

MOLE CONCEPT AND STOICHIOMETRY

MOLE CONCEPT

Introduction

Mole concept deals with determining or counting the number of particles. Since the number of particles is in large quantities, it becomes very difficult to deal with large numbers; therefore, these particles are placed in groups called **moles**.

One mole of a substance has 6.02×10^{23} particles. The particles of substances grouped into moles can be molecules, atoms, ions, electrons, radicals, protons or any other specified particles.

For example, 1 mole of magnesium atoms contains 6.02×10^{23} atoms, 1 mole of magnesium ions contains 6.02×10^{23} ions, 1 mole of $\text{H}_2\text{SO}_4(\text{aq})$ molecules contains 6.02×10^{23} molecules of $\text{H}_2\text{SO}_4(\text{aq})$.

A **mole** is the amount of substance which contains 6.02×10^{23} particles. A mole can as well be defined as the amount of substance that contains the same number of particles as the number of particles in 12grams of carbon-12.

The number 6.02×10^{23} is called Avogadro 's number or constant and it is denoted by letter L.

Molar Mass

Molar mass is the mass of one mole of a substance.

It is equal to the relative atomic mass expressed in grams.

The relative atomic masses of all elements have already been established.

Relative atomic mass of an element is the average mass of isotopes of the element compared to one twelfth the mass of one atom of carbon-12.

Isotopes and Isotopy

Isotopes

Are atoms of the same element with the same number of protons but different number of neutrons.

Isotopes therefore have different mass numbers.

For elements that show isotopy, the most abundant (common) isotope is taken to be the representative of all the element.

The abundance is usually given as a percentage.

Isotopy

This is the existence of atoms of the same element with the same number of protons but different number of neutrons.

Examples of isotopes

Element	Symbol	Atomic number	Isotopes	Abundance
Hydrogen	H	1	${}^1_1\text{H}$	9.99%
			${}^2_1\text{H}$	0.01%
			${}^3_1\text{H}$	Rare
Carbon	C	6	${}^{12}_6\text{C}$	98.9%
			${}^{13}_6\text{C}$	1.1%
			${}^{14}_6\text{C}$	Trace
Chlorine	Cl	17	${}^{35}_{17}\text{Cl}$	75%
			${}^{37}_{17}\text{Cl}$	25%

The atomic mass of substances or elements are obtained from the periodic table and the abundance is always given as a percentage.

The mass of carbon atom is taken to be 12 atomic mass units (amu). This was recommended by the IUPAC (International Union of Pure and Applied Chemistry) and one twelfth of the mass of carbon atom is one.

Relative Atomic Mass (RAM) / A_r

This is the mass of one atom of an element compared to $\frac{1}{12}$ (a twelfth) of the mass of one atom of Carbon-12 isotope.

Note;

R.A.M has no units because it is a ratio of the same quantity.

$$\text{R.A.M} = \frac{\text{mass of an atom of an element}}{\left(\frac{1}{12}\right) \text{ of the mass of Carbon-12}}$$

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Relative Atomic Mass of an element that shows isotopy is dependent on the relative abundance of each isotope and the atomic masses of the isotopes.

Example

Neon has three isotopes, $^{20}_{10}\text{Ne}$ (percentage abundance 90.5%); $^{21}_{10}\text{Ne}$ (percentage abundance of 0.3%) and $^{22}_{10}\text{Ne}$ (percentage abundance of 9.2%). Calculate the relative atomic mass of neon.

Solution

$$\begin{aligned} \text{R.A.M} &= \left(\frac{90.5 \times 20}{100} \right) + \left(\frac{0.3 \times 21}{100} \right) + \left(\frac{9.2 \times 22}{100} \right) \\ &= \underline{20.187} \end{aligned}$$

Examples of some R.A.M of elements are given below;

Atoms	Relative atomic mass	Molar mass
Hydrogen	1	1g
Carbon	12	12g
Oxygen	16	16g
Sodium	23	23g
Magnesium	24	24g
Sulphur	32	32g
Aluminium	27	27g
Copper	64	64g
Silver	108	108g
Lead	207	207g
Chlorine	35.5	35.5g
Calcium	40	40g
Potassium	39	39g
Tin	119	119g
Zinc	65	65g
Nickel	59	59g
Nitrogen	14	14g
Iron	56	56g
Phosphorus	31	31g

Relative formula mass/RFM (Mr) or molecular mass/RMM (Mr)

Generally, these two can be defined as the mass of one mole of a compound. They are obtained by adding the relative atomic masses of the atoms present in a compound/molecule. They both have no units.

RMM is reserved for covalent compounds or molecules.

RFM is reserved for ionic compounds. This is because they are composed of aggregates of ions arranged in crystal lattice.

Important terms

Lattice means a structure consisting of strips crossed and fastened together but with spaces left between, just the way you use some wire fences.

A molecule is a structure of atoms connected by covalent bonds. Not all chemical compounds are a molecule.

A particle is just a small portion of matter.

Examples

Calculate the formula/molecular masses of the following compounds.

a) Water, H₂O (H=1, O=16)

$$= (1 \times 2) + 16$$

$$= \underline{18}$$

b) Oxygen molecule, O₂ (O=16)

$$= 2 \times 16$$

$$= \underline{32}$$

c) Sodium sulphate, Na₂SO₄ (Na=23, S=32, O=16)

$$= (2 \times 23) + 32 + (4 \times 16)$$

$$= 46 + 32 + 64$$

$$= \underline{142}$$

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d) Copper (II) sulphate crystals, $\text{CuSO}_4 \cdot 5\text{H}_2\text{O}$ (Cu=64, S=32, O=16, H=1)

$$= 64 + 32 + (4 \times 16) + 5(2 \times 1 + 16)$$

$$= \underline{250}$$

Exercise

Calculate the relative formula/molecular masses of the following compounds.

a) $\text{FeSO}_4 \cdot 7\text{H}_2\text{O}$, (Fe=56, S=32, O=16, H=1)

b) $(\text{NH}_4)_2\text{SO}_4$, (N=14, H=1, S=32, O=16)

c) Al_2O_3 (Al=27, O=16)

d) $\text{Ca}(\text{HCO}_3)_2$ (Ca=40, H=1, C=12, O=16)

e) Magnesium hydroxide (H=1., O=16, Mg=24)

f) Sodium carbonate decahydrate

g) Sodium hydroxide

h) Calcium hydroxide

i) Hydrated copper (II) sulphate/copper (II) sulphate pentahydrate

j) Chalk

k) Sand

l) Ammonia

m) Carbon dioxide (C=12, N=14, O=16, Na=23, S=32, Ca=40, Cu=64, Pb=207)

Calculating number of particles

This is based on the relationship that one mole of a substance contains 6.02×10^{23} particles.

Examples

1. Calculate the number of particles in the following compounds.

a) Water (H_2O)

Number of molecules in 1 mole of $\text{H}_2\text{O} = 6.02 \times 10^{23}$ molecules

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1 mole of H₂O contains 2 moles of H atoms, therefore, number of H atoms in 1 mole of H₂O

$$\begin{aligned} &= (2 \times 6.02 \times 10^{23}) \\ &= \underline{12.04 \times 10^{23} \text{ atoms.}} \end{aligned}$$

Number of O atoms in 1 mole of H₂O = 6.02×10^{23} .

b) 2 moles of oxygen molecules, (2O₂)

1 mole of oxygen molecules contains 6.02×10^{23} molecules

2 moles of oxygen molecules contain $(2 \times 6.02 \times 10^{23}) = 12.04 \times 10^{23}$ molecules

1 mole of oxygen molecule contains 2 moles of oxygen atoms

2 moles of oxygen molecules contain (2x2) moles of oxygen atoms

1 mole of oxygen atom contains 6.02×10^{23} atoms

4 moles of oxygen atoms contain $(4 \times 6.02 \times 10^{23})$ atoms

$$= \underline{24.08 \times 10^{23} \text{ atoms.}}$$

c) Number of hydrogen ions in 2 moles of H₂SO₄

1 mole of H₂SO₄ contains 2 moles of H⁺ ions

2 moles of H₂SO₄ contains (2x2) moles of H⁺ ions

1 mole of H⁺ contains 6.02×10^{23} ions

4 moles of H⁺ contains $(4 \times 6.02 \times 10^{23})$ ions

$$= \underline{24.08 \times 10^{23} \text{ H}^+ \text{ ions}}$$

d) Total number of ions in 1 mole of (NH₄)₂SO₄



1 mole of (NH₄)₂SO₄ contains 3 moles of ions

1 mole of ion contains 6.02×10^{23} ions

3 moles of ions contain $(3 \times 6.02 \times 10^{23})$ ions

$$= \underline{18.06 \times 10^{23} \text{ ions}}$$

2. Calculate the number of atoms in the following

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a) 0.25 moles of calcium

1 mole of calcium contains 6.02×10^{23} atoms

$$\begin{aligned} 0.25 \text{ moles of calcium contain } & \frac{0.25 \times 6.02 \times 10^{23}}{1} \text{ atoms} \\ & = \underline{1.5 \times 10^{23} \text{ atoms}} \end{aligned}$$

b) 8 moles of sulphur

1 mole of sulphur contains 6.02×10^{23} atoms

$$\begin{aligned} 8 \text{ moles of sulphur contain } & \frac{8 \times 6.02 \times 10^{23}}{1} \text{ atoms} \\ & = \underline{4.8 \times 10^{23} \text{ atoms}} \end{aligned}$$

c) 0.4 g of oxygen atoms

1 mole of oxygen atom contains 6.02×10^{23} atoms

16g of oxygen contains 6.02×10^{23} atoms

1g of oxygen contains $\frac{1 \times 6.02 \times 10^{23}}{1}$ atoms

$$\begin{aligned} 0.4\text{g of oxygen contains } & \frac{1 \times 0.4 \times 6.02 \times 10^{23}}{1 \times 1} \text{ atoms} \\ & = \underline{1.5 \times 10^{24} \text{ atoms}} \end{aligned}$$

Calculate the following (Al=27, H=1, O=16, S=32, Ca=40, Na=23, Hg=201, Cl=35.5)

a) Number of atoms in 2 moles of sodium

b) Number of molecules in 5 moles of hydrogen (H_2)

c) Number of ions in 1 mole of $\text{Al}_2(\text{SO}_4)_3$

d) Number of hydroxyl ions in 2 moles of $\text{Ca}(\text{OH})_2$

2. Calculate the number of particles in the following

a) 0.1 moles of sodium atoms

b) 0.5 moles of chlorine atoms

c) 0.3 moles of calcium atoms

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3. Use the value of 6.02×10^{23} mol⁻¹ for the Avogadro constant to find the number of atoms in;
- a) 2.0×10^{-3} g of calcium
 - b) 16g of magnesium ribbon
 - c) 27g of sodium carbonate decahydrate or hydrated sodium carbonate.
 - d) 5.0×10^{-6} g of argon
 - e) 1.00×10^{-10} g of mercury
 - f) How many P₄ molecules and atoms are present in 30g of phosphorus? (P=31)
 - g) Sucrose is a compound commonly known as sugar. Its molecular formula is C₁₂H₂₂O₁₁. Calculate the number of atoms of each element in 4.32g of sucrose. (Hint; get the RMM first)

Converting number of particles to mass

Examples

1) Calculate the mass of sodium with 1.5×10^{22} sodium atoms. (Na=23, L= 6.02×10^{23} atoms)

6.02×10^{23} atoms are contained in 1 mole of sodium

6.02×10^{23} atoms are contained in 23g of sodium

1 atom is contained in $\frac{1 \times 23}{6.02 \times 10^{23}}$ g of sodium

1.5×10^{22} atoms are contained in $\frac{1 \times 23 \times 1.5 \times 10^{22}}{6.02 \times 10^{23}}$ g of sodium

= 0.575g of sodium

2) How many grams of calcium contain?

6.02×10^{23} atoms are contained in 1 mole of calcium

6.02×10^{23} atoms are contained in 40g of calcium

1 atom is contained in $\frac{1 \times 40}{6.02 \times 10^{23}}$ g of calcium

6.02×10^{23} atoms are contained in $\frac{1 \times 40 \times 6.02 \times 10^{23}}{6.02 \times 10^{23}}$ g of calcium

= 40g of calcium

3) 1.5×10^{23} atoms

6.02×10^{23} atoms are contained in 1 mole of calcium

6.02×10^{23} atoms are contained in 40g of calcium

1 atom is contained in $\frac{1 \times 40}{6.02 \times 10^{23}}$ g of calcium

1.5×10^{23} atoms are contained in $\frac{1 \times 40 \times 1.5 \times 10^{23}}{6.02 \times 10^{23}}$ g of calcium

= 10g of calcium

4) 3.0×10^{23} atoms

6.02×10^{23} atoms are contained in 1 mole of calcium

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6.02×10^{23} atoms are contained in 40g of calcium

1 atom is contained in $\frac{1 \times 40}{6.02 \times 10^{23}}$ g of calcium

3.0×10^{23} atoms are contained in $\frac{1 \times 40 \times 3.0 \times 10^{23}}{6.02 \times 10^{23}}$ g of calcium
= 20g of calcium

Self-Test

1. Convert 2.03×10^{23} particles of sodium into mass. (Na = 23)
2. 6×10^{18} particles of magnesium
3. 3×10^{13} particles of lead (II) nitrate
4. 2.4×10^{20} particles of glucose
5. 5.08×10^3 particles of methane
6. Methane is a gas used for cooking. Calculate the number of atoms of each element in in 40g of methane.
7. How many molecules are present in;
 - a) 9g of water
 - b) 17g of ammonia

Converting masses to moles

Relationships

For atoms, 1 mole is equivalent to Relative atomic mass (RAM) in grams. For example: 1 mole of Cu=64g; 1 mole of S = 32g; 1 mole of H= 1g.

For molecules, 1 mole is equivalent to relative molecular mass (RMM) or relative formula mass (RFM). For example: 1 mole of H₂O= 18g; 1 mole of CaCO₃= 100g; 1 mole of H₂SO₄= 98g.

Note;

$$\text{Number of moles} = \frac{\text{Given mass}}{\text{R.A.M}} ; \text{ for atoms}$$

$$\text{Number of moles} = \frac{\text{Given mass}}{\text{R.M.M}} ; \text{ for molecules}$$

Examples

1). Calculate the number of moles of in 4g of oxygen molecule (O₂), (O=16, S=32)

Solution

$$\text{RMM of O}_2 = (2 \times 16) = 32$$

32g is contained in 1 mole of O₂

$$1 \text{ g is contained in } \frac{1 \times 1}{32} \text{ moles of O}_2$$

$$4\text{g is contained in } \frac{1 \times 1 \times 4}{32 \times 1} \text{ moles of O}_2$$

$$= \underline{0.125 \text{ moles of O}_2}$$

160g of sulphur (S)

$$\text{RAM} = 32$$

32g is contained in 1 mole of Sulphur

$$1 \text{ g is contained in } \frac{1 \times 1}{32} \text{ moles of Sulphur}$$

$$160\text{g is contained in } \frac{1 \times 1 \times 160}{1} \text{ moles of Sulphur}$$

$$= \underline{5.0 \text{ moles of Sulphur}}$$

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2). Calculate the number of moles in the following molecules;

a) 2 g of Calcium oxide

$$\text{RMM of CaO} = 40 + 16 = 56$$

56g is contained in 1 mole of CaO

$$1 \text{ g is contained in } \frac{1 \times 1}{56} \text{ moles of CaO}$$

2g is contained in moles of CaO

$$= \underline{0.036 \text{ moles of CaO}}$$

b) 4 g of sodium hydroxide

$$\text{RMM of NaOH} = 23 + 16 + 1 = 40$$

40g is contained in 1 mole of NaOH

$$1 \text{ g is contained in } \frac{1 \times 1}{40} \text{ moles of NaOH}$$

$$2\text{g is contained in } \frac{2 \times 1 \times 1}{1 \times 40} \text{ moles of NaOH}$$

$$= \underline{0.1 \text{ moles of NaOH}}$$

Exercise

Given (O=16, Ca=40, C=12, S=32, H=1, Cu=64, Ag=108,

Mg=24) Calculate the number of moles in

a) 21.6 g of silver

b) 12g of magnesium

c) 6g of ammonia

d) 88g of carbon dioxide

e) 22.2g of calcium chloride

f) 5g of copper

Converting moles to masses

Examples

Given (O=16, Ca=40, C=12, S=32, H=1, Cu=64)

Calculate the mass in

a) 0.23 moles of sodium

1 mole of sodium weighs 23g

$$\begin{aligned} 0.23 \text{ moles of sodium weighs } & \frac{23 \times 0.23}{1} \text{ g} \\ & = \underline{5.29 \text{ g}} \end{aligned}$$

b) 7.1 moles of chlorine molecules RMM of Cl₂

$$= 35.5 \times 2$$

$$= 71$$

1 mole of chlorine weighs 71g

1 moles of chlorine weighs $\frac{1 \times 71}{1}$ g

$$= \underline{504.1 \text{ g of chlorine}}$$

c) 0.1 moles of potassium carbonate

RMM of K₂CO₃ = (39X2) + 12 + (16X3)

$$= 138$$

1 mole of potassium carbonate weighs 138g

0.1 moles of chlorine weighs $\frac{0.1 \times 138}{1}$ g

$$= \underline{13.8 \text{ g of chlorine}}$$

Exercise

Calculate the mass of;

a) 0.1 moles of sodium atom

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- b) 0.3 moles of chlorine molecules
- c) 0.125 moles of sodium carbonate
- d) 0,123 moles of calcium hydroxide
- e) 0.25 moles of sodium hydroxide
- f) 0.1 moles of sulphuric acid
- g) 0.28 moles of lead (II) nitrate
- h) 0.5 moles of iron (III) chloride
- i) 0.05 moles of calcium atoms
- j) 0.05 moles of copper (II) carbonate
- k) 0.2 moles of lead (IV) oxide
- l) Malaria is an infectious disease that affects human beings. It can easily be treated by taking artemether lumefantrine (Coartem) tablets, which have the molecular formula $C_{46}H_{58}Cl_3NO_6$. One tablet contains 20 mg of coartem. The prescription of for an infected adult is 4x2 for 3 days.
 - i. How many tablets are taken by an adult?
 - a) On daily basis
 - b) For a full dose
 - ii. On daily basis:
 - a) How many moles of coartem are swallowed?
 - b) How many molecules of coartem are swallowed?
 - c) How many atoms of carbon are orally swallowed?

Calculations percentage composition by mass

From the formula of a compound, we can calculate the percentage by mass of each element in a compound.

$$\text{Percentage composition of an element} = \frac{\text{Mass of an element}}{\text{Formula mass}} \times 100$$

Examples

Given (O=16, Ca=40, C=12, S=32, H=1, Cu=64)

Calculate the percentage composition by mass of;

a) Oxygen in calcium carbonate

Formula mass of $\text{CaCO}_3 = 40 + 12 + (3 \times 16) = 100\text{g}$

Mass of oxygen = $(3 \times 16) = 48\text{g}$

Percentage of oxygen

$$= \frac{48}{100} \times 100$$

$$= \underline{48\%}$$

b) Water in $\text{CuSO}_4 \cdot 5\text{H}_2\text{O}$

Formula mass of $\text{CuSO}_4 \cdot 5\text{H}_2\text{O} = 64 + 32 + (4 \times 16) + 5(1 \times 2 + 16) = 250\text{g}$

Mass of water, $\text{H}_2\text{O} = 5(1 \times 2 + 16) = 90\text{g}$

Percentage of oxygen

$$= \frac{90}{250} \times 100$$

$$= \underline{36\%}$$

c) Nitrogen in nitrogen dioxide

Formula mass of $\text{NO}_2 = 14 + (2 \times 16) = 46\text{g}$

Mass of nitrogen = 14g

$$\text{Percentage of oxygen} = \frac{14}{46} \times 100$$

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$$= \underline{30.4\%}$$

d) Oxygen in baking powder, NaHCO_3

$$\text{Formula mass of } \text{NaHCO}_3 = 23 + 1 + 12 + (3 \times 16) = 84\text{g}$$

$$\text{Mass of oxygen} = (3 \times 16) = 48\text{g}$$

Percentage of oxygen

$$= \frac{48}{84} \times 100$$

$$= \underline{57.14\%}$$

Exercise

Calculate the percentage of nitrogen in each of the following;

- a) Ammonium chloride (NH_4Cl)
- b) Sodium nitrate (NaNO_3)
- c) Ammonium sulphate, $(\text{NH}_4)_2\text{SO}_4$

Given that; N=14, H=1, Cl=35.5, Na=23, O=16, S=32.

2. Calculate the percentage of water of crystallization in sodium carbonate crystals, $\text{Na}_2\text{CO}_3 \cdot 10\text{H}_2\text{O}$

A metal sulphate, $\text{X}_2(\text{SO}_4)_3$ contains 28% by mass of metal X. Determine the Relative Atomic Mass of X and the Relative Molecular Mass of $\text{X}_2(\text{SO}_4)_3$. (**X=56, RMM= 390**)

Empirical and Molecular formulae Empirical formula

This is the simplest formula of a compound which expresses the ratio in which different atoms present in one molecule exists.

Examples of empirical and molecular formulae

Compound	Molecular formula	Empirical formula
Water	H ₂ O	H ₂ O
Ammonia	NH ₃	NH ₃
Ethene	C ₂ H ₄	CH ₂
Benzene	C ₆ H ₆	CH
Glucose	C ₆ H ₁₂ O ₆	CH ₂ O

Steps involved in calculating empirical formula

1. Write down the symbols of the elements present
2. Write down the percentage composition or composition by mass below the symbols
3. Find the number of moles of each element by dividing the percentage composition or mass by Relative Atomic Mass

$$\text{Number of moles} = \frac{\text{mass}}{\text{RMM or RAM}}$$

Find the mole ratio of the elements by dividing the moles with the smallest number

4. Write down the empirical number.

If the mole ratio is in fractions;

- a) Round off to the nearest whole number if it is very close to the whole number.
- b) Multiply by a small number that converts the fraction to a whole number if the fraction is not close to a whole number.

Molecular formula

This is a formula that shows the actual number of each atom present in one molecule of a compound. The molecular formula is a multiple of the empirical formula, so, from the empirical formula, the molecular formula can be determined.

Molecular formula= (Empirical formula) n = Molecular mass; where, n is number to be determined

Examples

1. a) Calculate the empirical formula of a compound containing 80% carbon and 20% hydrogen.
b) If the molar mass of the compound is 30g, determine its molecular formula.

Solution

a)

Elements	C	H
Percentage composition	80	20
Number of moles	$\frac{80}{12} = 6.7$	$\frac{20}{1} = 20$
Divide by smallest	$\frac{6.7}{6.7}$	$\frac{20}{6.7}$
Mole ratio	1	3

The empirical formula is CH₃

b)

(Empirical formula) n = Molar mass

$$(\text{CH}_3) n = 30$$

$$(12+3) n = 30$$

$$15n = 30$$

$$n = 2$$

The molecular formula is therefore, (CH₃)₂ or C₂H₆

1. Calculate the empirical formula of a compound containing 28% of iron, 24% sulphur and the rest being oxygen. (Fe = 56, S = 32, O = 16)

Solution

Percentage composition of oxygen

$$= 100 - (28 + 24)$$

$$= 100 - 52$$

$$= 48\%$$

Elements present	Fe	S	O
Percentage composition	28	24	48
Number of moles	$\frac{28}{56} = 0.5$	$\frac{24}{32} = 0.75$	$\frac{48}{16} = 3$
Divide by smallest	$\frac{0.5}{0.5}$	$\frac{0.75}{0.5}$	$\frac{3}{0.5}$
Mole ratio	1	1.5	6
Multiply by 2 throughout	2	3	12

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The empirical formula is $\text{Fe}_2\text{S}_3\text{O}_{12}$ or $\text{Fe}_2(\text{SO}_4)_3$

2. A hydrocarbon contains 85.7% carbon and its relative molecular mass is 28. Work out its molecular formula.

Solution

Percentage of hydrogen $100\% - 85.7\% = 14.3\%$

Elements present	C	H
Percentage composition	85.7	14.3
Number of moles	$\frac{85.7}{12} = 7.1$	$\frac{14.3}{1} = 14.3$
Divide by smallest	$\frac{7.1}{7.1}$	$\frac{14.3}{7.1}$
Mole ratio	1	2

The empirical formula is CH_2

(Empirical formula) n = Molar mass $(\text{CH}_2) n = 28$

$$(12+2) n = 28$$

$$14n = 28$$

$$n = 2$$

The molecular formula is therefore $(\text{CH}_2)_2$ or C_2H_4

Exercise

1. A compound X consists of carbon 40%, hydrogen 6.7% and the rest being oxygen. If the RMM is 60, determine its molecular formula. (C=12, H=1, O=16) (**Ans. CH_2O**)
2. A hydrocarbon is made up of 92.3% carbon and has molecular formula of 78g. Calculate its empirical and molecular formula. (**Answer CH**)
3. Calculate the empirical formula of the compound formed when 1.8g of carbon forms 2.4g of a hydrocarbon. (**Answer CH_4**)
4. Given that 0.24g of magnesium reacted with 0.16g of oxygen. Find the empirical formula. (O=16, Mg=24).
5. Calculate the molecular formula of a hydrocarbon with empirical formula CH_2 and molecular mass of 28g. (Answer C_2H_4)

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6. Calculate the empirical formula of a salt with the following composition, copper 25%, sulphur 12.8%, oxygen 25.6% and water 36.0% (Answer $\text{CuSO}_4 \cdot 5\text{H}_2\text{O}$)
7. Calculate the empirical formula of a hydrated salt with the following composition, sodium 16.09%, carbon 9.20%, oxygen 16.78% and water 62.93% (Answer $\text{Na}_2\text{C}_2\text{O}_3 \cdot 10\text{H}_2\text{O}$)
8. Find the empirical formulae of the compounds formed in the reactions described below.
 - a) 10.800g magnesium form 18.000g of an oxide (Answer= MgO)
 - b) 3.400g calcium form 9.435g of a chloride (Answer= CaCl_2)
 - c) 3.528g iron form 10.237g of a chloride. (Answer = FeCl_3)
9. Calculate the empirical formulae of the compounds from which the following analytical results were obtained.
 - a) 27.3%C, 72.7%O (Answer= CO_2)
 - b) 53.0%C, 47.0%O (Answer= C_3O_2)
 - c) 29.1%Na, 40.5%S,30.4%O (Answer= $\text{Na}_2\text{S}_2\text{O}_3$)
 - d) 32.4%Na, 22.5%S, 45.0%O (Answer= Na_2SO_4)

Calculations of masses from equations

Moles and mole ratios can be used to calculate the number of substances reacting and products formed. This requires that a correctly balanced equation is written. Such an equation is known as a **stoichiometric equation**.

STOICHIOMETRY

Stoichiometry is the relationship between amounts of reactants and products in a chemical reaction. A stoichiometric equation is an equation in which the reactants and products are correctly balanced.

Steps involved in the calculation

1. Write down a balanced equation for the reaction
2. Write down the moles of substances that concerns the question
3. Convert the moles into grammes

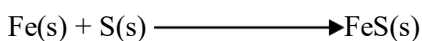
Examples

1. Calculate the mass of iron (II) sulphide formed by heating 64g of sulphur with excess iron filling. (S=32, Fe=56)

Solution

Molecular mass of FeS= 56 + 32 = 88g

Equation for reaction;



1 mole of sulphur forms 1 mole of Iron (II)sulphide

32g of S forms 88g of FeS

1 g of S forms $\frac{1 \times 88}{32}$ g of FeS

64g of S forms $\frac{64 \times 1 \times 88}{32}$ g of FeS
= 176g of FeS

2. What is the mass of magnesium required to form 55g of magnesium oxide.

Solution

Molecular mass of MgO = 24 + 16 = 40g

Equation for reaction;

(2x40) g of MgO is formed by (2x24) g of Mg



2 moles of magnesium oxide are formed 2 moles of magnesium

(2x40) g of MgO is formed by (2x24) g of Mg

80g of MgO is formed by 48g of Mg

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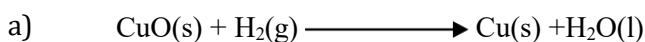
1 g of MgO is formed by $\frac{1 \times 48}{80}$ g of Mg

55g of MgO is formed by $\frac{1 \times 48 \times 55}{1 \times 80}$ g of Mg

$$= \underline{33 \text{ g of Mg}}$$

3. a) Calculate the mass of copper formed when 3.2g of copper (II)oxide is completely reduced to the metal by hydrogen gas.
- b) How many grams of water was produced
- c) Calculate the mass of hydrogen used in the experiment

Solution



1 mole of CuO forms 1 mole of Cu (64+16) g of CuO forms 64g of Cu

80g of CuO forms 64g of Cu

1 g of CuO forms $\frac{1 \times 64}{80}$ g of Cu

3.2g of CuO forms $\frac{32 \times 1 \times 64}{1 \times 80}$ g of Cu

$$= \underline{2.6 \text{ g of Cu}}$$

b) 1 mole of CuO forms 1 mole of H₂O

(64+16) g of CuO forms (1x2+16) g of H₂O

80g of CuO forms 18g of H₂O

1 g of CuO forms $\frac{1 \times 18}{80}$ g of H₂O

3.2g of CuO forms $\frac{1 \times 18 \times 3.2}{1 \times 80}$ g of H₂O

$$= \underline{0.7 \text{ g of H}_2\text{O}}$$

c) 1 mole of CuO reacts with 1 mole of H₂

(64+16) g of CuO reacts with (1x2) g of H₂

80g of CuO reacts with 2g of H₂

1 g of CuO reacts with $\frac{1 \times 2}{80}$ g of H₂

3.2g of CuO reacts with $\frac{1 \times 2 \times 3.2}{1 \times 80}$ g of H₂

$$= \underline{0.08 \text{ g of H}_2}$$

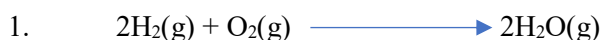
Exercise

1. A solution of 8.1g of NaOH was neutralized by hydrochloric acid. Calculate the mass of sodium chloride produced when the solution was evaporated to dryness.
(C=12, Na=23, O=16, H=1, Cl=35.5,) (**Answer =11.85g**)
2. Calculate the mass of residue left when 2.40g of sodium hydrogen carbonate is decomposed by heat. (**Answer =1.51g**)
3. Calculate the loss in mass when 100g of calcium carbonate is heated to constant mass. (Ca=40, C=12, O=16) (**Answer =44g**)
4. 76.5g of calcium hydrogen carbonate was heated strongly. What was the mass of carbon dioxide formed? (**Answer =20.78g**)
5. What mass of sodium oxide would be made from 1.5 g of sodium? (**Answer=2.02g**)

Avogadro's law

The law states that *“equal volume of gases at the same temperature and pressure contains the same number of molecules”*

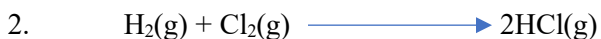
Avogadro's law gives an interpretation to Gay-Lussac's law in terms of molecules of gases. Consider the following examples.



2 volumes of hydrogen combine with 1 volume of oxygen to form 2 volumes of steam

Is interpreted as;

2 molecules of hydrogen combine with 1 molecule of oxygen to form 2 molecules of steam



1 volume of hydrogen combines with 1 volume of chlorine to form 2 volumes of hydrogen chloride gas

Is interpreted as

1 molecule of hydrogen combines with 1 molecule of chlorine to form 2 molecules of hydrogen chloride gas.

This law is quite important because it enables us to change from a statement about volumes of gases to the same statement about moles of gases and vice versa.

The relationship between vapor density and Relative Molecular Mass can be deduced from Avogadro's law and is expressed as;

$2 \times \text{Vapor Density} = \text{Relative Molecular Mass}$

I.e. Relative Molecular Mass is twice the value of vapor density.

It follows from Avogadro's law that if equal volumes of gases contain equal numbers of molecules, then the volume occupied by one mole must be the same for all gases. It is called the **gas molar volume**.

Gas molar volume

The molar gas volume is the volume occupied by one mole of a gas. It is the same for all gases under the same conditions of temperature and pressure. Values for molar gas volumes are given in the table below.

Condition	Value for molar gas volume
Standard temperature and pressure (s.t.p)	22.4l or 22.4dm ³ or 22400cm ³
Room temperature and pressure (r.t.p)	24l or 24dm ³ or 24000cm ³

Example

1. Determine the number of moles in the following gaseous volumes at s.t.p;

- a) 1.2 dm³ of nitrogen
- b) 300cm³ of ammonia

Solution

a) 22.4 dm³ is occupied by 1 mole of nitrogen

1 dm³ is occupied by $\frac{1 \times 1}{22.4}$ moles of nitrogen

1.2 dm³ is occupied by moles of nitrogen

$$= \underline{0.05 \text{ moles of nitrogen at s.t.p}}$$

b) 22400 cm³ is occupied by 1 mole of ammonia

1 cm³ is occupied by $\frac{1 \times 1}{22400}$ moles of ammonia

300 cm³ is occupied by $\frac{300 \times 1 \times 1}{1 \times 22400}$ moles of ammonia

$$= \underline{0.013 \text{ moles of ammonia at s.t.p}}$$

2. Determine at r.t.p the volume and mass of

- a) 0.04 moles of hydrogen
- b) 0.2 moles of carbon dioxide

Solution

a) 1 mole of hydrogen occupies 24l at r.t.p

0.04 moles of hydrogen occupy $\frac{0.04 \times 24}{1}$ l at r.t.p

$$= \underline{0.96 \text{ l of hydrogen at r.t.p}}$$

RMM for H₂ = (1x2) = 2

1 mole of hydrogen weighs 2g

0.04 moles of hydrogen weighs $\frac{0.04 \times 2}{1}$ g

$$= \underline{0.08 \text{ g of hydrogen}}$$

b) 1 mole of carbon dioxide occupies 24l at r.t.p

0.2 moles of carbon dioxide occupy $\frac{0.2 \times 24}{1}$ l at r.t.p

$$= \underline{4.8 \text{ l of carbon dioxide at r.t.p}}$$

RMM for CO₂ = 12 + (16x2) = 44

1 mole of carbon dioxide weighs 44g

0.2 moles of carbon dioxide weighs $\frac{0.2 \times 44}{1}$ g

$$= \underline{8.8 \text{ g of carbon dioxide}}$$

3. Calculate the RMM of Y given that 0.8 g of Y occupies 560cm³ at s.t.p.

Solution

560 cm³ of Y weighs 0.8g

1 cm³ of Y weighs $\frac{1 \times 0.8}{560}$ moles of nitrogen

22400 cm³ of Y weighs $\frac{1 \times 0.8 \times 22400}{1 \times 560}$ moles of nitrogen

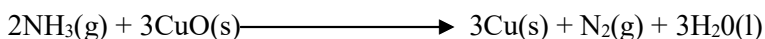
$$= 32 \text{ g}$$

The RMM of Y is 32.

Calculation of masses and volumes

Examples

1. 0.2 moles of ammonia gas measured at s.t.p were passed over copper (II)oxide. The equation for the reaction is;



Calculate;

- The number of moles of copper (II)oxide used
- The mass of copper formed
- The volume of ammonia used at s.t.p

(Cu=64, O=16, N=14, H=1, 1 mole of a gas occupies 22.4 l at s.t.p)

Solution

- a). From the equation;

2 moles of NH₃ reacts with 3 moles of CuO

1 mole of NH₃ reacts with $\frac{1 \times 3}{2}$ moles of CuO

0.2 moles of NH₃ reacts with $\frac{0.2 \times 1 \times 3}{1 \times 2}$ moles of CuO

$$= \underline{0.3 \text{ moles of CuO}}$$

- b). From the equation;

2 moles of NH₃ produce 3 moles of Cu

2 moles of NH₃ produce (3x64) g of Cu

1 mole of NH₃ produces $\frac{1 \times 3 \times 64}{2}$ moles of Cu

0.2 moles of NH₃ produces $\frac{0.2 \times 1 \times 3 \times 64}{1 \times 2}$ moles of Cu

$$= \underline{19.2 \text{ g of Cu}}$$

- c). 1 mole of NH₃ occupies 22.4l at s.t.p

0.2 moles of NH₃ occupies $\frac{0.2 \times 22.4}{1}$ l at s.t.p

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$$= \underline{4.48 \text{ l of NH}_3 \text{ at s.t.p}}$$

2). 2.5 g of CuCO_3 were heated to constant mass. Determine;

- a) The mass of the residue
- b) The volume of gas produced at s.t.p

(Cu=64, O=16, C=12, 1 mole of a gas occupies 24 dm³ at s.t.p)

Solution



a). From the equation

1 mole of CuCO_3 produces 1 mole of CuO

(64+12+16x3) g of CuCO_3 produces (64+16) g of CuO

124g of CuCO_3 produces 80g of CuO

1 g of CuCO_3 produces $\frac{1 \times 80}{124}$ g of CuO

2.5g of CuCO_3 produces $\frac{1 \times 80 \times 2.5}{124 \times 1}$ g of CuO

$$= \underline{1.61 \text{ g of CuO (residue)}}$$

b). From the equation

1 mole of CuCO_3 produces 1 mole of CO_2

(64+12+16x3) g of CuCO_3 produces 24dm³ of CO_2

124g of CuCO_3 produces 24dm³ of CO_2

1 g of CuCO_3 produces $\frac{1 \times 24}{124}$ dm³ of CO_2

2.5 g of CuCO_3 produces $\frac{1 \times 24 \times 2.5}{124 \times 1}$ dm³ of CO_2

$$= \underline{0.484 \text{ dm}^3 \text{ of CO}_2}$$

3). From the equation;



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Calculate the;

- Volume of chlorine at s.t.p required to react with 8 g of iron
- Mass of iron (III) chloride formed

Solution

a). From the equation;

2 moles of Fe react with 3 moles of Cl₂

(2x56) g of Fe react with (3x22.4) dm³ of Cl₂

112g of Fe react with 67.2dm³ of Cl₂

1 g of Fe react with $\frac{1 \times 67.2}{112}$ dm³ of Cl₂

8 g of Fe react with $\frac{1 \times 67.2 \times 8}{112 \times 1}$ dm³ of Cl₂

$$= \underline{4.8 \text{ dm}^3 \text{ of Cl}_2 \text{ at s.t.p}}$$

b). From the equation;

2 moles of Fe produce 2 moles of FeCl₂

(2x56) g of Fe produces 2x(56+35.5x2) g of FeCl₂

112g of Fe produces 325g of FeCl₂

1 g of Fe produces $\frac{1 \times 325}{112}$ g of FeCl₂

8 g of Fe produces $\frac{1 \times 325 \times 8}{112 \times 1}$ g of FeCl₂

$$= \underline{23.01 \text{ g of FeCl}_2}$$

Exercise

- What volume of hydrogen at s.t.p would be used if 40g of nitrogen combined with it to form ammonia.
- 0.1 mole of lead nitrate is completely decomposed on heating according to the reaction below;



- What volume at s.t.p of nitrogen dioxide and oxygen were evolved

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- b) What was the mass of the residue left
- 2). 1.4 litres of ammonia gas measured at s.t.p were passed over hot copper (II)oxide. The equation for the reaction is



Calculate;

- a) The number of moles of ammonia used
- b) The number of moles of nitrogen gas used
- c) The mass of copper formed
- 3). 3.1g of a carbonate XCO_3 are heated to constant mass; 2.0g of the metal oxide are formed. When heated in a stream of hydrogen for sufficient time, the oxide is reduced to 1.6g of pure metal.
- a) Give the equation for the two reactions
- b) Calculate the minimum volume of hydrogen at s.t.p needed to reduce the oxide
- c) Calculate the volume of carbon dioxide at s.t.p produced when the carbonate was completely decomposed.
- 4). Calculate the mass of sulphur deposited when 8.4 dm^3 of chlorine oxidizes hydrogen sulphide.

Volumetric analysis

In an investigation to determine the nature of matter in a substance, a chemist focuses on two main questions:

- a) What are the components of the substance?
- b) What amount of each component is present in the substance?

In an attempt to answer question (a) and (b), a chemist carries out **qualitative** and **quantitative analysis** respectively. Most of the reactions which a chemist carries out take place in solution.

Consider the reaction below;



If we are to determine the volume of A needed to completely react with a given amount or volume of B, the answer is provided practically through volumetric analysis.

In volumetric analysis, quantities of substances (often acids or alkalis) are estimated by analytical processes involving measurements of volumes of solutions using pipettes, burettes and measuring cylinders (for approximate measurement). Weighing may also be involved. Most of the work in volumetric analysis is based upon molar (M) solutions.

Standard and Molar solutions

A standard solution

This is a solution of known concentration. Examples of standard solutions are; solution containing 12g of sodium chloride in one litre of a solution; a solution containing 2 moles of solute in 1 dm³ etc.

The substance that is used to prepare a standard solution is known as a **primary standard**.

A Molar solution

This is a solution that contains one mole of a substance in a solution of one litre. In other words, it is a solution containing one mole of solute in one litre.

Other related terms are;

Concentration; this is the amount of solutes in a given volume of solution.

Molarity; this is the number of moles of solute in one litre of a solution. The unit is mol/dm³ or mol/l. The molarity of a solution is commonly denoted by letter M, e.g. 0.2M NaOH which mean 1 litre of a solution containing 0.2 moles of NaOH. 1 litre (1 l) = 1 cubic decimetre (1 dm³) = 1000 cubic centimetre (1000 cm³)

Calculations on molarity and masses

Examples

Calculate the molarities of the following solution given (Na=23, O=16, H=1, C=12, Cl=35.5, S=32)

- a) 13.5g of copper (II) chloride in 1dm³ of solution
- b) 4.0g of sodium hydroxide in 400cm³ of solution
- c) 53g of anhydrous sodium carbonate in 2 dm³ of solution

Solution

a) RMM of CuCl₂
= 64+(35.5x2)
= 135

135g is contained in 1 mole of CuCl₂

1 g is contained in $\frac{1 \times 1}{135}$ moles of CuCl₂

13.5g is contained in $\frac{1 \times 1 \times 13.5}{135 \times 1}$ moles of CuCl₂
= 0.1M of CuCl₂

b) 400cm³ of solution contains 4.0g of NaOH

c) 1cm³ of solution contains $\frac{1 \times 4.0}{400}$ g of NaOH

1000cm³ of solution contains $\frac{1 \times 4.0 \times 1000}{400 \times 1}$ g of NaOH
= 10g/l of NaOH (concentration in grams/litre)

RMM of NaOH

= 23+16+1
= 40

40g is contained in 1 mole of NaOH

1 g is contained in $\frac{1 \times 1}{40}$ moles of NaOH

10 g is contained in $\frac{1 \times 1 \times 10}{40 \times 1}$ moles of NaOH
= 0.25M NaOH

d) 2dm³ of solution contains 53g of Na₂CO₃

1dm³ of solution contains $\frac{1 \times 53}{2}$ g of Na₂CO₃
= 26.5g/dm³ of Na₂CO₃ (concentration in grams/litre)

RMM of Na₂CO₃= 23x2+12+16x3

= 106

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106g is contained in 1 mole of Na_2CO_3

1 g is contained in $\frac{1 \times 1}{106}$ moles of Na_2CO_3

10 g is contained in $\frac{1 \times 1 \times 10}{106 \times 1}$ moles of Na_2CO_3
 $= \underline{0.25\text{M Na}_2\text{CO}_3}$

In general, **Molarity** = $\frac{\text{Concentration } (\frac{\text{g}}{\text{l}})}{\text{Molar mass}}$

When the concentration of a solute in grams per litre and the RMM are known then the molarity can be calculated from the above expression.

Note;

The use of formula is not so much recommended and workings should be from first principle.

2. Calculate the mass of the named substance needed to make;
 - a) 0.1 dm^3 of 2M sodium sulphate solution
 - b) 1 l of 0.25M sodium hydroxide solution
 - c) 25 cm^3 of 0.1M potassium carbonate solution
 - d) 500 cm^3 of 0.05M sodium carbonate solution

Solution

a) 1 dm^3 of solution contains 2 moles of Na_2SO_4

0.1 dm^3 of solution contains $\frac{0.1 \times 2}{1}$ moles of Na_2SO_4
 $= 0.2$ moles of Na_2SO_4

RMM for Na_2SO_4

$$= 23 \times 2 + 32 + 16 \times 4$$

$$= 142$$

1 mole of Na_2SO_4 weighs 142g

0.1 moles of Na_2SO_4 weighs $\frac{0.1 \times 142}{1}$ g
 $= \underline{28.4\text{g}}$

b) 1l of solution contains 0.25 moles of NaOH

RMM for NaOH

$$= 23 + 16 + 1$$

$$= 40$$

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1 mole of NaOH weighs 40g

0.25 moles of NaOH weighs $\frac{0.25 \times 40}{1}$ g

$$= \underline{10\text{g}}$$

c) 1000cm³ of solution contains 0.1moles of K₂CO₃

1 cm³ of solution contains $\frac{1 \times 0.1}{1000}$ moles of K₂CO₃

25 cm³ of solution contains $\frac{1 \times 0.1}{1000}$ moles of K₂CO₃

$$= 0.0025 \text{ moles of K}_2\text{CO}_3$$

RMM of K₂CO₃

$$= 39 \times 2 + 12 + 16 \times 3$$

$$= 138$$

1 mole of K₂CO₃ weighs 138g

0.0025 moles of K₂CO₃ weighs $\frac{0.0025 \times 138}{1}$ g

$$= \underline{0.345\text{g}}$$

Calculating number of moles of ions in standard solutions

Examples

1). Calculate the number of moles of hydrogen ions in 25cm³ of a 0.2 M sulphuric acid.

2). Calculate the number of moles of potassium ions in 35cm³ of 0.12 M potassium carbonate solution.

Solution

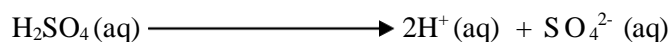
1. 1000cm³ of solution contains 0.2moles of H₂SO₄

1 cm³ of solution contains $\frac{1 \times 0.2}{1000}$ moles of H₂SO₄

25 cm³ of solution contains $\frac{1 \times 0.2 \times 25}{1000 \times 1}$ moles of H₂SO₄

$$= 0.005 \text{ moles of H}_2\text{SO}_4$$

From the equation of ionization of H₂SO₄,



1 mole of H₂SO₄ produces 2 moles H⁺

0.005 moles of H₂SO₄ produces $\frac{0.005 \times 2}{1}$ moles H⁺

$$= \underline{0.01 \text{ moles H}^+}$$

2. 1000cm³ of solution contains 0.12 moles of K₂CO₃

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1 cm³ of solution contains $\frac{1 \times 0.12}{1000}$ moles of K₂CO₃

35 cm³ of solution contains $\frac{1 \times 0.12 \times 35}{1000 \times 1}$ moles of K₂CO₃

$$= 0.0042 \text{ moles of K}_2\text{CO}_3$$

From the equation of ionization of K₂CO₃,



1 mole of K₂CO₃ produces 2 moles K⁺

0.0042 moles of K₂CO₃ produces $\frac{0.0042 \times 2}{1}$ moles K⁺

$$= \underline{0.0084 \text{ moles K}^+}$$

ACID-BASE REACTION (TITRATION)

This is a method of volumetric analysis in which a solution (usually a standard solution) is added from a burette to another solution (usually whose concentration is unknown) until the reaction is complete. An indicator changes color immediately the reaction is complete or when the **end point** is reached. Most titrations at this level are acid-base titrations.

Common laboratory indicators and color changes

Indicator	Color in acidic solution	Color in alkaline solution
Phenolphthalein	Colorless	Purple
Methyl orange	Red/pink	Yellow
Litmus	Red	Blue

Procedure for acid-base titration

- ❖ Wash the pipette with distilled water. Use the pipette to deliver either 20.0cm³ or 25.0cm³ of the alkali into a clean conical flask. Add a few drops (2 or 3 drops) of indicator.
- ❖ Wash the burette with distilled water then a little of the Level with acid solution and run out the acid through the tap. Fill Burette meniscus the burette above the 0 cm³ mark and run a little of the acid out to bring the meniscus of the acid to the 0 cm³ mark or slightly below it. Take the burette reading as V₁ cm³.
- ❖ Run the acid solution from the burette drop wise. Use your left hand to open the tap and your right hand to swirl the conical flask (unless you are left-handed). Stop when the indicator just changes color. This is the end point the titration.
- ❖ Take the burette reading again V₂ cm³. Subtract (V₂- V₁) cm³ to get the —**titre** (i.e. the volume of the acid needed to neutralize the known volume of alkali).
- ❖ Repeat the titration. Obtain an average titre. From this volume you can calculate the unknown concentration.

Note;

The first titration is regarded as a trial run (rough titration) and may not be very accurate, therefore the value may not be used in computing the average volume. Values used in calculating the average volume must be close to each other.

Specimen readings

Example

Neutralization of 0.1M NaOH solution with a solution of HCl

Volume of pipette used = 25.0cm³

Number of titrations	1	2	3
Final burette reading/ cm ³	14.80	30.00	15.00
Initial burette reading/ cm ³	0.00	15.00	0.00
Volume of acid used/ cm ³	14.80	15.00	15.00

Value used to calculate average volume of acid used: 15.00 cm³ and 15.00 cm³

Average volume of acid used;

$$\begin{aligned} &= \frac{15.00 + 15.00}{2} \\ &= 15.00 \text{ cm}^3 \end{aligned}$$

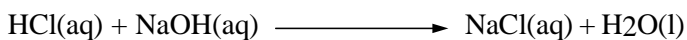
Calculate;

- The number of moles of sodium hydroxide that reacted
- The number of moles of hydrochloric acid that reacted
- The molarity of the hydrochloric acid (i.e. concentration in mol/litre)

Method of calculation

a).

Write the equation for the reaction that took place. This gives you the mole ratio of reaction between the acid and the alkali.



1 mole of HCl neutralizes 1 mole of NaOH

Work out the number of moles of the standard solution. In this case it is the alkali (NaOH) as its concentration is known (25.0 cm³ contains of 0.1M NaOH).

Moles of NaOH that reacted

1000cm³ of solution contains 0.1 moles of NaOH

1 cm³ of solution contains $\frac{1 \times 0.1}{1000}$ moles of NaOH

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$$25 \text{ cm}^3 \text{ of solution contains } \frac{1 \times 0.1 \times 25}{1000 \times 1} \text{ moles of NaOH} \\ = \underline{0.0025 \text{ moles of NaOH}}$$

b).

Work out the number of moles of the acid that reacted by relating the number of moles of the alkali to the mole ratio of reaction between the acid and alkali.

From the equation of reaction;

1 mole of NaOH reacts with 1 mole of HCl

$$0.0025 \text{ moles of NaOH reacts with } \frac{0.0025 \times 1}{1} \text{ moles of HCl} \\ = \underline{0.0025 \text{ moles of HCl}}$$

(since the mole ratio of the reaction of the HCl: NaOH is 1:1, so, the number of moles of HCl = 0.0025 moles)

c).

Now work out the molarity of the acid

15cm³ of solution contains 0.0025moles of HCl

1cm³ of solution contains $\frac{1 \times 0.0025}{15}$ moles of HCl

$$1000\text{cm}^3 \text{ of solution contains } \frac{1 \times 0.0025 \times 1000}{1 \times 15} \text{ moles of HCl} \\ = \underline{0.167 \text{ mol/l of HCl}}$$

The molarity of the HCl is 0.167M

4.

From the molarity, you can proceed and work out the concentration in grams/litre, if the Molar mass is known. (e.g. for 0.167M HCl in the above case)

Molar mass of HCl = (1+35.5) g = 36.5g

1 mole of HCl weighs 36.5g

$$0.167 \text{ moles of HCl weighs } \frac{0.167 \times 36.5}{1} \text{ g} \\ = \underline{6.1 \text{ g/l}}$$

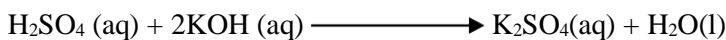
Other examples

1. 25cm³ of sulphuric acid of concentration 0.15mol/dm³ neutralized 31.2cm³ of potassium hydroxide solution. Find the concentration of the KOH solution in mol/l and in grams/litre.

Solution

Write the equation for the reaction

Equation for the reaction;



1 mole of acid reacts with 2 moles of alkali

Calculate the number of moles of the acid (standard solution) that reacted

1000cm³ of solution contains 0.15 moles of H₂SO₄

1 cm³ of solution contains $\frac{1 \times 0.15}{1000}$ moles of H₂SO₄

25 cm³ of solution contains $\frac{1 \times 0.15 \times 25}{1000 \times 1}$ moles of H₂SO₄

$$= \underline{0.00375 \text{ moles of H}_2\text{SO}_4}$$

Relate the number of moles of acid to the mole ratio of the reaction to find the number of moles of the alkali that reacted

1 mole of H₂SO₄ reacts with 2 moles of KOH

0.00375 moles of H₂SO₄ reacts with $\frac{0.00375 \times 2}{1}$ moles of KOH

$$= \underline{0.0075 \text{ moles of KOH}}$$

Calculate the molarity (concentration in mol/dm³) of the alkali

31.2cm³ of solution contains 0.0075 moles of KOH

1cm³ of solution contains $\frac{1 \times 0.0075}{31.2}$ moles of KOH

1000cm³ of solution contains $\frac{1 \times 0.0075 \times 1000}{31.2 \times 1}$ moles of KOH

$$= 0.24 \text{ mol/l of KOH}$$

The concentration of the KOH is 0.24 mol/l of KOH

Proceed and calculate the concentration in g/l

Molar mass of KOH = (39+16+1) = 56g

1 mole of KOH weighs 56g

0.24 moles of KOH weighs $\frac{0.24 \times 56}{1}$ g

$$= 13.44\text{g}$$

The concentration of the KOH is 13.44 g/l

2. Determine the molarity of acids in the following solutions.

a) 16.0cm³ of 1.5M sodium hydroxide neutralized by 20.0cm³ of hydrochloric acid

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b) 25.0cm³ of 0.2M ammonia solution neutralized by 20.0cm³ of nitric acid

Solution

a). Equation of reaction



1000cm³ of solution contains 1.5moles of NaOH

1 cm³ of solution contains $\frac{1 \times 1.5}{1000}$ moles of NaOH

16 cm³ of solution contains $\frac{1 \times 1.5 \times 16}{1000 \times 1}$ moles of NaOH
= 0.024 moles of NaOH

From the equation;

1 mole of NaOH reacts with 1 mole of HCl

0.024 moles of NaOH reacts with $\frac{0.024 \times 1}{1}$ moles of HCl
= 0.024 moles of HCl

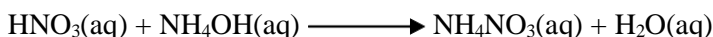
20.0cm³ of solution contains 0.024moles of HCl

1cm³ of solution contains $\frac{1 \times 0.024}{20.0}$ moles of HCl

1000cm³ of solution contains $\frac{1 \times 0.024 \times 1000}{20.0 \times 1}$ moles of HCl
= 1.2 M HCl

The concentration of the HCl is 1.2M

b). Equation for reaction;



1000cm³ of solution contains 0.2moles of NH₄OH

1 cm³ of solution contains $\frac{1 \times 0.2}{1000}$ moles of NH₄OH

25cm³ of solution contains $\frac{1 \times 0.2 \times 25}{1000 \times 1}$ moles of NH₄OH
= 0.005 moles of NH₄OH

From the equation;

1 mole of NH₄OH reacts with 1 mole of HNO₃

0.005 moles of NH₄OH reacts with moles of HNO₃
= 0.005 moles of HNO₃

20.0cm³ of solution contains 0.005 moles of HNO₃

1cm³ of solution contains $\frac{1 \times 0.005}{20.0}$ moles of HNO₃

1000cm³ of solution contains moles of HNO₃
= 0.25 M HNO₃

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The concentration of the HNO_3 is 0.25M

Activity of integration

In most activities we do at homes like preparing meals, preparing juice, we use the knowledge of stoichiometry and moles because there is need to take measurements and volumes of the substances we use.

You expect Tr. Kisule and other seven guests in your home to attend your sister's birthday. Your parents have assigned you a role of taking charge of the planning and preparation of the different food stuffs for the lunch that will be served on the birthday.

Your parents expect a nice dish, meat sauce, fruit salad and cocktail juice to be prepared on that day.

Write a report to your parents explaining the quantities of terms needed for the lunch. (think of the support materials)

