أولمبياد الموهبة الوطني
NATIONAL GIFTED OLYMPIAD

Phase 2 Package

Scientific Track – NSMO

Chemistry

Chemistry Scientific Team

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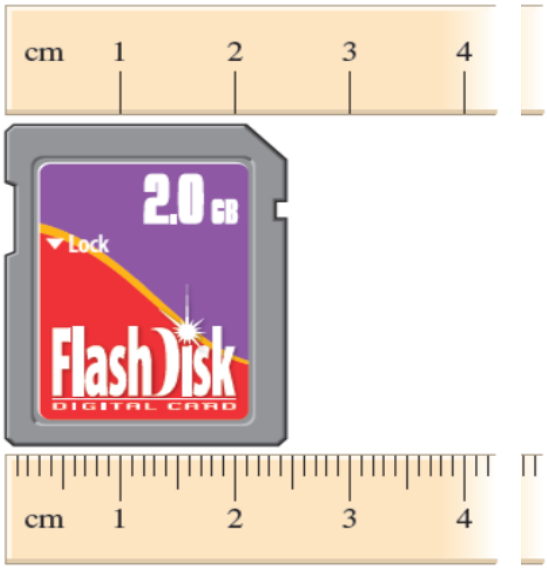
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01 Elementary Principles

Significant Figures

An inexact number must be reported in such a way as to indicate the uncertainty in its value, this is done using significant figures. Significant figures are the meaningful digits in a reported number.

Consider the measurement of the memory card "in the adjacent figure" using the ruler above it, The card's width is between 2 and 3 cm. We may record the width as 2.5 cm, but because there are no gradations between 2 and 3 cm on this ruler, we are estimating the second digit. Although we are certain about the number 2 in 2.5, we are not certain about the number 5.



The last digit in a measured number is referred to as the uncertain digit; and the uncertainty associated with a measured number is generally considered to be ± 1 in the decimal place of the last recorded digit. Thus, when we report the width of the memory card to be 2.5 cm, we are implying that its width is 2.5 ± 0.1 cm, and that its actual width may be as low as 2.4 cm or as high as 2.6 cm. Each of the digits in a measured number, including the uncertain digit, is a significant figure. The reported width of the memory card, 2.5 cm, contains two significant figures.

A ruler with millimeter gradations would enable us to be certain about the second digit in this measurement and to estimate a third digit. Now consider the measurement of the memory card using the ruler below it. We may record the width as 2.45 cm. Again, we estimate one digit beyond those we can read. The reported width of 2.45 cm contains three significant figures. Reporting the width as 2.45 cm implies that the width is 2.45 ± 0.01 cm.

Rules of significant figures

(A) Rules for determining how many significant figures are in a number:

(1) Non-Zero digits (1 – 9) and Zeros that are in between two non-zero digits are always significant.

(2) Leading zeroes are never significant.

(3) Trailing zeroes are only significant if a decimal point is present in the number.

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*Examples

	a. 809,231	b. 0.00456	c. 2300	d. 130.00
The explanation	Zero in between DOES count	Leading zeroes do NOT count	Trailing zeros do NOT count	Trailing zeros DO count befor the decimal
The number of significant figures	6	3	2	5

(B) Rules for performing Addition / Subtraction:

The final answer is written so that it has the same number of decimal places as the measurement that has the fewest decimal places.

*Examples

	a. 420.03 + 299.270 + 99.068	b. 504.09 - 246.8 - 119.32
The explanation	This number is the least precise (2 decimal places). So the answer MUST BE rounded to 2 decimal places.	Least Precise
Answer	818.37	138.0

(C) Rules for performing Multiplication / Division:

The final answer is written so that it has the same number of Significant Figures as the measurement with the fewest Significant Figures.

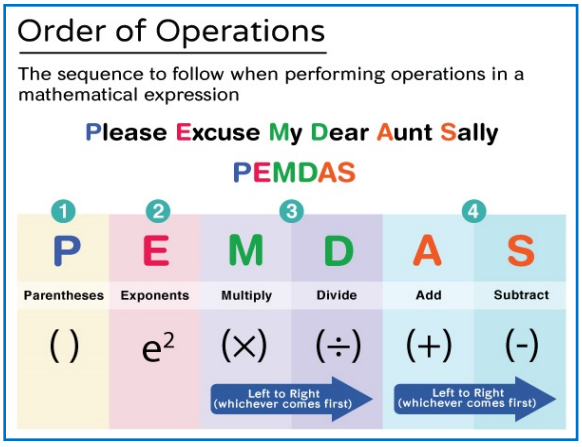
*Examples

	a. (2400)(3.45)(16.21)	b. 0.9935 × 10.48 × 13.4
The explanation	This number only has 2 Significant Figures , so the answer must be rounded to 2 Significant Figures .	Only has 3 Significant Figures, so answer needs rounded to 3 Significant Figures.
Answer	130000 or 1.3×10^5	140. or 1.40×10^2

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(D) Rules for Performing a Combination of Operations:

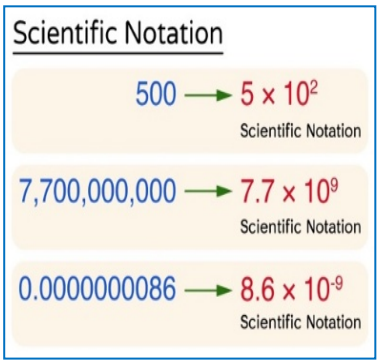
The previous rules listed for A/S and M/D still apply along with incorporating the rules for order of operations (PEMDAS). Be cautious, however, that you do not round the answer too early.



*Example

1.4 × 2.639 + 117.25	
Step 1: Multiply (1.4 × 2.639) + 117.25 (3.6946) + 117.25 Because this is multiplication, the answer to this part can only have 2 Significant Figures (3.7). Make note of this, but do NOT round yet.	(1.4 × 2.639) + 117.25 (3.7) + 117.25 Rounded too early. Should have been left as 3.6946
Step 2: Addition 3.6946 + 117.25 = 120.9946 = 120.9 1 2 decimal place	

Scientific notation

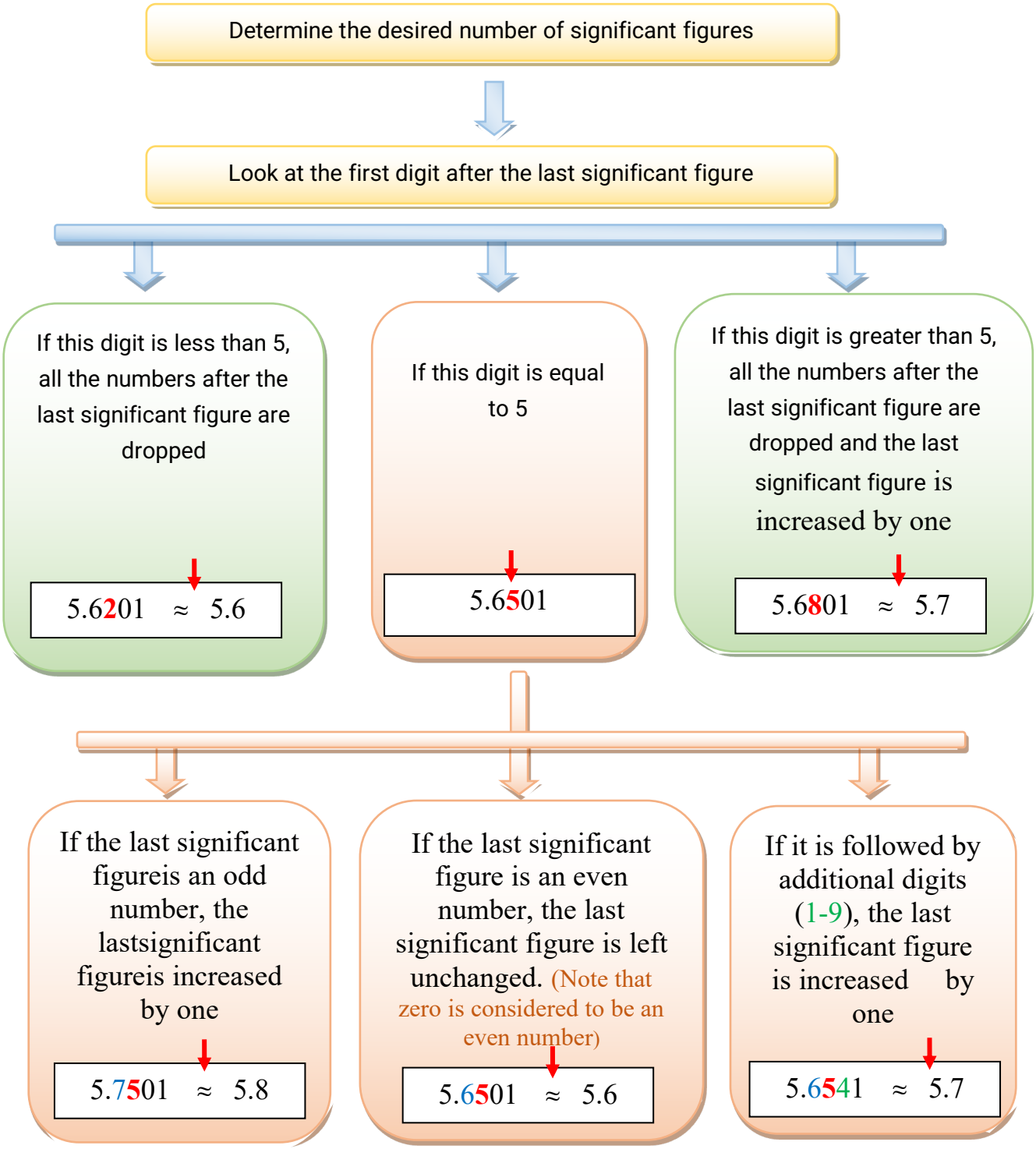


Chemists often deal with numbers that are either extremely large or extremely small. For example, in 1 g of the element hydrogen there are roughly 602,200,000,000,000,000,000,000 hydrogen atoms. Each hydrogen atom has a mass of only 0.00000000000000000000000166 g. These numbers are cumbersome to handle, and it is easy to make mistakes when using them in arithmetic computations. Consequently, when working with very large and very small numbers, we use a system called scientific notation. Regardless of their magnitude, all numbers can be expressed in the form: $N\ 10^n$

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Rounding Off

It is usually necessary to round off the final answer so that it has the proper number of significant figures. To round off a number, follow the procedure below.



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Fundamental and derived units of measurement



International System of Units (SI):

The International System of Units (SI) is the most widely used system of measurement units in the world.

*Examples of Fundamental Units

Quantity	SI unit	Symbol
Length	meter	m
Mass	kilogram	kg
Time	second	s
Temperature	kelvin	K
Electric current	ampere	A
Amount of substance	mole	mol
Luminous intensity	candela	cd

*Examples of Derived Units

Quantity	SI unit	Symbol
Area	square meter	m^2
Volume	cubic meter	m^3
Density	kilogram per cubic meter	kg/m^3
Force	newton	$N = kg.m/s^2$
Energy	joule	J = N.m



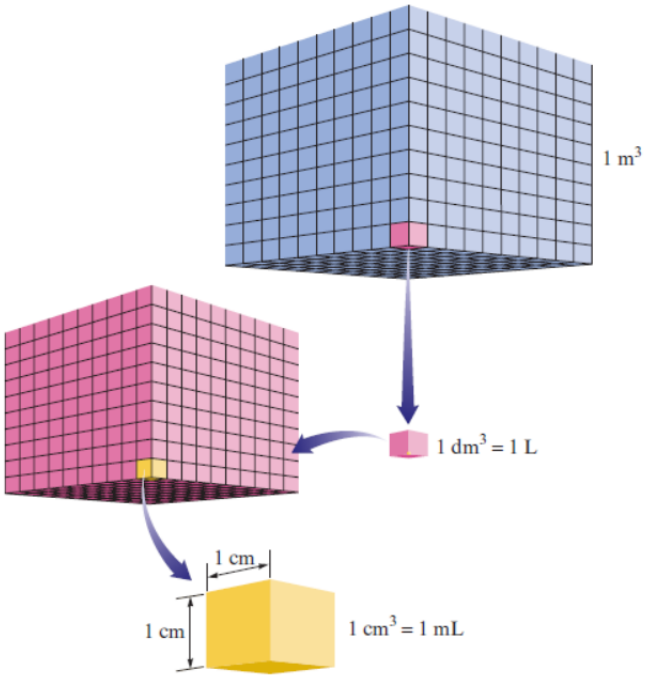
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Prefixes Used in the SI System

Prefix	Symbol	Meaning	Exponential Notation*
exa	E	1,000,000,000,000,000,000	10^{18}
peta	P	1,000,000,000,000,000	10^{15}
tera	T	1,000,000,000,000	10^{12}
giga	G	1,000,000,000	10^{9}
mega	M	1,000,000	10^{6}
kilo	k	1,000	10^{3}
hecto	h	100	10^{2}
deka	da	10	10^{1}
—	—	1	10^{0}
deci	d	0.1	10^{-1}
centi	c	0.01	10^{-2}
milli	m	0.001	10^{-3}
micro	μ	0.000001	10^{-6}
nano	n	0.000000001	10^{-9}
pico	p	0.000000000001	10^{-12}
femto	f	0.000000000000001	10^{-15}
atto	a	0.000000000000000001	10^{-18}

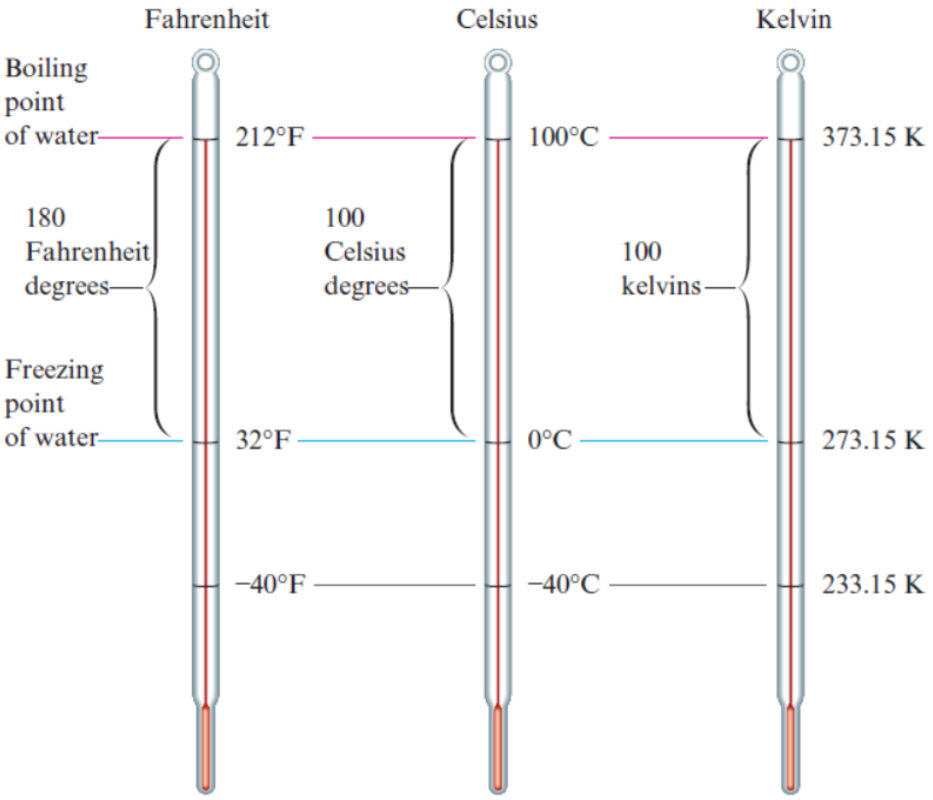
The most commonly encountered prefixes are shown in blue

$1\ L = 1\ (dm)^3 = (10\ cm)^3 = 1000\ cm^3$



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$K = °C + 273.15$



***Examples:**

Example 1-1: convert 2.3 cm^3 to m^3 unit.

$2.3\ cm^3 = 2.3\ (cm)^3 = 2.3\ (10^{-2}\ m)^3 = 2.3 \times 10^{-6}\ m^3$

Example 1-2: convert 5000 μs^{-1} to s^{-1} unit.

$5000\ \mu s^{-1} = 5000\ (\mu s)^{-1} = 5000\ (10^{-6}\ s)^{-1} = 5 \times 10^{9}\ s^{-1}$

Example 1-3: Speed of sound = 330 m/s, convert the unit to km/h.

$$330\ m/s = 330\ \frac{\cancel{m}}{\cancel{s}} \times \frac{1\ km}{1000\ \cancel{m}} \times \frac{3600\ \cancel{s}}{1\ h} = 1188\ km/h$$

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Exercises

(1-1)

State the number of significant digits in each measurement:

1) 2804 m	2) 2.84 km	3) 5.029 m
.....	.....	.....
4) 0.003068 m	5) 4.6×10^5 m	4) 0.003068 m
.....	.....	.....
7) 750 m	8) 75 m	9) 75,000 m
.....	.....	.....
10) 75.00 m	11) 75,000.0 m	12) 10 cm
.....	.....	.....

(1-2)

Round the following numbers as indicated:

(a) To four figures:

3.682417	21.860051	375.6523	112.511	45.4673
.....	.....	.....	.....	.....

(b) To one decimal place:

1.3511	2.473	5.687524	7.555	8.235
.....	.....	.....	.....	.....

(c) To two decimal places:

22.494	79.2588	0.03062	3.4125	41.86632
.....	.....	.....	.....	.....

(1-3)

Solve the following problems and report answers with appropriate number of Significant Figures.

1) 6.201 cm + 7.4 cm + 0.68 cm + 12.0 cm = ...

2) 1.6 km + 1.62 m + 1200.0 cm = ...

3) 8.264 g - 7.8 g = ...

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4) 10.4168 m - 6.0 m = ...

5) 12.00 m + 15.001 kg = ...

6) 1.31 cm × 2.3 cm = ...

7) 5.7621 m × 6.201 m = ...

8) 20.2 cm: 7.410 s = ...

9) 40.002 g / 13.000005 g = ...

(1-4)

Express the following numbers in their equivalent scientific notation:

1) 123,876.3

.....

2) 1,236,840

.....

3) 422000

.....

4) 0.000000000000211

.....

5) 0.000238

.....

6) 0.0000205

.....

(1-5)

Perform the following mathematical operations and express the result to the correct number of Significant Figures.

1) $\dfrac{(1.00866 - 1.00728)}{6.02205 \times 10^{23}}$ = ...

2) $\dfrac{(14.86 + 13.7) \times (65.346 - 4.10)}{(43.888 - 32.888)}$ = ...

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(1-6)

The floor area of a science laboratory is 60 m^2. Find the area in cm^2 and mm^2.

.....

.....

.....

.....

(1-7)

The density of Mercury is 13.6 g/cm^3. Convert its unit to Mg/mm^3.

.....

.....

.....

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(1-8)

The density of sea water is 1024 $kg.m^{-3}$. Find the density in $g.cm^{-3}$.

.....

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.....

.....

(1-9)

In a measurement of air pollution, air was drawn through a filter at the rate of 26.2 liters per minute for 48.0 hours. The filter gained 0.0241 grams in mass because of entrapped solid particles. Express the concentration of solid contaminants in the air in units of micrograms per cubic meter.

.....

.....

.....

.....

(1-10)

NewYork City's 7.9 million people in 1978 had a daily per capita consumption of 656 liters of water. How many metric tons (10^3 kg) of sodium fluoride (45% fluorine by mass) would be required per year to give this water a tooth strengthening dose of 1 part (by mass) fluorine per million parts water ?

(The density of water is 1.000 g/cm^3 or 1.000 kg/L).

.....

.....

.....

.....

.....

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Exercise answers

(1-1)

1) 2804 m
.....4.....

2) 2.84 km
.....3.....

3) 5.029 m
.....4.....

4) 0.003068 m
.....4.....

5) 4.6×10^5 m
.....2.....

4) 0.003068 m
.....3.....

7) 750 m
...2 or 3.....

8) 75 m
.....2.....

9) 75,000 m
...2, 3, 4, or 5.....

10) 75.00 m
.....4.....

11) 75,000.0 m
.....6.....

12) 10 cm
.....1 or 2.....

(1-2)

(a) To four figures:

3.682417	21.860051	375.6523	112.511	45.4673
...3.682.....	...21.86....	...375.7....	... 112.5....	...45.47....

(b) To one decimal place:

1.3511	2.473	5.687524	7.555	8.235
...1.4.....	.....2.5....	... 5.7.....	...7.6.....	...8.2.....

(c) To two decimal places:

22.494	79.2588	0.03062	3.4125	41.86632
.....22.49....	...79.26.....	...0.03.....	...3.41.....	...41.87.....

(1-3)

1) 6.201 cm + 7.4 cm + 0.68 cm + 12.0 cm = 26.3 cm

2) 1.6 km + 1.62 m + 1200.0 cm = 1.6 km or 1613.6 m or 161362.0 cm

3) 8.264 g - 7.8 g = 0.5 g

4) 10.4168 m - 6.0 m = 4.4 m

5) 12.00 m + 15.001 kg = can't add m and kg

6) 1.31 cm × 2.3 cm = 3.0 cm^2

7) 5.7621 m × 6.201 m = 35.73 m^2

8) 20.2 cm: 7.410 s = 2.73 cm/s

9) 40.002 g / 13.000005 g = 3.0771

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(1-4)

1) 123,876.3
1.238763×10^5

2) 1,236,840
1.23684×10^6

3) 422000
4.22×10^5

4) 0.000000000000211
2.11×10^{-13}

5) 0.000238
2.38×10^{-4}

6) 0.0000205
2.05×10^{-5}

(1-5)

1) $$\frac{(1.00866 - 1.00728)}{6.02205 \times 10^{23}} = 2.29 \times 10^{-27}$$

2) $$\frac{(14.86 + 13.7) \times (65.346 - 4.10)}{(43.888 - 32.888)} = 159$$

(1-6)

$1\ m = 10^2\ cm\ (100\ cm)$

$1\ m^2 = (10^2\ cm)^2 = 10^4\ cm^2$

$$60\ m^2 = 60\ \cancel{m}^2 \times \frac{10^4\ cm^2}{1\ \cancel{m}^2} = 60 \times 10^4\ cm^2 = 6 \times 10^5\ cm^2$$

and $1\ cm = 10\ mm$

$1\ cm^2 = (10\ mm)^2 = 10^2\ mm^2$

$$60\ m^2 = 6 \times 10^5\ \cancel{cm}^2 \times \frac{10^2\ mm^2}{1\ \cancel{cm}^2} = 6 \times 10^7\ mm^2$$

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(1-7)

$1\ Mg = 10^6\ g\ ,\ 1\ cm^3 = (10\ mm)^3 = 10^3\ mm^3$

$$13.6\ g/cm^3 = 13.6\ \frac{\cancel{g}}{\cancel{cm}^3} \times \frac{1\ Mg}{10^6\ \cancel{g}} \times \frac{1\ \cancel{cm}^3}{10^3\ mm^3} = 13.6 \times 10^{-9}\ Mg/mm^3$$

(1-8)

$1\ kg = 10^3\ g\ ,\ 1\ m^3 = (10^2\ cm)^3 = 10^6\ cm^3$

$$1024\ kgm^{-3} = 1024\ \frac{\cancel{kg}}{\cancel{m}^3} \times \frac{10^3\ g}{1\ \cancel{kg}} \times \frac{1\ \cancel{m}^3}{10^6\ cm^3} = 1.024\ gcm^{-3}$$

(1-9)

$$\frac{(0.0241\ g)\ (10^6 \mu g/1\ g)}{(48.0\ h)(60\ min/h)(\ 26.2\ L/1\ min)(1\ dm^3/1\ L)(1\ m^3/10^3\ dm^3)} = 319\ \mu g/m^3$$

(1-10)

A good start is to calculate the mass of water, in tons, required per year.

$(7.9 \times 10^6$ persons) (656 L water /1 person · day) (365 days/year) (1 kg water/1 L water) (1 metric ton/1000 kg) $= 1.89 \times 10^9$ metric tons water / year .

Note that all units cancel except metric tons of water/yr; it is needed for the next step. Now, set up and calculate the total mass of sodium fluoride, in tons, required each year.

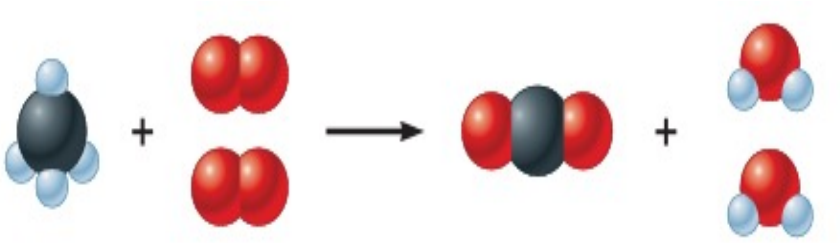
(1.89 $\times 10^9$ metric tons water / year) (1 ton fluorine/10^6 tons of water) (1 ton sodium fluoride/0.45 ton of fluorine) $= 4.2 \times 10^3$ tons of sodium fluoride / year .

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02 Chemical Reactions

A chemical reaction is a process in which a set of starting materials (reactants) is converted into new materials (products) that are completely different in their chemical and physical properties from the original materials.

For example, when methane burns, the chemical bonds in the reactant molecules are broken, and new chemical bonds are formed to rearrange the atoms and produce the product molecules. The key point is that the types of atoms do not change (a carbon atom remains carbon, and a hydrogen atom remains hydrogen); what changes is how they are bonded to one another.



The Importance of Chemical Reactions in Our Lives

Chemical reactions are an integral part of our daily lives and the universe. Their importance includes:

- **biological processes:** such as photosynthesis in plants to produce food and oxygen, the digestive process in our bodies to convert food into energy, and cellular respiration to produce energy.
- **Industry and technology:** production of medicines, fuels (combustion of gasoline in car engines), plastics, fertilizers, and detergents.
- **Environment:** such as the rusting reaction (the reaction of iron with oxygen), and natural element cycles such as the carbon and water cycles.

Signs of a Chemical Reaction:

1) **Qualitative evidence (observed by the senses or simple experiments).**
   a) Color change.
      - Example: When potassium permanganate solution is mixed with some reducing agents, the purple color disappears.
      - Explanation: A change in the composition of ions/molecules results in a difference in light absorption.
   b) It smells new or different.
      - Example: The sulfurous smell resulting from the reaction of some sulfides with acids (H_2S).
      - Explanation: Production of gases or volatile compounds with a distinctive odor.

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c) gas formation (gas release).

- Example: Effervescence occurs when sodium bicarbonate is mixed with acid, forming CO_2. This can be detected by passing the gas through limewater ($Ca(OH)_2$), where a white turbidity is formed, indicating the presence of CO_2.
- Explanation: The formation of gaseous molecules resulting from the rearrangement of atoms.

d) Precipitate (a solid that settles out).

- Example: Mixing sodium sulfate solution with sodium chloride solution does not produce a precipitate, but mixing sodium sulfate with barium nitrate produces a white precipitate ($BaSO_4$).
- Explanation: Formation of a new compound with low solubility.

e) Change in energy (change in heat, light, sound).

- Example: Emission of light or sparks: Combustion or strong oxidation reactions.
- Explanation: Change in bond energy — forming new bonds releases energy or absorbing energy to break bonds.

f) Change in electrical conductivity.

- Example: Dissolving an acid or base or a reaction that results in the formation of ions increases or decreases the electrical conductivity of the solution.
- Explanation: The number of Ions in the solution changes.

2) Quantitative evidence.

a) Mass change measurement.

- In a mass-closed reaction, the total mass does not change (law of conservation of mass).
- In open experiments: a loss of mass indicates the release of a gas; a gain in mass may indicate the absorption of a gas (e.g., a metal absorbing oxygen from the air in an oxidation reaction).

b) Heat of reaction measurements.

- Using a calorimeter, the heat released or absorbed can be accurately measured, confirming that an energy transfer has occurred associated with the formation or breaking of bonds.

c) Concentration measurements.

- Tracking the decrease in concentration of one substance and the formation of another using spectroscopy, chromatography, or electrical conductivity measurement provides strong evidence that a reaction has occurred, showing the amount of product produced and the rate of the reaction.

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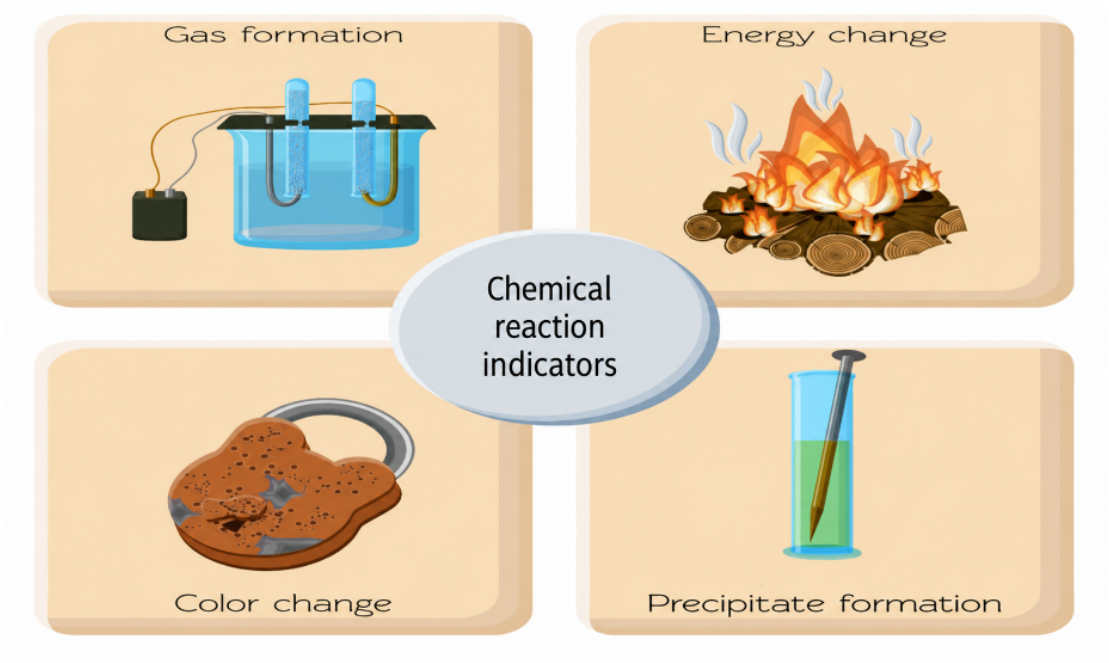


Figure 2-1: Some examples of chemical reactions

Representing chemical reactions

Chemical equations are the basic and most effective tool for representing chemical reactions. They are a universal and concise language used by chemists to describe what happens at the molecular level during a reaction.

The purpose of representing a chemical reaction is:

Representing a chemical reaction is writing a symbolic description of what happens during a reaction, using element symbols, chemical formulas, and arrows instead of writing the reaction out in words.

The basic components of a chemical equation

1) Reactants

These are the starting materials that start the reaction. They are written to the left of the arrow (or to the left of the equal sign in some cases).

2) Products

The new substances that result from the reaction. They are written on the right side of the arrow (or to the right of the equal sign in some cases).

3) Arrow

The arrow (⟶) points from reactants to products, and is usually read as "produces" or "reacts to give."

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Symbols used in equations

Symbol	Meaning	Example
→	Produces or gives. Separates reactants from products and indicates an irreversible, one-way reaction.	A + B → C
⇌	Reversible reaction. Indicates that the reaction reaches equilibrium and occurs in both forward and reverse directions.	A + B ⇌ C + D
+	Reacts with or in addition to. Separates reactants and/or products.	H_2 + O_2
(s)	Refers to the solid state.	NaCl(s)
(l)	Refers to the liquid state.	H_2O(l)
(g)	Refers to the gaseous state.	CO_2(g)
(aq)	Refers to an aqueous solution.	HCl(aq)
Δ or heat	Placed above the arrow. Indicates that the reaction requires heating to occur.	$CaCO_3 \xrightarrow{\Delta} CaO + CO_2$
Pressure or pressure value	Placed above the arrow. Indicates that the reaction requires a specific pressure.	$N_2 + 3H_2 \xrightarrow{200\ atm} 2NH_3$
Catalyst, such as Pt	Placed above the arrow. Indicates the use of a catalyst to speed up the reaction without being consumed.	$C_2H_4 + H_2 \xrightarrow{Pt} C_2H_6$
↓ (Precipitate)	Placed after a chemical formula. Indicates the formation of a precipitate, an insoluble solid, in the products.	AgCl↓
↑ (Gas evolved)	Placed after a chemical formula. Indicates the evolution of a gas in the products.	CO_2↑

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Word chemical equations

A word equation is an expression written in words to show the reactants and products of a chemical reaction without using chemical symbols or formulas.

reactant (1) + reactant (2) ⟶ product(1) + product(2)

Example 1:

Methane reacts with oxygen to produce water and carbon dioxide.

Methane +oxygen ⟶ carbon dioxide + water

Exercise (2-1):

Write the word equation for the reaction of bromine gas with hydrogen to produce hydrogen bromide gas.

.....

.....

.....

Exercise (2-2):

Write the word equation for the heat decomposition of ammonium nitrate to form nitrous oxide (laughing gas) and water vapor.

.....

.....

.....

Exercise (2-3):

When solid ammonium nitrate is heated, it decomposes into various gases. This decomposition produces nitrous oxide (laughing gas) and water vapor.

The dinitrogen oxide gas is then passed over a hot magnesium ribbon, producing magnesium oxide and nitrogen gas.Which of the following word equations correctly represents the complete series of these two reactions?

	The equation
a)	ammonium nitrate ⟶ dinitrogen oxide + water vapor Then dinitrogen oxide + magnesium ⟶ magnesium nitride + water vapor
b)	ammonium nitrate ⟶ nitric oxide + oxygen Then nitric oxide + magnesium ⟶ magnesium oxide + nitrogen
c)	Ammonium nitrate ⟶ dinitrogen oxide + water vapor Then dinitrogen oxide + magnesium ⟶ magnesium oxide + nitrogen
d)	ammonium nitrate ⟶ nitrogen dioxide + water Then nitrogen dioxide + magnesium ⟶ magnesium nitride + oxygen

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Symbolic chemical equations

Symbolic chemical equations are the most accurate and scientific way to represent chemical reactions. These equations use symbols and formulas for elements and compounds, along with numbers written before them called coefficients, to show what is happening at the molecular level.

Coefficient: The number written before a reactant or product. Coefficients are usually integers and are not written if the value is one. Coefficients describe the simplest whole number ratio of the quantities of the reactants and products.

Balancing chemical equations

Balancing chemical equations means making the number of atoms of each element equal on both sides of the equation (reactants and products), so that the law of conservation of mass is achieved, which states that mass can neither be created nor destroyed, but rather is transformed from one form to another.

The importance of balancing

The most important property of a symbolic equation is that it must be balanced. Figure 2-2 shows the steps for balancing chemical equations.

- **Definition:** Balancing an equation means adjusting the coefficients to ensure that the number of atoms of each element is equal on both sides of the equation (the reactants and the products).
- **Goal:** To achieve the law of conservation of mass, which states that mass is neither created nor destroyed in a chemical reaction.
- **Method:** Only the coefficients are changed; subscripts in the formula must not be changed, as this changes the identity of the substance.

Steps for balancing chemical equations

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Steps for Balancing Chemical Equations

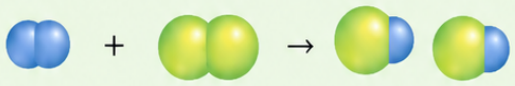
Step	Process	Example
1	**Write the unbalanced chemical equation.** Make sure to include the correct formulas for all reactants and products. Identify the different elements and count the number of atoms of each element in the reactants and products.	$H_{2(g)} + Cl_{2(g)} \rightarrow HCl_{(g)}$  Hydrogen molecule — Chlorine molecule — Hydrogen chloride molecule
2	**Count the atoms of each element in the reactants.** In this example, 2 hydrogen atoms and 2 chlorine atoms.	$H_2 + Cl_2 \rightarrow$ 2 H atoms — 2 Cl atoms
3	**Count the atoms of each element in the products.** In this example, 1 hydrogen atom and 1 chlorine atom.	HCl 1 H atom 1 Cl atom
4	**Change the coefficients to make the number of atoms of each element the same on both sides of the equation.** Do not change the subscripts in the chemical formulas, as this would change the identity of the substance.	$H_2 + Cl_2 \rightarrow 2HCl$ 2 H atoms — 2 Cl atoms — 2 H atoms + 2 Cl atoms  2 H atoms — 2 Cl atoms — 2 H atoms, 2 Cl atoms
5	**Write the balanced equation with the smallest possible whole-number coefficients.** The coefficients should be in the simplest possible ratio. In this example, $2HCl : 1Cl_2 : 1H_2$ is already the simplest ratio, because it cannot be divided further without making the coefficients non-integers.	$H_{2(g)} + Cl_{2(g)} \rightarrow 2HCl_{(g)}$ $1H_2 : 1Cl_2 : 2HCl$ $1 : 1 : 2$
6	**Check your work.** Make sure the chemical equation is balanced correctly and that the number of atoms of each element is the same on both sides.	$H_{2(g)} + Cl_{2(g)} \rightarrow 2HCl_{(g)}$ 2 H atoms — 2 Cl atoms — 2 H atoms, 2 Cl atoms The number of hydrogen atoms and chlorine atoms is the same on both sides of the equation.

Figure 2-2: Steps for balancing chemical equations

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Example 2:

Write the balanced chemical symbol equation for the reaction of iron(III) chloride with sodium hydroxide in water to produce solid iron(III) hydroxide and sodium chloride.

Souliotion:

Textual data:

Iron(III) chloride + sodium hydroxide → iron(III) hydroxide + sodium chloride.

Writing the correct chemical formulas for substances:

The Substance	Chemical Formula
Iron(III) chloride	$FeCl_3$
sodium hydroxide	NaOH
Iron(III) hydroxide	$Fe(OH)_3$
sodium chloride	NaCl

Writing the unbalanced equation:

$$FeCl_3 + NaOH \rightarrow Fe(OH)_3 + NaCl$$

Balancing the equation:

1. Iron (Fe):

1 in reactants and 1 in products → ✔ Balanced.

2. Chlorine (Cl):

3 in $FeCl_3$ → We need 3 in NaCl → We put 3 in front of NaCl.

3. Sodium (Na):

3 in NaCl → So we put 3 in front of NaOH in the reactants.

4. Hydroxide (OH):

We now have 3 in $(OH)_3$ in the reactants = 3 in $Fe(OH)_3$ → ✔ balanced.

Final balanced equation:

$$FeCl_{3(aq)} + 3NaOH_{(aq)} \rightarrow Fe(OH)_{3(s)} + 3NaCl_{(aq)}$$

Exercise (2-4):

Write the word equation and the symbol chemical equation for the following reaction: When solid potassium chlorate is heated, solid potassium chloride and oxygen gas are produced.

.....

.....

.....

.....

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Exercise (2-5):

Express the following word equations as balanced symbolic equations:

calcium carbonate + hydrochloric acid → calcium chloride + water + CO_2
.....
potassium hydroxide + sulfuric acid → potassium sulfate + water
.....
tin(II) sulfate + potassium phosphate → tin(II) phosphate + potassium sulfate
.....
aluminum + hydrochloric acid → aluminum chloride + hydrogen gas
.....

Exercise (2-6):

Select the balanced equation from the following options:

a) $2Al + 6HCl \rightarrow 3AlCl_3 + 3H_2$

b) $Ba(ClO_3)_2 \rightarrow BaCl_2 + 2O_2$

c) $2Li_3PO_4 + 3CaCl_2 \rightarrow 6LiCl + Ca_3(PO_4)_2$

d) $C_5H_{12} + 6O_2 \rightarrow 3CO_2 + 5H_2O$

Exercise (2-7):

Balance the following equations:

a)

……ZnS + ……HCl → ……$ZnCl_2$ + ……H_2S				
Element	Zn	S	H	Cl
Reactants				
Product				

b)

……HCl + ……Cr → ……$CrCl_3$ + ……H_2			
Element			
Reactants			
Product			

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c)

……$FeCl_3$ + ……Na_2CO_3 → ……$Fe_2(CO_3)_3$ + ……NaCl				
Element				
Reactants				
Product				

d)

……Na_2O + ……$(NH_4)_2SO_4$ → ……Na_2SO_4 + ……H_2O + ……NH_3					
Element					
Reactants					
Product					

Exercise (2-8):

Balance the following equations:

1) .....$C_2H_4O_2$ + ...O_2 → .....CO_2 + ...H_2O
2) .....$Mn(NO_2)_2$ + ...$BeCl_2$ → .....$Be(NO_2)_2$ + ...$MnCl_2$
3) .....AgBr + ...$GaPO_4$ → .....Ag_3PO_4 + ...$GaBr_3$
4) .....H_2SO_4 + ...$B(OH)_3$ → .....$B_2(SO_4)_3$ + ...H_2O
5) .....$NaNO_3$ + ...PbO → .....$Pb(NO_3)_2$ + ...Na_2O
6) .....$Ca_3(PO_4)_2$ + ... H_2SO_4 → .....$CaSO_4$ + ...H_3PO_4
7) .....Al + ...H_2SO_4 → .....$Al_2(SO_4)_3$ + ... H_2
8) .....Al_4C_3 + ... H_2O → .....$Al(OH)_3$ + ... CH_4
9) .....Na + ... H_2O → ..... NaOH + ...H_2

Types of Chemical Reactions

Chemical reactions vary into several main types, including:

1) **Combustion reactions:** It is a type of rapid chemical reaction that occurs between a reactant and an oxidizing agent (usually atmospheric oxygen), producing energy.

Example 3: the complete combustion of methane (natural gas)

$$CH_{4(g)} + 2O_{2(g)} \longrightarrow CO_{2(g)} + 2H_2O_{(g)} + \text{Energy}$$

2) **Formation reactions:** the combination of two or more substances to form a new compound.

   The product is always just one compound. This is the clearest indication that a reaction is a synthesis reaction. These are often reactions that release energy in the form of heat and light. This makes sense because bond formation is often accompanied by the release of energy.

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General form of the equation:

$$A + B \longrightarrow AB$$

Formation reactions can occur between:

- Two elements to form a compound:

Example 4: Hydrogen reacts with oxygen to form water.

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

- Element and compound combine to form a new compound:

Example 5: Sulfur dioxide reacts with oxygen to form sulfur trioxide.

$$2SO_{2(g)} + O_{2(g)} \longrightarrow 2SO_{3(g)}$$

- Two compounds combine to form a new, more complex compound.

Example 6: Calcium oxide (quicklime) reacts with water to form calcium hydroxide (slaked lime).

$$CaO_{(s)} + H_2O_{(l)} \longrightarrow Ca(OH)_{2(s)}$$

3) **Decomposition reactions:** It is a chemical reaction in which one compound breaks down into two or more simpler substances (either elements or simpler compounds). Simply put, it is a "breaking" or "decomposition" reaction.

General form of the equation:

$$AB \longrightarrow A + B$$

Because decomposition reactions involve the breaking of chemical bonds, they often require the addition of energy. This energy can take various forms, such as heat, light, or electricity.

Example 7: Ammonium nitrate decomposes into nitrous oxide and water.

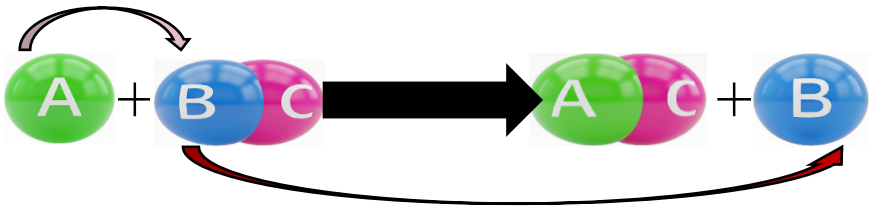
$$NH_4NO_{3(s)} \longrightarrow N_2O_{(g)} + 2H_2O_{(g)}$$

4) **Simple substitution reactions:** substitution of one element for another in a compound.

General form of the equation:

$$A + BC \longrightarrow AC + B$$

In this type of reaction, an active (more reactive) element reacts with a compound, replacing one of the components of that compound (the less reactive element). The result is the formation of a new compound and a free element.



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Types of simple substitution reactions:

- **Substitution of one metal for another:**
  It occurs when the free metal is more reactive than the metal in the compound.

Example 8:

$$Zn + CuSO_4 \rightarrow ZnSO_4 + Cu$$

- **Substitution of hydrogen from acids by a metal:**

Some active metals displace hydrogen in acids.

Example 9:

$$Mg + 2HCl \rightarrow MgCl_2 + H_2$$

- **Substitution of a nonmetal for another nonmetal:**

It often occurs among halogens (e.g., F_2, Cl_2, Br_2, I_2).

Example 10:

$$Cl_2 + 2NaBr \rightarrow 2NaCl + Br_2$$

A reaction only occurs if the free element is more reactive than the element it will replace. The order of reactivity is known as the chemical activity series. As shown in Figure 2-3.

Metal activity series																				
The lowest		→																	**The highest**	
Au	**Ag**	**Cu**	**Pb**	**Sn**	**Ni**	**Co**	**Fe**	**Cr**	**Zn**	**Mn**	**Ti**	**Al**	**Mg**	**Na**	**Ca**	**Sr**	**Ba**	**Cs**	**K**	**Rb**	**Li**
Nonmetal activity series																				
The lowest		→																	**The highest**	
																	I_2	Br_2	Cl_2	F_2

Figure 2-3: Activity sequence of some elements

Exercise (2-9):

Which of the following reactions occur between halogens and halide salts?

a) $F_{2(g)} + FeI_{2(aq)} \rightarrow FeF_{2(aq)} + I_{2(s)}$

b) $I_{2(s)} + MnBr_{2(aq)} \rightarrow MnI_{2(aq)} + Br_{2(l)}$

c) $Cl_{2(g)} + SrF_{2(aq)} \rightarrow SrCl_{2(aq)} + F_{2(g)}$

d) $Br_{2(l)} + CoCl_{2(aq)} \rightarrow CoBr_{2(aq)} + Cl_{2(g)}$

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Exercise (2-10):

Predict whether each of the following reactions will occur, indicate the products formed, and then weigh them:

1)	$Mg_{(s)} + ZnCl_{2(aq)} \rightarrow$
2)	$F_{2(g)} + 2HCl_{(aq)} \rightarrow$
3)	$Al_{(s)} + AlCl_{3(aq)} \rightarrow$
4)	$Fe_{(s)} + Na_3PO_{4(aq)} \rightarrow$
5)	$2Al_{(s)} + 3Pb(NO_3)_{2(aq)} \rightarrow$

Exercise (2-11):

In a series of laboratory experiments, the following tests were conducted using three unknown metals: X, Y, and Z, and solutions of their salts: XCl_2, YCl_2, and ZCl_2.

Experiment 1: When metal X was placed in the YCl_2 solution, a shiny layer was observed on metal X and a color change was observed in the solution.

Experiment 2: When metal Y was placed in the ZCl_2 solution, no noticeable change occurred after a sufficient period of time.

Experiment 3: When metal Z was placed in the XCl_2 solution, a precipitate was observed and a clear change was observed in the solution.

Based on these observations, which of the following statements is true?

	Statement
a)	Metal X is the most chemically active, and its ability to lose electrons is the least among the three metals.
b)	Metal Y is the most chemically reactive, and the chemical equation for the third reaction is: $Y_{(s)} + XCl_{2(aq)} \rightarrow YCl_{2(aq)} + X_{(s)}$
c)	The order of metals from least reactive to most reactive is: Y < X < Z
d)	A simple substitution reaction can only occur if the metal in solution (in the form of an ion) is more reactive than the free metal.

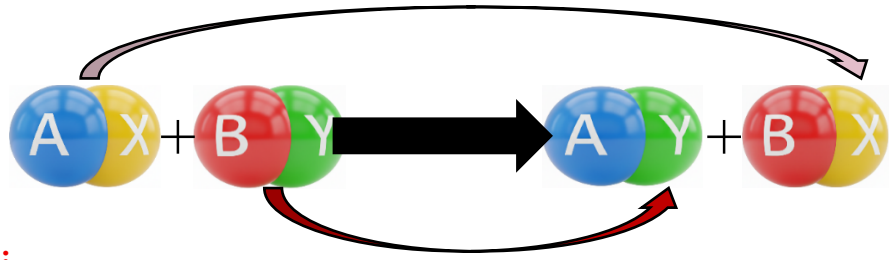
5) **Double replacement reactions:** the exchange of ions between two compounds to form two new compounds.

General form of the equation:

$$AX + BY \rightarrow AY + BX$$

Where A and B are cations (positive ions) and X and Y are anions (negative ions). In this reaction, both cations exchange places with the corresponding anions to form new compounds.

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Example 11:

$$2NaOH_{(aq)} + CuCl_{2(aq)} \rightarrow 2NaCl_{(aq)} + Cu(OH)_{2(s)}$$

Double replacement reactions are simple and orderly ion exchange processes (cations and anions), yet they are essential in many chemical processes, from the formation of precipitates in laboratories to neutralization reactions that adjust the acidity of solutions to the production of gases in various applications. The key is always to create a stable product that drives the reaction to completion. Figure 2-4 shows the steps for writing a balanced chemical equation for double displacement reactions.

Basic steps for writing balanced chemical equations for double replacement reactions:

Step		Example
1	Write the chemical formula for the reactants.	$Ba(NO_3)_2 + H_2SO_4$
2	Identify the positive and negative ions in each compound.	$Ba(NO_3)_2$ contains Ba^{2+} and NO_3^-. H_2SO_4 contains H^+ and SO_4^{2-}.
3	Match each positive ion with the negative ion in the other compound.	SO_4^{2-} pairs with Ba^{2+}. NO_3^- pairs with H^+.
4	Write the chemical formulas of the products using step 3.	$BaSO_4$ and HNO_3
5	Write the complete chemical equation for the double replacement reaction.	$Ba(NO_3)_{2(aq)} + H_2SO_{4(aq)} \rightarrow BaSO_{4(s)} + HNO_{3(aq)}$ Check solubility to confirm that the reaction occurs.
6	Balance the equation.	$Ba(NO_3)_{2(aq)} + H_2SO_{4(aq)} \rightarrow BaSO_{4(s)} + 2HNO_{3(aq)}$

Figure 2-4: Steps for writing double replacement equations

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Exercise (2-12):

Hydrochloric acid is added to two beakers, one containing sodium chloride solution and the other silver nitrate solution. The addition results in a white precipitate in one of the beakers.

1) Which of the beakers contains the precipitate?

.....

2) What is the precipitate formed?

.....

3) Write a chemical equation to show the reaction.

.....

4) Classify this reaction.

.....

Exercise (2-13):

Classify the following reactions according to the types you have studied:

	The Equation	Classification
1)	Sulfur trioxide + water → sulfuric acid	
2)	Magnesium + iron(III) chloride → magnesium chloride + iron	
3)	When oxygen fluoride gas is heated, oxygen gas and fluorine gas are produced.	
4)	Hydrogen and oxygen are obtained from water by electrolysis.	
5)	Table salt is obtained from the reaction of sodium and chlorine.	

Exercise (2-14):

Classify the following reactions according to the types you have studied:

1) $Q + XZ \rightarrow X + QZ$ (.....)

2) $Q + Z \rightarrow QZ$ (.....)

3) $RX + TW \rightarrow RW + TX$ (.....)

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Exercise (2-15):

The following table represents the results of an experiment conducted by a chemist researcher in which different metals (minerals) were added to aqueous solutions of different ions of those metals, where the positive sign (+) represents the occurrence of a reaction and the negative sign (-) represents the non-occurrence of a reaction.

Metal / solution	Al	Cd	Co	Mn	Ni	Pb
$AlCl_3$		-	-	-	-	-
$CdCl_2$	+		-	+	-	-
$CoCl_2$	+	+		+	-	-
$MnCl_2$	+	-	-		-	-
$NiCl_2$	+	+	+	+		-
$PbCl_2$	+	+	+	+	+	

Study the table above carefully and then answer the following questions:

i. Which metal is more active? .....

ii. Which one is less active?.....

iii. Can a cobalt chloride solution be stored in a container made of manganese? Why?

.....

iv. Can aluminum chloride solution be stored in a lead container? Why?

.....

v. If you were asked to make containers to store all the solutions in the table above individually, what metal would you choose from among the previous metals to make the containers?

.....

Write a chemical equation for the reaction of cobalt with nickel chloride? Indicating the type of reaction?

.....

.....

Reactions in aqueous solutions

Aqueous solution reactions are chemical reactions that occur when the reactants (solutes) are dissolved in water. In these reactions, water acts as the solvent and is the largest component of the solution.

Classification of substances in water:

Compounds dissolved in water are classified based on their ability to dissociate into ions (conductivity of electricity).

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Type	Properties	Example
Strong electrolytes	They dissociate or ionize completely in water to form a large number of ions. Most soluble ionic compounds, strong acids, and strong bases are included.	**NaCl** (table salt) **HCl** (hydrochloric acid)
Weak electrolytes	They partially ionize in water to form a small number of ions. Includes weak acids and bases.	**CH_3COOH** (acetic acid)
Non-electrolytes	They dissolve in water but do not ionize or dissociate into ions, and remain in the form of molecules.	**$C_{12}H_{22}O_{11}$** (table sugar)

Dissolution of Molecular Compounds in Aqueous Solutions:

Molecular (or covalent) compounds are substances composed of atoms held together by covalent bonds. When these compounds dissolve in an aqueous solution, they behave in one of two main ways:

1) **Dissolving without ionization (formation of non-electrolytes):**

This is the most common behavior of ordinary molecular compounds, especially those not classified as strong acids or bases. As shown in Figure 2-5.

Example 12:

Sugar (sucrose):

$$C_{12}H_{22}O_{11(s)} \xrightarrow{H_2O} C_{12}H_{22}O_{11(aq)}$$

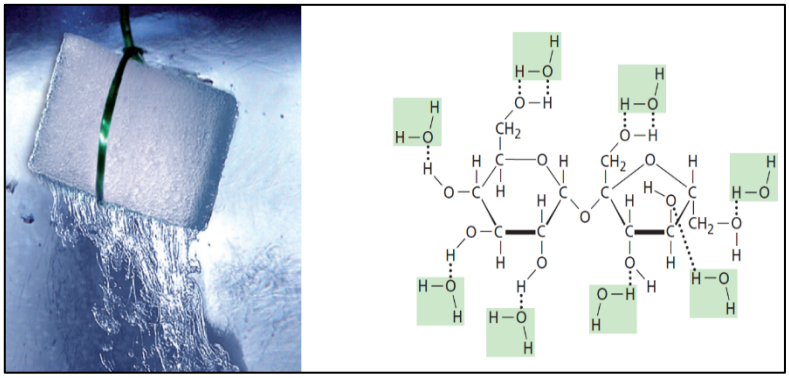


Figure 2-5: Dissolving sugar in water

2) **Dissolution with ionization (formation of electrolytes):**

In special cases, some polar molecules react with water molecules to form ions. This process is called ionization. As shown in Figure 2-6.

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Example 13:

$$HCl_{(g)} + H_2O_{(l)} \rightarrow H_3O^+_{(aq)} + Cl^-_{(aq)}$$

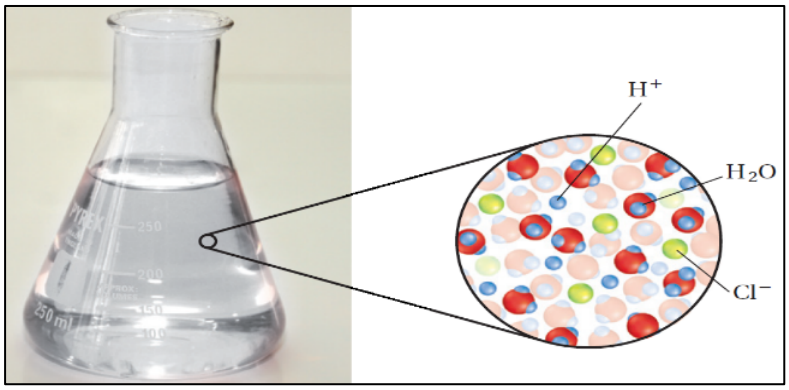


Figure 2-6: Electrolyte solubility in water

Dissolution of Ionic Compounds in Aqueous Solutions

Ionic compounds dissolve in polar solutions (such as aqueous solution) by dissociating into freely moving positive and negative ions. This process is known as dissolution or hydration when the solvent is water. Figure 2-7 shows the process of dissolving table salt in water.

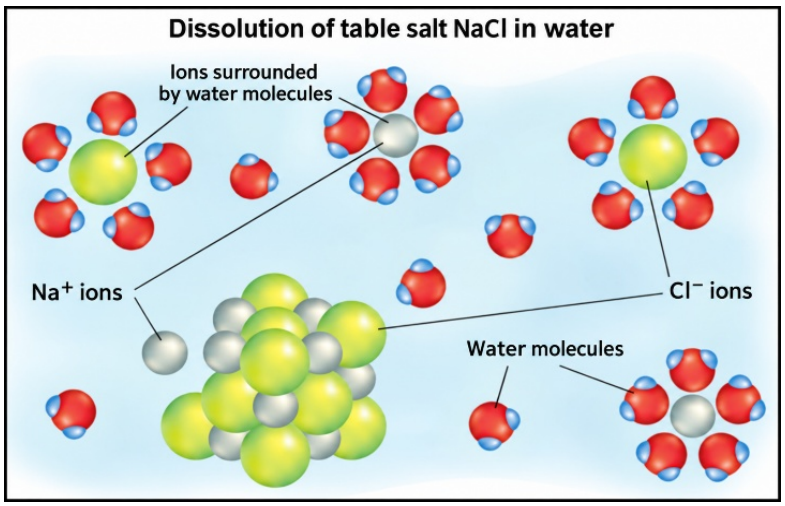


Figure 2-7: Dissolution of ionic compounds in water

Describing reactions with ionic equations:

Ionic equations are used to describe reactions in aqueous solutions more accurately, especially for strong electrolytes.

1) **Complete (overall) ionic equation:** All strong electrolytes dissolved in water are shown as dissociated ions, while the precipitate, water, or gas (or weak electrolyte) is written in its molecular formula.

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Example 14:

$$AgNO_{3(aq)} + NaCl_{(aq)} \rightarrow AgCl_{(s)} + NaNO_{3(aq)}$$

Write the complete ionic equation for this reaction as follows:

$$Ag^{+}_{(aq)} + NO_3^{-}{}_{(aq)} + Na^{+}_{(aq)} + Cl^{-}_{(aq)} \rightarrow AgCl_{(s)} + Na^{+}_{(aq)} + NO_3^{-}{}_{(aq)}$$

2) **Net (final) ionic equation:** It is obtained by eliminating spectator ions, which are ions that appear unchanged on both the reactant and product sides and do not actually participate in the chemical reaction.

Example 15:

$$Ag^{+}_{(aq)} + NO_3^{-}{}_{(aq)} + Na^{+}_{(aq)} + Cl^{-}_{(aq)} \rightarrow AgCl_{(s)} + Na^{+}_{(aq)} + NO_3^{-}{}_{(aq)}$$

To write the net ionic equation for this reaction, remove the spectator ions:

$$Ag^{+}_{(aq)} + \cancel{NO_3^{-}{}_{(aq)}} + \cancel{Na^{+}_{(aq)}} + Cl^{-}_{(aq)} \rightarrow AgCl_{(s)} + \cancel{Na^{+}_{(aq)}} + \cancel{NO_3^{-}{}_{(aq)}}$$

Final result:

$$Ag^{+}_{(aq)} + Cl^{-}_{(aq)} \rightarrow AgCl_{(s)}$$

Spectator ions: ions present on both sides of an ionic equation (reactants and products) without undergoing any change in their chemical or physical state.

Types of reactions in aqueous solutions:

A double replacement reaction occurs and is considered complete only if some of the ions are removed from the solution to form one of the following products:

1) **Precipitate formation (Precipitation reactions):**

This is the most common type. The positive and negative ions combine to form an insoluble ionic compound (solid) that separates from the solution. Figure 2-8 shows the process of precipitation of copper(II) hydroxide.

Example 16:

$$2NaOH_{(aq)} + CuCl_{2(aq)} \rightarrow 2NaCl_{(aq)} + Cu(OH)_{2(s)}$$

The complete ionic equation:

$$2Na^{+}_{(aq)} + 2OH^{-}_{(aq)} + Cu^{2+}_{(aq)} + 2Cl^{-}_{(aq)} \rightarrow 2Na^{+}_{(aq)} + 2Cl^{-}_{(aq)} + Cu(OH)_{2(s)}$$

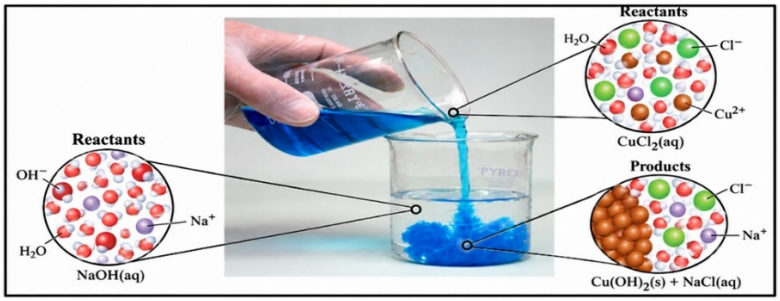
Net ionic equation:

$$2\cancel{Na^{+}_{(aq)}} + 2OH^{-}_{(aq)} + Cu^{2+}_{(aq)} + \cancel{2Cl^{-}_{(aq)}} \rightarrow \cancel{2Na^{+}_{(aq)}} + \cancel{2Cl^{-}_{(aq)}} + Cu(OH)_{2(s)}$$

$$2OH^{-}_{(aq)} + Cu^{2+}_{(aq)} \rightarrow Cu(OH)_{2(s)}$$

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Figure 2-8:
Precipitation Reactions



2) Formation of a water molecule (neutralization reactions):

It occurs between an acid and a base. The positive hydrogen ion (H^+) from the acid combines with the negative hydroxide ion (OH^-) from the base to form a neutral water molecule (H_2O). Figure 2-9 shows the process of water formation.

Example 17:

$$HBr_{(aq)} + NaOH_{(aq)} \rightarrow NaBr_{(aq)} + H_2O_{(l)}$$

The complete ionic equation:

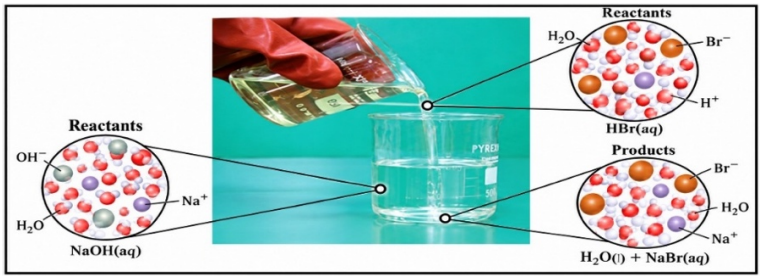
$$H^+_{(aq)} + Br^-_{(aq)} + Na^+_{(aq)} + OH^-_{(aq)} \rightarrow Na^+_{(aq)} + Br^-_{(aq)} + H_2O_{(l)}$$

Net ionic equation:

$$H^+_{(aq)} + \cancel{Br^-_{(aq)}} + \cancel{Na^+_{(aq)}} + OH^-_{(aq)} \rightarrow \cancel{Na^+_{(aq)}} + \cancel{Br^-_{(aq)}} + H_2O_{(l)}$$

$$H^+_{(aq)} + OH^-_{(aq)} \rightarrow H_2O_{(l)}$$

Figure 2-9:
Neutralization reactions



3) Gas formation:

One of the products is a gas that leaves the solution. This gas often results from the decomposition of an unstable intermediate compound, such as acetic acid with baking soda (sodium hydrogen carbonate), to produce CO_2 gas.

Example 18:

Chemical equation:

$$NaHCO_{3(aq)} + HCl_{(aq)} \rightarrow NaCl_{(aq)} + H_2O_{(l)} + CO_{2(g)}$$

The reaction takes place in two stages:

1. $NaHCO_{3(aq)} + HCl_{(aq)} \rightarrow NaCl_{(aq)} + H_2CO_{3(g)}$
2. $H_2CO_{3(g)} \rightarrow H_2O_{(l)} + CO_{2(g)}$

$$NaHCO_{3(aq)} + HCl_{(aq)} + \cancel{H_2CO_{3(g)}} \rightarrow \cancel{H_2CO_{3(g)}} + NaCl_{(aq)} + H_2O_{(l)} + CO_{2(g)}$$

The complete ionic equation:

$$Na^+_{(aq)} + HCO_3^-{}_{(aq)} + H^+_{(aq)} + Cl^-_{(aq)} \rightarrow Na^+_{(aq)} + Cl^-_{(aq)} + H_2O_{(l)} + CO_{2(g)}$$

Net ionic equation:

$$\cancel{Na^+_{(aq)}} + HCO_3^-{}_{(aq)} + H^+_{(aq)} + \cancel{Cl^-_{(aq)}} \rightarrow \cancel{Na^+_{(aq)}} + \cancel{Cl^-_{(aq)}} + H_2O_{(l)} + CO_{2(g)}$$

$$HCO_3^-{}_{(aq)} + H^+_{(aq)} \rightarrow H_2O_{(l)} + CO_{2(g}$$

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Exercise (2-16):

Write the balanced, fully ionic, and final ionic chemical equations when potassium iodide solution reacts with lead nitrate solution to form solid lead iodide.

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Exercise (2-17):

Iron(III) pyrophosphate, sold under the trade name Triphosphate, is an iron substitute medical product used for the maintenance of hemoglobin in the blood of adults. It is prepared by reacting a solution of calcium pyrophosphate ($Ca_2P_2O_7$) with a solution of ferric chloride ($FeCl_3$) to produce a solution of calcium chloride ($CaCl_2$) and a solid of iron(III) pyrophosphate.

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Exercise (2-18):

Write the balanced, fully ionic, and final ionic chemical equations for the chemical reaction that occurs when sodium carbonate and manganese chloride (**V**) solutions are mixed to form a precipitate containing manganese.

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Exercise (2-19):

When studying the product of the reaction of barium acetate with sodium sulfite, it was found that the lattice energy of the barium sulfite compound was very high. After filtering the reaction, the precipitate was mixed with a hydrochloric acid solution.

Write the complete ionic equation and the net ionic equation for the latter reaction.

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Exercise (2–20):

Four experiments were conducted in a chemistry laboratory by mixing aqueous solutions of different salts, and the observations were as follows:

Based on the principles of double replacement reactions, Which of the following statements is NOT correct?

Experiment	Added Aqueous Solutions	Observation After Mixing
1	Potassium bromide + sodium nitrate	No noticeable change occurred, and the solution remained clear.
2	Copper sulfate + sodium hydroxide	A thick blue precipitate formed.
3	Ammonium chloride + potassium hydroxide	A gas with a pungent odor was released.
4	Sulfuric acid + barium hydroxide	A white precipitate formed, and the temperature of the solution increased significantly.

	Phrase
a)	Double replacement reactions always occur unless all products are produced as soluble aqueous solutions.
b)	The reaction in experiment (2) resulted in the formation of sodium sulfate and a precipitate of copper hydroxide.
c)	The reaction in experiment (4) is a precipitation reaction only, as the heat of reaction is not related to the conditions for the occurrence of double replacement.
d)	The reaction that did not occur in experiment (1) would have occurred if we had replaced $NaNO_3$ with $AgNO_3$.

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Exercise answers

Exercise (2-1):

Bromine+ hydrogen ⟶ Hydrogen Bromide

Exercise (2-2):

$$\text{Ammonium nitrate} \xrightarrow{\Delta} \text{nitrous oxide + water}$$

Exercise (2-3):

C

Exercise (2-4):

Solid potassium chlorate ⟶Solid potassium chloride + Oxygen gas

$$2KClO_{3(s)} \rightarrow 2KCl_{(s)} + 3O_{2(g)}$$

Exercise (2-5):

calcium carbonate + hydrochloric acid ⟶ calcium chloride + water + CO_2 $CaCO_3 + 2HCl \rightarrow CaCl_2 + H_2O + CO_2$
potassium hydroxide + sulfuric acid ⟶ potassium sulfate + water $2KOH + H_2SO_4 \rightarrow K_2SO_4 + 2H_2O$
tin(Ⅱ) sulfate + potassium phosphate ⟶ tin(Ⅱ) phosphate + potassium sulfate $3SnSO_4 + 2K_3PO_4 \rightarrow Sn_3(PO_4)_2 + 3K_2SO_4$
aluminum + hydrochloric acid ⟶ aluminum chloride + hydrogen gas $2Al + 6HCl \rightarrow 2AlCl_3 + 3H_2$

Exercise (2-6):

c) $2Li_3PO_4 + 3CaCl_2 \rightarrow 6LiCl + Ca_3(PO_4)_2$

Exercise (2-7):

a)

$ZnS + 2HCl \rightarrow ZnCl_2 + H_2S$				
Element	Zn	S	H	Cl
Reactants	1	1	2	2
Product	1	1	2	2

b)

$6HCl + 2Cr \rightarrow 2CrCl_3 + 3H_2$			
Element	H	Cl	Cr
Reactants	6	6	2
Product	6	6	2

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c)

$2FeCl_3 + 3Na_2CO_3 \rightarrow Fe_2(CO_3)_3 + 6NaCl$				
Element	Fe	Cl	Na	CO_3
Reactants	2	6	6	3
Product	2	6	6	3

d)

$Na_2O + (NH_4)_2SO_4 \rightarrow Na_2SO_4 + H_2O + 2NH_3$					
Element	Na	O	N	H	S
Reactants	2	5	2	8	1
Product	2	5	2	8	1

Exercise (2-8):

1) $C_2H_4O_2 + 2O_2 \rightarrow 2CO_2 + 2H_2O$
2) $Mn(NO_2)_2 + BeCl_2 \rightarrow Be(NO_2)_2 + MnCl_2$
3) $3AgBr + GaPO_4 \rightarrow Ag_3PO_4 + GaBr_3$
4) $3H_2SO_4 + 2B(OH)_3 \rightarrow B_2(SO_4)_3 + 6H_2O$
5) $2NaNO_3 + PbO \rightarrow Pb(NO_3)_2 + Na_2O$
6) $Ca_3(PO_4)_2 + 3H_2SO_4 \rightarrow 3CaSO_4 + 2H_3PO_4$
7) $2Al + 3H_2SO_4 \rightarrow Al_2(SO_4)_3 + 3H_2$
8) $Al_4C_3 + 12H_2O \rightarrow 4Al(OH)_3 + 3CH_4$
9) $2Na + 2H_2O \rightarrow 2NaOH + H_2$

Exercise (2-9):

a) $F_{2(g)} + FeI_{2(aq)} \rightarrow FeF_{2(aq)} + I_{2(s)}$

Exercise (2-10):

1)	$Mg_{(s)} + ZnCl_{2(aq)} \rightarrow MgCl_{2(aq)} + Zn_{(s)}$
2)	$F_{2(g)} + 2HCl_{(aq)} \rightarrow Cl_{2(g)} + 2HF_{(aq)}$
3)	$Al_{(s)} + AlCl_{3(aq)} \rightarrow$ NR
4)	$Fe_{(s)} + Na_3PO_{4aq)} \rightarrow$ NR
5)	$2Al_{(s)} + 3Pb(NO_3)_{2(aq)} \rightarrow 3Pb_{(s)} + 2Al(NO_3)_{3(aq)}$

Exercise (2-11):

C

Exercise (2-12):

1) Silver nitrate beaker
2) Silver chloride
3) $AgNO_{3(aq)} + HCl_{(aq)} \rightarrow AgCl_{(s)} + HNO_{3(aq)}$
4) Double displacement reaction

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Exercise (2-13):

	The Equation	Classification
1)	Sulfur trioxide + water → sulfuric acid	Formation
2)	Magnesium + iron(III) chloride → magnesium chloride + iron	Simple substitution
3)	When oxygen fluoride gas is heated, oxygen gas and fluorine gas are produced.	Decomposition
4)	Hydrogen and oxygen are obtained from water by electrolysis.	Decomposition
5)	Table salt is obtained from the reaction of sodium and chlorine.	Formation

Exercise (2-14):

1) Q + XZ → X + QZ (Simple substitution)
2) Q + Z → QZ (Formation)
3) RX + TW → RW + TX (Duble substitution)

Exercise (2-15):

i. Al
ii. Pb
iii. No, because a reaction occurs between cobalt chloride and manganese.
iv. Yes, because no reaction occurs between aluminum chloride and lead
v. Lead
vi. $Co + Ni^{2+} \rightarrow Co^{2+} + Ni$

Reaction type: Single Displacement

Exercise (2-16):

Chemical equation:

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

The complete ionic equation:

$$2K^+_{(aq)} + 2I^-_{(aq)} + Pb^{2+}_{(aq)} + 2NO_3^-{}_{(aq)} \rightarrow 2K^+_{(aq)} + 2NO_3^-{}_{(aq)} + PbI_{2(s)}$$

Final ionic equation:

$$2I^-_{(aq)} + Pb^{2+}_{(aq)} \rightarrow PbI_{2(s)}$$

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Exercise (2-17):

Chemical equation:

$$3Ca_2P_2O_{7(aq)} + 4FeCl_{3(aq)} \rightarrow 6CaCl_{2(aq)} + Fe_4(P_2O_7)_{3(s)}$$

The complete ionic equation:

$$6Ca^{2+}_{(aq)} + 3P_2O_7^{4-}{}_{(aq)} + 4Fe^{3+}_{(aq)} + 12Cl^-_{(aq)} \rightarrow 6Ca^{2+}_{(aq)} + 12Cl^-_{(aq)} + Fe_4(P_2O_7)_{3(s)}$$

Final ionic equation:

$$3P_2O_7^{4-}{}_{(aq)} + 4Fe^{3+}_{(aq)} \rightarrow Fe_4(P_2O_7)_{3(s)}$$

Exercise (2-18):

Chemical equation:

$$5Na_2CO_{3(aq)} + 2MnCl_{5(aq)} \rightarrow 10NaCl_{(aq)} + Mn_2(CO_3)_{5(s)}$$

The complete ionic equation:

$$10Na^+_{(aq)} + 5CO_3^{2-}{}_{(aq)} + 2Mn^{5+}_{(aq)} + 10Cl^-_{(aq)} \rightarrow 10Na^+_{(aq)} + 10Cl^-_{(aq)} + Mn_2(CO_3)_{5(s)}$$

Final ionic equation:

$$5CO_3^{2-}{}_{(aq)} + 2Mn^{5+}_{(aq)} \rightarrow Mn_2(CO_3)_{5(s)}$$

Exercise (2-19):

Chemical equation:

$$BaSO_{3(s)} + 2HCl_{(aq)} \rightarrow BaCl_{2(aq)} + SO_{2(g)} + H_2O_{(l)}$$

The complete ionic equation:

$$BaSO_{3(s)} + 2H^+_{(aq)} + 2Cl^-_{(aq)} \rightarrow Ba^{2+}_{(aq)} + 2Cl^-_{(aq)} + SO_{2(g)} + H_2O_{(l)}$$

Net ionic equation:

$$BaSO_{3(s)} + 2H^+_{(aq)} + 2Cl^-_{(aq)} \rightarrow Ba^{2+}_{(aq)} + 2Cl^-_{(aq)} + SO_{2(g)} + H_2O_{(l)}$$

$$BaSO_{3(s)} + 2H^+_{(aq)} \rightarrow Ba^{2+}_{(aq)} + SO_{2(g)} + H_2O_{(l)}$$

Exercise (2-20):

C

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03 Measuring Matter

In chemistry, the term "measurement of matter" refers to the precise determination of the amount of a substance, especially when dealing with chemical reactions that require specific ratios of reactants and products. Because atoms and molecules are so small that they cannot be counted individually, chemists adopted the mole as the basic unit of measurement for "amount of matter" in the International System of Units (SI)..

Mole: The standard unit for counting particles (atoms, molecules, ions, or chemical formula units) in chemistry.

Particles	Example
Atoms	Atoms of elements (Fe, Cu, Ag, etc.)
Molecules	Elemental molecules (O_2, F_2, P_4, etc.) and compound molecules (H_2O, CO_2, etc.)
Ions	Monatomic ions (Cl^-, Cr^{2+}, etc.) and polyatomic ions (OH^-, NH_4^+, etc.)
Chemical formula units	Covalent network compounds (SiO_2) and ionic compounds (NaCl, Al_2O_3)

A mole is also defined as the amount of a substance that contains the same number of particles as 12 g of the carbon-12 isotope. This constant is called Avogadro's number (N_A).

$$N_A = 6.02x10^{23}$$

The relationship between moles and the number of particles

Avogadro's number can be used to convert between moles and the number of particles according to the following relationship:

Mathematical relationship:

$$n = \frac{N}{N_A} = \frac{\text{Number of Particles}}{\text{Avogadro Number}}$$

Example 1:

How many moles are there in $5.75x10^{24}$ atoms of aluminum?

$$n = \frac{N}{N_A}$$

$$n = \frac{5.75x10^{24}}{6.02x10^{23}}$$

n = 9.55 mol

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Exercise (3-1):

Which of the following options contains the largest number of oxygen atoms?

a) 10.0 mol O_3
b) 3.0 mol O_2
c) 16.0 mol O_2
d) 16.0 mol H_2O

Exercise (3-2):

We have an unknown sample of ozone gas (O_3) distributed among three different containers:

- The first container contains 0.75 moles of ozone gas (O_3) molecules.
- The second container contains half the number of molecules in the first container.
- The third container contains 3.011×10^{23} molecules more than the second container.

What is the total number of moles of ozone (O_3) molecules present in the three containers together?

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Mass and mole

The mass of one mole is the mass of one mole of a substance, that is, the mass of a sample containing Avogadro's number of particles.

Molar mass (M_m)

The molar mass of an element is the mass of one mole of atoms of that element (g/mol).

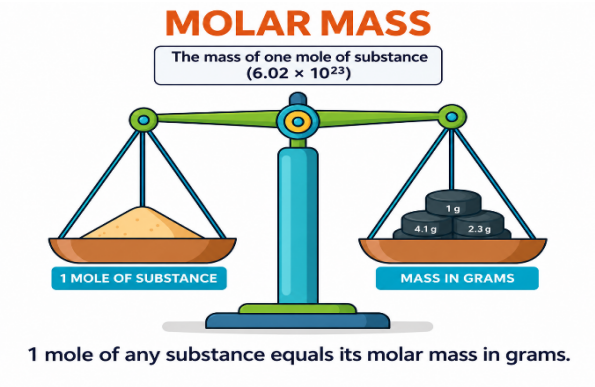


Figure 3-1: Molar mass representation

The molar mass of an element is equal to the relative atomic mass of that element (as found in the periodic table), but in g/mol rather than atomic mass units (amu).

Example 2:

molar mass O_2:

O : 2 × 16.00 g/mol = 32.00 g/mol

Molar mass (M) = 32.00 g/mol

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Exercise (3-3):

Which of the following cannot be considered one mole?

a) 65.34 g Zn
b) 26.0 g Fe
c) 6.02×10^{23} Ag atom
d) 6.02×10^{23} O_2 molecule

Mathematical relationship:

$$\text{Number of moll} = \frac{mass}{\text{molar mass}}$$

$$n = \frac{m}{M}$$

Example 3:

Calculate the mass in grams of aluminum Al which has 3.75 mol?

$m = n \times M$

$m = 3.75\ mol \times 27\ g/mol$

$m = 101.25\ g$

Exercise (3-4):

What is the molar mass of $C_8H_9NO_2$?

a) 162 g/mol
b) 76 g/mol
c) 151 g/mol
d) 70 g/mol

Exercise (3-5):

We have a sample of aluminum sulfate ($Al_2(SO_4)_3$), if the portion representing one-third of the total mass of the sample contains 0.8700 moles of sulfate groups (SO_4^{2-}).
What is the total mass of the original aluminum sulfate sample in grams?

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The relationship between the number of particles and mass

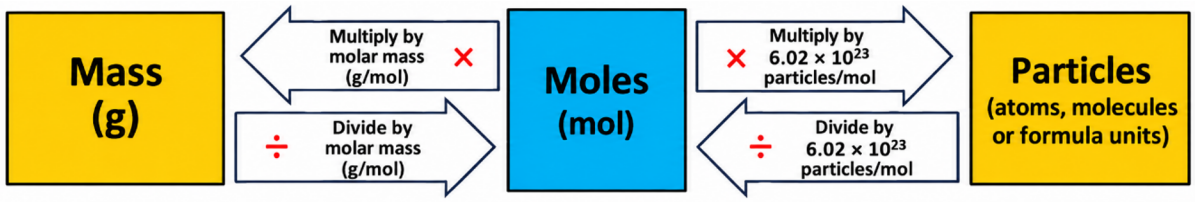


Figure 3-2: Some moles relationships

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Example 4:

How many atoms are in 11.5 g of mercury Hg?

Solution

$N = \frac{m}{M} \times 6.02\times10^{23}$

$N = \frac{11.5}{200.6} \times 6.02\times10^{23}$

$N = 3.45 \times 10^{22}$ atom

Exercise (3-6):

How many moles of ions are present in 2 mol of sodium fluoride NaF?

a) 2 mol
b) 0.5 mol
c) 4 mol
d) 1 mol

Exercise (3-7):

We have a solid mixture consisting of two pure compounds: sodium chloride (NaCl) and calcium chloride ($CaCl_2$). The total mass of the mixture is 55.7 g. When the mixture is completely dissolved in water, the resulting solution is found to contain 6.022×10^{23} chloride ions (Cl^-).

How many formula units of sodium chloride (NaCl) are in this mixture?

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Relationships between the mole and the chemical formula of compounds

The relationship between the mole and the chemical formula of compounds is based on three basic concepts.

- **Molar mass and its relationship to the chemical formula of compounds**

The chemical formula of a compound is used to determine the number of atoms of each element, and then the molar mass of the compound is calculated by adding the atomic masses of all the atoms that make it up as they appear in the formula.

Example 5:

To calculate the molar mass of water (H_2O).

H : 2 × 1.008 = 2.016 g/mol

O : 1 × 16.00 = 16.00 g/mol

M = 2.016 + 16.00 = 18.016 g/mol

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Exercise (3-8):

Calculate the molar mass of Hexaammine cobalt(III) chloride, whose chemical formula is $[Co(NH_3)_6]Cl_3$?

a) 184.35 g/mol
b) 267.52 g/mol
c) 496.93 g/mol
d) 189.44 g/mol

- **Avogadro's number and its relationship to the chemical formula of compounds**

A chemical formula represents the ratio of the number of particles (or moles) of the elements that make up a compound.

Example 6:

1 mole of methane gas (CH_4) contains 6.02×10^{23} molecules of methane. Based on the formula, this mole also contains:

1 mole of carbon (C) atoms (6.02×10^{23} atoms). Because there is only one carbon atom in the formula.

4 moles of hydrogen atoms (H) ($4 \times 6.02 \times 10^{23}$ atoms). Because there are only four hydrogen atoms in the formula.

Exercise (3-9):

Glucose ($C_6H_{12}O_6$) is an important sugar. If you have a sample of glucose with a mass of 180.16 g, calculate the total number of oxygen atoms in this sample?

a) 6.02×10^{23} oxygen atoms
b) 3.61×10^{24} oxygen atoms
c) 1.82×10^{23} oxygen atoms
d) 2.97×10^{22} oxygen atoms

- **Molar ratios and their relationship to the chemical formula of compounds**

The subscripts in a chemical formula indicate the number of moles of atoms per mole of the compound.

Example 7:

What is the molar ratio of carbon dioxide (CO_2)?

The mole ratio of carbon to oxygen is 1 mole to 2 moles per 1 mole of CO_2.

Relative percentage composition of the chemical formula

Percent composition is the percentage by mass of each element in a compound or sample. This percentage is constant for a pure compound, regardless of the sample size.

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Mathematical relationship:

$$\text{Mass percentage of an element} = \frac{\text{Mass of element in compound}}{\text{Molar mass of the compound}} \times 100$$

or

$$\text{Percentage by mass} = \frac{\text{Number of atoms of the element} \times \text{Molar mass of the element}}{\text{Molar mass of the compound}} \times 100$$

Important Note: The sum of the mass percentages of all the elements in a compound must be equal to or very close to 100%.

Example 8:

Determine the percentage composition of carbon dioxide CO_2?

M = (1×C) + (2×O)

M = (1×12.01) + (2×16.00)

M = 44.01g/mol

C% = $\frac{12.01}{44.01} \times 100$

C% = 27.29%

O% = $\frac{2\times16.00}{44.01} \times 100$

O% = 72.71%

Exercise (3-10):

An anhydrous ionic compound consists of chromium (Cr) and a sulfate group (SO_4^{2-}). If the chemical formula of the compound is $Cr_2(SO_4)_3$ (chromium(III) sulfate), what is the mass percentage of oxygen (O) in the compound?

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Exercise (3-11):

We have an unknown chemical compound with the formula $X(OH)_2$, where X is an unknown metallic element. The mass percentage of oxygen (O) in this compound is 37.20%. What is the total number of moles of oxygen (O) atoms present in this 15.0 g sample? What element is this?

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Empirical formula

It is a chemical formula that expresses the simplest whole number ratio of the atoms of the elements that make up a compound. It does not represent the actual number of atoms in a molecule. It is often derived from empirical data (observation or experiment). The empirical formula may be the same as the molecular formula of some compounds, such as water (H_2O) and carbon dioxide (CO_2). Figure 3-3 shows a simplified step in the process of finding the initial formula.

Steps to find the empirical formula

	Steps
1	If the data are mass percentages, assume you have a sample with a mass of 100 g. This conversion simplifies the calculations, as the percentage of an element becomes equal to its mass in grams.
2	Convert the masses of elements obtained (either from direct data or after conversion in step 1) to moles. Use the atomic mass of the element to convert its mass to the number of moles.
3	Determine the simplest mole ratio by dividing the number of moles of each element (which you obtained in step 2) by the smallest number of moles between them.
4	If the results in step 3 are integers, skip this step. If they contain decimals, convert them to integers by multiplying all values by the smallest integer that makes them all integers.
5	Write the empirical formula using the integers from steps 3 or 4 as subscripts for the elements in the chemical formula in order.

Figure 3-3: Steps to find the empirical formula

Example 9:

An organic compound consisting of 40.0% carbon (C), 6.7% hydrogen (H), and 53.3% oxygen (O). (Atomic masses: C = 12.0, H = 1.0, O = 16.0).

Element	Mass (Step 1)	Number of Moles (Step 2)	Mole Ratio (Step 3)	Integers (Step 4)
C	40.0 g	40.0 / 12.0 = 3.33	3.33 / 3.33 = 1.0	1
H	6.7 g	6.7 / 1.0 = 6.7	6.7 / 3.33 = 2.01	2
O	53.3 g	53.3 / 16.0 = 3.33	3.33 / 3.33 = 1.0	1

Since the integers are C = 1, H = 2, O = 1, the empirical formula is CH_2O.

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Exercise (3-12):

There are two binary compounds of mercury and oxygen. Heating each compound causes the compound to decompose, releasing oxygen gas into the atmosphere and leaving a residue of pure mercury. Heating 0.6498 g of one compound leaves a residue of 0.6018 g, and heating 0.4172 g of the other compound results in a mass loss of 0.016 g. Determine the empirical formula for each compound.

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Exercise (3-13):

An unknown organic compound consists of carbon (C), hydrogen (H), oxygen (O), and sulfur (S).

A 5.000 g sample of the compound was completely combusted, producing 6.600 g of carbon dioxide (CO_2) and 2.700 g of water (H_2O). A separate analysis showed that the compound contains 25.66% sulfur (S) by mass. Find the empirical formula for this compound.

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Molecular formula

It is the formula that shows the true composition of the compound, i.e. the number of atoms of each element in the molecule, not the simple ratio between them.

The difference between molecular formula and empirical formula

Comparison	Empirical formula	Molecular formula
Definition	Show the simplest integer ratio between atoms.	Shows the actual number of atoms in a molecule.
Example	CH_2O	$C_6H_{12}O_6$
The relationship between them	Molecular formula = (n × empirical formula)	

How to find the molecular formula:

a) Calculate the empirical formula of the compound.
b) Calculate the molar mass of the empirical formula.
c) Calculate the following ratio:

$$n = \frac{The\ actual\ molar\ mass\ of\ the\ compound}{Molar\ mass\ of\ the\ empirical\ formula}$$

d) Multiply the empirical formula symbols by n to get the molecular formula.

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Example 10:

A compound has the empirical formula CH_2 and a molar mass of 56.12 g/mol. What is its molecular formula?

a) Molar mass of the empirical formula CH_2:

$$(1\times12.01) + (2\times1.01) = 14.03 \text{ g/mol}$$

b) Number of repetition units n:

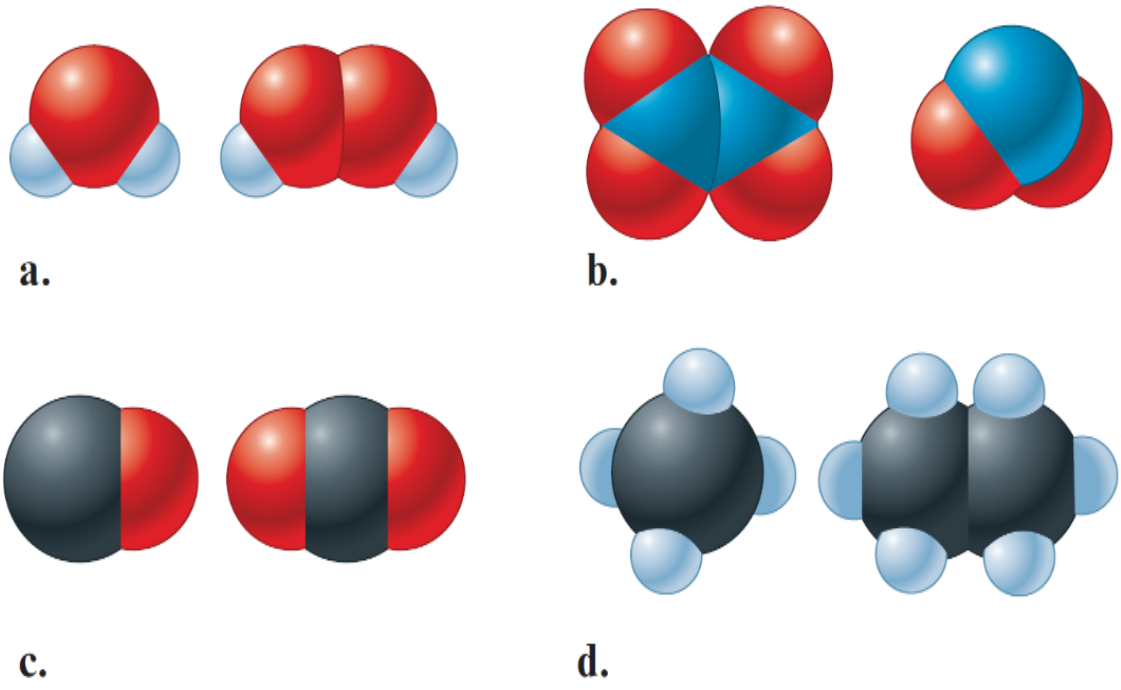
$$n = \frac{56.12}{14.03} \approx 4$$

c) Molecular formula:

$$(CH_2)\times4 = C_4H_8$$

Exercise (3-14):

Which of the following compounds have the same empirical formulas?



Exercise (3-15):

Cumene is a compound containing only carbon and hydrogen that is used in the production of acetone and phenol in the chemical industry. Combustion of 47.6 mg cumene produces some CO_2 and 42.8 mg water. The molar mass of cumene is between 115 and 125 g/mol. Determine the empirical and molecular formulas.

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Exercise (3-16):

Vitamin C (molar mass 176.12 g/mol) is a compound of C, H, and O, found in many natural sources, especially citrus fruits. When a 1.000 g sample of vitamin C is burned in a combustion device, we obtain the following data:

Mass of carbon dioxide absorbed after combustion: 85.35 g

Mass of carbon dioxide absorbed before combustion: 83.85 g

Mass of water absorbed after combustion: 37.96 g

Mass of water absorbed before combustion: 37.55 g

What is the molecular formula for vitamin C?

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Formulas of Hydrates

A hydrated salt is a solid ionic compound that contains a specific, definite number of water molecules trapped within its crystal lattice. These water molecules are called "water of crystallization."

An anhydrous salt is the same salt but with all of the water of crystallization molecules removed from it (usually by heating).

The presence of water of crystallization gives the hydrated salt different properties from the anhydrous salt, including color (e.g., copper sulfate changes from aqueous blue to anhydrous white as shown in Figure 3-4).

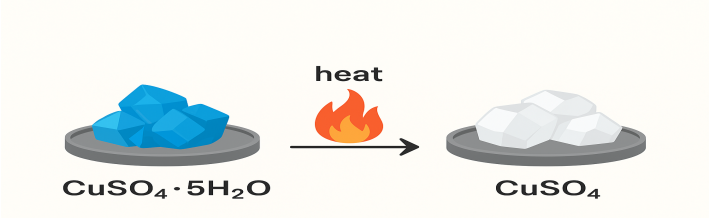


Figure 3-4: Copper sulfate changes from aqueous blue to anhydrous white

The formula for a hydrated salt is written by combining the formula for the anhydrous salt with water molecules, separated by a dot (·).

General formula:

$$\text{anhydrous salt} \cdot x H_2O$$

where:

Hydrated salt: The chemical formula of an ionic compound (salt).

x: is an integer (molar) that represents the number of moles of water molecules associated with each mole of salt.

Example 11:

$CuSO_4 \cdot 5H_2O$ Copper(II) Sulfate Pentahydrate

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Naming of hydrates

1. Naming an ionic (anhydrous) salt is the usual way (name the cation then name the anion).
2. Add a syllable to indicate the number of water molecules associated with it using Greek prefixes.
3. Ending the name with the word "water" or "hydrate".

Prefix	Number of Water Molecules	Formula	Name
Monohydrate	1	$(NH_4)_2C_2O_4 \cdot H_2O$	Ammonium oxalate monohydrate
Dihydrate	2	$CaCl_2 \cdot 2H_2O$	Calcium chloride dihydrate
Trihydrate	3	$NaC_2H_3O_2 \cdot 3H_2O$	Sodium acetate trihydrate
Tetrahydrate	4	$FePO_4 \cdot 4H_2O$	Iron(III) phosphate tetrahydrate
Pentahydrate	5	$CuSO_4 \cdot 5H_2O$	Copper(II) sulfate pentahydrate
Hexahydrate	6	$CoCl_2 \cdot 6H_2O$	Cobalt(II) chloride hexahydrate
Heptahydrate	7	$MgSO_4 \cdot 7H_2O$	Magnesium sulfate heptahydrate
Octahydrate	8	$Ba(OH)_2 \cdot 8H_2O$	Barium hydroxide octahydrate
Decahydrate	10	$Na_2CO_3 \cdot 10H_2O$	Sodium carbonate decahydrate

Figure 3-5: Examples of hydrated salts

How to determine the formula of hydrated salt:

1) The mass of the original sample of hydrated salt is measured.
2) Heating: Heating the hydrated salt to remove the water of crystallization (evaporate it) and obtain the anhydrous salt.
3) The mass of anhydrous salt is measured after all the water has been removed.
4) Calculating the mass of water: Subtract the mass of the anhydrous salt from the mass of the hydrated salt to determine the mass of the lost water of crystallization.
5) Determine the molar ratio (x):
    - Convert the mass of water lost and the mass of anhydrous salt remaining into moles.
    - Moles of water are divided by moles of anhydrous salt:

$$X = \frac{\text{number of moles of water}}{\text{Number of moles of anhydrous salt}}$$

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Example 12:

A 2.50 g sample of blue hydrated copper sulfate ($CuSO_4.xH_2O$) was placed in a beaker and heated. 1.59 g of white anhydrous copper sulfate ($CuSO_4$) remained after heating. What is the formula of the hydrated salt? What is its name?

m = 2.50 g - 1.59 g = 0.91 g

$$n_{(CuSO4)} = \frac{m}{M}$$

$$n_{(CuSO4)} = \frac{1.59}{159.6}$$

$n_{(CuSO4)}$ = 0.00996 mol

$$n_{(H2O)} = \frac{m}{M}$$

$$n_{(H2O)} = \frac{0.91}{18.02}$$

$n_{(H2O)}$ = 0.05 mol

$$X = \frac{\text{number of moles of water}}{\text{Number of moles of anhydrous salt}}$$

$$X = \frac{0.05}{0.00996} \approx 5$$

Copper(II) Sulfate Pentahydrate $CuSO_4.5H_2O$

Exercise (3–17):

Magnesium sulfate hydrate ($MgSO_4.xH_2O$) is used as a laxative. A sample of this hydrate was analyzed and found to be 51.22% water (H_2O) by mass. If the mass percentage of magnesium (Mg) in this hydrate is 9.93%, what is the value of the factor (x) (the number of water of crystallization molecules) in the chemical formula?

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Exercise (3–18):

An unknown hydrated salt with the formula $Na_2CO_3.xH_2O$ (hydrated sodium carbonate) contains nitrogen (N) as an impurity in the anhydrous sodium carbonate at a rate of 2.80% by mass of the anhydrous salt (Na_2CO_3). When 5.00 g of this hydrated salt was heated until all the water of crystallization had disappeared, the mass of the remaining material was 3.00 g. Find the value of x?

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Exercise answers

Exercise (3-1):

c) 16.0 mol O_2

Exercise (3-2):

$N_1 = 0.75\text{mol} \times 6.022\times10^{23}$ molecules/mol

$N_1 = 4.5165\times10^{23}$ molecules

$N_2 = \frac{N_1}{2}$

$N_2 = \frac{4.5165\times10^{23}}{2}$

$N_2 = 2.25825\times10^{23}$ molecules

$n_2 = \frac{N_2}{N_A}$

$n_2 = \frac{2.25825\times10^{23}}{6.022\times10^{23}} = 0.375\text{mol}$

$\Delta n = \frac{3.011\times10^{23}}{6.022\times10^{23}} = 0.50\text{mol}$

$n_3 = n_2 + \Delta n$

$n_3 = 0.375 + 0.50 = 0.875\text{mol}$

$n_t = n_1 + n_2 + n_3$

$n_t = 0.75 + 0.375 + 0.875 = 2.0$ mol

Exercise (3-3):

b) 26.0 g Fe

Exercise (3-4):

c) 151g/mol

Exercise (3-5):

$M = (2 \times M_{Al}) + (3 \times M_S) + (3 \times 4 \times M_O)$

$M = (2 \times 27.0) + (3 \times 32.1) + (3 \times 4 \times 16.0) = 342.3$ g/mol

$n_{(part)} = n(SO_4) \times \frac{1\text{ mol } Al_2(SO_4)_3}{3\text{ mol } SO_4}$

$n_{(part)} = 0.87\text{ mol} \times \frac{1}{3} = 0.29\text{ mol}$

$m_{(part)} = n_{(part)} \times M$

$m_{(part)} = 0.29\text{ mol} \times 342.3\text{ g/mol} = 99.267\text{ g}$

$m_{(t)} = 3 \times m_{(part)} = 3 \times 99.267 = 297.8\text{ g}$

Exercise (3-6):

c) 4 mol

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Exercise (3-7):

$n_{Cl^-} = \frac{number\ of\ Cl^-\ ions}{N_A}$

$n_{Cl^-} = \frac{6.022\times10^{23}\text{ion}}{6.022\times10^{23}\text{ion/mol}} = 1.00\text{ mol}$

$M_{NaCl} = 23.0 + 35.5 = 58.5\text{ g/mol}$

$M_{CaCl2} = 40.1 + (2\times35.5) = 111.1\text{ g/mol}$

$m_{Total} = n_1M_{NaCl} + n_2M_{CaCl2}$

$m_{Total} = 58.5n_1 + 111.1n_2$

$55.7 = 58.5n_1 + 111.1n_2$

$n_1 + 2n_2 = 1.00\text{ mol}$

$n_1 = 1.00 - 2n_2$

$55.7 = 58.5(1.00 - 2n_2) + 111.1n_2$

$55.7 = 58.5 - 117.0n_2 + 111.1n_2$

$55.7 - 58.5 = -5.9n_2$

$-2.8 = -5.9n_2$

$n_2 = \frac{2.8}{5.9} = 0.475\text{ mol}$

$n_1 = 1.00 - 2(0.475)$

$n_1 = 0.0508\text{ mol}$

$N_{NaCl} = n_1 \times N_A$

$N_{NaCl} = 0.0508 \times 6.022 \times10^{23} = 3.06\times10^{22}$ Formula unit

Exercise (3-8):

b) 267.52 g/mol

Exercise (3-9):

b) 3.61×10^{24} O atom

Exercise (3-10):

$M = (2 \times Cr) + (3 \times S) + (12 \times O)$

$M = (2 \times 52.0) + (3 \times 32.1) + (12 \times 1\ 6.0)$

$M = 104 + 96.3 + 192$

$M = 392.3\text{ g/mol}$

$(O)\ M = 12 \times 16.0 = 192.0\text{ g/mol}$

$O\% = \frac{\text{Mass of O in one mole of the compound}}{\text{Molar mass of the compound}} \times 100$

$O\% = \frac{192.0}{392.3} \times 100$

$O\% = 48.94\%$

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Exercise (3–11):

$O\% = \frac{\text{Mass of O in one mole of the compound}}{\text{Molar mass of the compound}} \times 100$

$m_O = 2 \times 16.00 = 32.00\ g$

$37.20\% = \frac{32.00}{M_{X(OH)_2}} \times 100$

$M_{(X(OH)2)} = \frac{32.00}{0.3720} \times 100$

$M_{(X(OH)2)} = 86.02\ g/mol$

$N_{compound} = \frac{m_{sample}}{M_{compound}} = \frac{15.0}{86.02} = 0.1744\ mol$

$n_O = C_{compound} \times \frac{2\ mol\ O}{1\ mol\ compound}$

$n_O = 0.1744\ mol \times 2 = 0.349\ mol$

$M_X = M_{X(OH)2} - M_{O2} - M_{H2}$

$M_X = 86.02 - 32.00 - (2 \times 1.00) = 52.02\ g/mol$

This mass is close to the atomic mass of the element chromium **Cr**.

Exercise (3–12):

Compound I: Mass O = 0.6498 g Hg_xO_y - 0.6018 g Hg = 0.0480 g O

$0.6018\ g\ Hg \times \frac{1\ mol}{200.6\ g} = 3.000 \times 10^{-3}\ mol\ Hg$

$0.0480\ g\ O \times \frac{1\ mol}{16\ g} = 3.00 \times 10^{-3}\ mol\ O$

The mole ratio between Hg and O is 1 : 1, so the empirical formula of compound I is HgO.

Compound II: Mass Hg = 0.4172 g Hg_xO_y - 0.016 g O = 0.4012 g Hg

$0.4012g\ Hg \times \frac{1\ mol}{200.6\ g} = 2.00 \times 10^{-3}\ mol\ Hg$; $0.016g\ O \times \frac{1\ mol}{16\ g} = 1.0 \times 10^{-3}\ mol\ O$

The mole ratio between Hg and O is 2 : 1, so the empirical formula is Hg_2O

Exercise (3–13):

$m_C = 6.600\ g\ CO_2 \times \frac{12.01\ g\ C}{44.01\ g\ CO_2} = 1.800\ g$

$m_H = 2.700\ g \times \frac{2 \times 1.01\ g\ H}{18.02\ g\ H_2O} = 0.303\ g$

m_S = total mass of sample $\times \frac{solfat\ \%}{100\%}$

$m_S = 5.000 \times 0.2566 = 1.283\ g$

m_O = total mass – ($m_C + m_H + m_S$)

$m_O = 5.000\ g - (1.800\ g + 0.303\ g + 1.283\ g) = 1.614\ g$

Element	Mass (Step 1)	Number of Moles (Step 2)	Mole Ratio (Step 3)	Integers (Step 4)
C	1.800 g	1.800/12.0 = 0.1499	0.1499/0.0400 = 3.747	3.747×4 = 15
H	0.303 g	0.303/1.0 = 0.3030	0.3030/0.0400 = 7.500	7.500×4 = 30
S	1.283 g	1.283/32.07 = 0.0400	0.0400/0.0400 = 1	1×4 = 4
O	1.614 g	1.614/16 = 0.1009	0.1009/0.0400 = 2.522	2.522×4 = 10

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The empirical formula of this complex compound is:

$C_{15}H_{30}S_4O_{10}$

Exercise (3-14):

b.

Exercise (3-15):

$42.8\ mg\ H_2O \times \frac{1\ g}{1000\ mg} \times \frac{2.016\ gH}{18.02\ gH_2O} \times \frac{1000\ mg}{1\ g} = 4.79\ mg\ H$

$Mass\ \%\ H = \frac{4.79\ mg\ H}{47.6\ mg\ cumene} = \times 100 = 10.1\%\ H$

mass % C = 100.0 - 10.1 = 89.9% C

Suppose we have 100 grams of the compound:

$89.9\ g\ C \times \frac{1\ mol\ C}{12.01\ g\ C} = 7.49\ mol\ C$

$10.1\ g\ H \times \frac{1\ mol\ H}{1.008\ g\ H} = 10.0\ mol\ H$

$\frac{10.0}{7.49} = 1.34 \approx \frac{4}{3}$

the mole H to C ratio is 4 : 3. The empirical formula is C_3H_4.

Empirical formula mass ≈ 3(12) + 4(1) = 40 g/mol.

The molecular formula must be $(C_3H_4)_3$ or C_9H_{12} because the molar mass of this formula will be between 115 and 125 g/mol (molar mass ≈ 3 × 40 g/mol = 120 g/mol).

Exercise (3-16):

$m_{CO2} = m_{(CO2)(after)} - m_{(CO2)(before)}$

$m_{CO2} = 85.35\ g - 83.85\ g = 1.50\ g\ CO_2$

$m_C = 1.50\ g\ CO_2 \times \frac{12.01g\ C}{44.01g\ CO_2} = 0.409\ g\ C$

$m_{H2O} = m_{(H2O)(after)} - m_{(H2O)(before)}$

$m_{H2O} = 37.96\ g - 37.55\ g = 0.41\ g\ H_2O$

$m_H = 0.41\ g\ H_2O \times \frac{2 \times 1.01\ g\ H}{18.02\ g\ H_2O} = 0.046\ g\ H$

$m_O = m_{total} - (m_C + m_H)$

$m_O = 1.000\ g - (0.409\ g + 0.046\ g) = 0.545\ g\ O$

Element	Mass (Step 1)	Number of Moles (Step 2)	Mole Ratio (Step 3)	Integers (Step 4)
C	0.409 g	0.409/12.0 = 0.034	0.034/0.034 = 1.0	1.0×3 = 3
H	0.046 g	0.046/1.0 = 0.046	0.046/0.034 = 1.35	1.35×3 = 4
O	0.545 g	0.545/16 = 0.034	0.034/0.034 = 1.0	1.0×3 = 3

The empirical formula of this complex compound is:

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$C_3H_4O_3$

$$n = \frac{\text{Molar mass of the empirical formula}}{\text{Actual molar mass of the compound}} = \frac{176.12}{88.06} = 2$$

Molecular formula:

$$(C_3H_4O_3)\times 2 = C_6H_8O_6$$

Exercise (3–17):

$\%(MgSO_4) = 100 - (H_2O)\%$

$\%(MgSO_4) = 100 - 51.22\%$

$\%(MgSO_4) = 48.78\%$

$$\%Mg = \frac{M_{Mg}}{M_{Total}} \times 100$$

$$9.93 = \frac{24.3}{M_{Total}} \times 100$$

$M_{Total} = 244.7$ g/mol

$M_{MgSO4} = 24.3 + 32.1 + (4\times16.0) = 24.3 + 32.1 + 64.0 = 120.4$ g/mol

$m_{H2O} = M_{Total} - M_{MgSO4}$

$m_{H2O} = 244.7 - 120.4$

$m_{H2O} = 124.3$ g/mol

$$x = \frac{m_{H2O}}{M_{H2O}} = \frac{124.3}{18}$$

$x = 6.905 \approx 7$

$MgSO_4.7H_2O$

Exercise (3–18):

Mass of water lost = Mass of hydrated salt - Mass of anhydrous salt

Mass of water lost = 5.00 g - 2.00 g = 3.00 g

$$\text{Mass of impurities} = 3.00 \times \times \frac{2.80}{100} = 0.084 \text{ g}$$

$m_{Na2CO3} = 3.00 - 0.084$ g $= 2.916$ g

$M_{Na2CO3} = (2\times23.0) + 12.0 + (3\times16.0) = 106.0$ g/mol

$M_{H2O} = (2 \times 1.0) + 16.0 = 18.0$ g/mol

$$n_{Na2CO3} = \frac{2.916\ g}{106.0\ g/mol} = 0.0275 \text{ mol}$$

$$n_{H2O} = \frac{2.00\ g}{18.0\ g/mol} = 0.1111 \text{ mol}$$

$$X = \frac{n_{H_2O}}{n_{Na_2CO_3}} = \frac{0.1111}{0.0275} = 4.0$$

Sodium carbonate tetrahydrate

$Na_2CO_3.4H_2O$

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PERIODIC TABLE

	1.	2.	3.	4.	5.	6.	7.	8.	9.	10.	11.	12.	13.	14.	15.	16.	17.	18.
	I.A																	VIII. A
1.	2.1 1 **H** 1.0	II. A											III. A	IV. A	V. A	VI. A	VII. A	2 **He** 4.0
2.	1.0 3 **Li** 6.9	1.5 4 **Be** 9.0											2.0 5 **B** 10.8	2.5 6 **C** 12.0	3.0 7 **N** 14.0	3.5 8 **O** 16.0	4.0 9 **F** 19.0	10 **Ne** 20.2
3.	0.9 11 **Na** 23.0	1.2 12 **Mg** 24.3	III. B	IV. B	V. B	VI. B	VII. B	VIII. B			I. B	II. B	1.5 13 **Al** 27.0	1.8 14 **Si** 28.1	2.1 15 **P** 31.0	2.5 16 **S** 32.0	3.0 17 **Cl** 35.5	18 **Ar** 40.0
4.	0.8 19 **K** 39.1	1.0 20 **Ca** 40.1	1.3 21 **Sc** 45.0	1.5 22 **Ti** 47.9	1.6 23 **V** 50.9	1.6 24 **Cr** 52.0	1.5 25 **Mn** 54.9	1.8 26 **Fe** 55.8	1.8 27 **Co** 58.9	1.8 28 **Ni** 58.7	1.9 29 **Cu** 63.5	1.6 30 **Zn** 65.4	1.6 31 **Ga** 69.7	1.8 32 **Ge** 72.6	2.0 33 **As** 74.9	2.4 34 **Se** 79.0	2.8 35 **Br** 79.9	36 **Kr** 83.8
5.	0.8 37 **Rb** 85.5	1.0 38 **Sr** 87.6	1.3 39 **Y** 88.9	1.4 40 **Zr** 91.2	1.6 41 **Nb** 92.9	1.8 42 **Mo** 95.9	1.9 43 **Tc** [98]	2.2 44 **Ru** 101.1	2.2 45 **Rh** 102.9	2.2 46 **Pd** 106.4	1.9 47 **Ag** 107.9	1.7 48 **Cd** 112.4	1.7 49 **In** 114.8	1.8 50 **Sn** 118.7	1.9 51 **Sb** 121.8	2.1 52 **Te** 127.6	2.5 53 **I** 126.9	54 **Xe** 131.3
6.	0.7 55 **Cs** 132.9	0.9 56 **Ba** 137.3	1.1 57 **La** 138.9	1.3 72 **Hf** 178.5	1.5 73 **Ta** 180.9	1.7 74 **W** 183.8	1.9 75 **Re** 186.2	2.2 76 **Os** 190.2	2.2 77 **Ir** 192.2	2.2 78 **Pt** 195.1	2.4 79 **Au** 197.0	1.9 80 **Hg** 200.6	1.8 81 **Tl** 204.4	1.8 82 **Pb** 207.2	1.9 83 **Bi** 209.0	2.0 84 **Po** [209]	2.2 85 **At** [210]	86 **Rn** [222]
7.	0.7 87 **Fr** [223]	0.9 88 **Ra** [226]	1.1 89 **Ac** [227]	104 **Rf** [261]	105 **Db** [262]	106 **Sg** [266]	107 **Bh** [264]	108 **Hs** [269]	109 **Mt** [268]	110 **Ds** [271]	111 **Rg** [272]	112 **Cn** [285]	113 **Nh** [284]	114 **Fl** [289]	115 **Mc** [289]	116 **Lv** [293]	117 **Ts** [294]	118 **Og** [294]

EN*
atomic number
Chem. symbol
relative atomic mass

1.1 58 **Ce** 140.1	1.1 59 **Pr** 140.9	1.1 60 **Nd** 144.2	1.1 61 **Pm** [145]	1.2 62 **Sm** 150.4	1.2 63 **Eu** 152.0	1.2 64 **Gd** 157.3	1.2 65 **Tb** 158.9	1.2 66 **Dy** 162.5	1.2 67 **Ho** 164.9	1.2 68 **Er** 167.3	1.2 69 **Tm** 168.9	1.1 70 **Yb** 173.0	1.2 71 **Lu** 175.0
1.3 90 **Th** 232.0	1.5 91 **Pa** 231.0	1.4 92 **U** 238.0	1.3 93 **Np** [237]	1.3 94 **Pu** [244]	1.3 95 **Am** [243]	1.3 96 **Cm** [247]	1.3 97 **Bk** [247]	1.3 98 **Cf** [251]	1.3 99 **Es** [252]	1.3 100 **Fm** [257]	1.3 101 **Md** [258]	1.3 102 **No** [259]	1.3 103 **Lr** [262]

* EN: electronegativity

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المصادر والمراجع العلمية

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