>3</sub>
Sulfur hexafluoride	SF <sub>6</sub>
Dinitrogen pentoxide	N <sub>2</sub> O <sub>5</sub>
Lithium nitride	Li <sub>3</sub> N
Sulfur trioxide	SO <sub>3</sub>
Aluminum sulfate	Al <sub>2</sub> (SO <sub>4</sub> ) <sub>3</sub>
Carbon monoxide	CO
Tin(II) chloride	SnCl <sub>2</sub>
Sodium hypochlorite	NaClO
Silver iodide	AgI
Potassium dichromate	K <sub>2</sub> Cr <sub>2</sub> O <sub>7</sub>
Copper(I) bromide	CuBr
Sodium dihydrogen phosphate	NaH <sub>2</sub> PO <sub>4</sub>
Calcium sulfate	CaSO <sub>4</sub>
Nitrogen trifluoride	NCl <sub>3</sub>
Strontium acetate	Sr(C <sub>2</sub> H <sub>4</sub> O <sub>2</sub> ) <sub>2</sub>
Dichlorine monoxide	Cl <sub>2</sub> O
Lithium hydrogen carbonate	LiHCO <sub>3</sub>
Zinc sulfide	ZnS
Magnesium sulfate	MgSO <sub>4</sub>
Chromium(III) hydroxide	Cr(OH) <sub>3</sub>
Disulfur dichloride	S <sub>2</sub> Cl <sub>2</sub>
Calcium carbonate	CaCO <sub>3</sub>
Copper(II) nitrate	Cu(NO <sub>3</sub> ) <sub>2</sub>

## Answer to Review Questions

Question number	Answer	Explanation
1	C	Ca is a metal and Cl is a nonmetal. Therefore, they form an ionic compound. Ca is a group IIA element and it forms $\text{Ca}^{2+}$ while Cl is a group VIIA element and it forms $\text{Cl}^-$ . To balance the charges, the ratio of $\text{Ca}^{2+}$ to $\text{Cl}^-$ must be 1 : 2 and so the formula becomes $\text{CaCl}_2$ , which matches the answer C.
2	E	Al is a metal and O is a nonmetal. Therefore, they form an ionic compound. Al is a group IIIA element and it forms $\text{Al}^{3+}$ while O is a group VIA element and it forms $\text{O}^{2-}$ . To balance the charges, the ratio of $\text{Al}^{3+}$ to $\text{O}^{2-}$ must be 2 : 3 and so the formula becomes $\text{Al}_2\text{O}_3$ , which matches the answer E.
3	D	Both P and H are nonmetals. Therefore, they form a covalent compound. P is a group VA element which shares three electrons while H shares one electron to achieve the electron configuration of He. So one P atom can share electrons with three H atoms. The simplest formula is $\text{PH}_3$ , which matches the answer D.
4	A	Both H and S are nonmetals. Therefore, they form a covalent compound. H shares one electron to achieve the electron configuration of He while S is a group VIA element and it shares two electrons. So one S atom can share electrons with two H atoms. The simplest formula is $\text{H}_2\text{S}$ , which matches the answer A.
101	F, T	The correct name for $\text{Cu}_2\text{O}$ is copper(I) oxide because $\text{Cu}^+$ ions instead of $\text{Cu}^{2+}$ ions are found in $\text{Cu}_2\text{O}$ . The roman numeral represents the charge on an ion or the oxidation number of an atom, despite of the number of atoms in a chemical formula.
102	T, T, CE	Since $\text{Na}^+(\text{aq})$ ions can be found in both the reactant, $\text{NaOH}(\text{aq})$ , and the product, $\text{NaCl}(\text{aq})$ , they do not take part in the reaction and act as spectator ions. Likewise, $\text{Cl}^-(\text{aq})$ ions are also found in both the reactant and the product. Thus, both $\text{Na}^+(\text{aq})$ and $\text{Cl}^-(\text{aq})$ can be neglected in the net ionic equation.

Question number	Answer	Explanation
5	C	<p>We can balance the chemical equation in the following steps:</p> <p>Original equation      <math>\dots\text{C}_2\text{H}_6(g) + \dots\text{O}_2(g) \rightarrow \dots\text{CO}_2(g) + \text{H}_2\text{O}(g)</math></p> <p>Balancing C      <math>\underline{1}\text{C}_2\text{H}_6(g) + \dots\text{O}_2(g) \rightarrow \underline{2}\text{CO}_2(g) + \dots\text{H}_2\text{O}(g)</math></p> <p>Balancing H      <math>\underline{1}\text{C}_2\text{H}_6(g) + \dots\text{O}_2(g) \rightarrow \underline{2}\text{CO}_2(g) + \underline{3}\text{H}_2\text{O}(g)</math></p> <p>Balancing O      <math>\underline{1}\text{C}_2\text{H}_6(g) + \underline{7/2}\text{O}_2(g) \rightarrow \underline{2}\text{CO}_2(g) + \underline{3}\text{H}_2\text{O}(g)</math></p> <p>Multiplying by 2      <math>\underline{2}\text{C}_2\text{H}_6(g) + \underline{7}\text{O}_2(g) \rightarrow \underline{4}\text{CO}_2(g) + \underline{6}\text{H}_2\text{O}(g)</math></p> <p>Therefore, the coefficient of <math>\text{H}_2\text{O}</math> is 6. It should be noted that this is only one of the many methods in arriving at the same answer.</p>
6	C	<p>We should balance the atoms first:</p> <p>Original equation      <math>\dots\text{Cu}^{2+} + \dots\text{I}^- \rightarrow \dots\text{CuI} + \dots\text{I}_2</math></p> <p>Balancing Cu      <math>\underline{1}\text{Cu}^{2+} + \dots\text{I}^- \rightarrow \underline{1}\text{CuI} + \dots\text{I}_2</math></p> <p>Balancing I      <math>\underline{1}\text{Cu}^{2+} + \underline{3}\text{I}^- \rightarrow \underline{1}\text{CuI} + \underline{1}\text{I}_2</math></p> <p>However, we notice that the number of charges on the left equals to <math>+2 + (-1) \times 3 = -1</math> which is not neutral as shown on the right. Careful inspection on <math>\text{Cu}^{2+}</math> and <math>\text{I}^-</math> reveals that the ratio of <math>\text{Cu}^{2+}</math> to <math>\text{I}^-</math> must be 1: 2 to keep the reactant side neutral. Therefore, we can try the following:</p> <p>Balancing charges      <math>\underline{1}\text{Cu}^{2+} + \underline{2}\text{I}^- \rightarrow \dots\text{CuI} + \dots\text{I}_2</math></p> <p>Balancing Cu      <math>\underline{1}\text{Cu}^{2+} + \dots\text{I}^- \rightarrow \underline{1}\text{CuI} + \dots\text{I}_2</math></p> <p>Balancing I      <math>\underline{1}\text{Cu}^{2+} + \underline{2}\text{I}^- \rightarrow \underline{1}\text{CuI} + \underline{1/2}\text{I}_2</math></p> <p>Multiplying by 2      <math>\underline{2}\text{Cu}^{2+} + \underline{4}\text{I}^- \rightarrow \underline{2}\text{CuI} + \underline{1}\text{I}_2</math></p> <p>It should be noted that this is only one of the many methods in arriving at the same answer.</p> <p>Therefore, only option III is correct and the answer is C.</p>



## Chapter 4 – Nomenclature and Chemical Equations