MOLE CONCEPT AND STOICHIOMETRY DETAILED NOTES AND SAMPLE QUESTIONS

(I regret for any mistake if noted**)**

S3 TERM TWO TOPIC

CBC NEW LOWER SECONDARY CURRICULUM

(CHEMISTRY)

BY

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DEDICATED TO YOU

The attached questions are almost enough for a student to have a general idea/concept about this region(**content/subtopic**) in chemistry, however, I advise a student to search for more related questions about this content area for **better results**.

CONTENT:

1. ***Mole concept and Stoichiometry***
2. ***Some sample questions on the above topic***
3. **Try so hard to answer the sample questions and look for more qns.**

Don't say tomorrow, it will be too late for chemistry revision, and yesterday is gone forever, you have got today to revise your chemistry!
"Revise as if tomorrow is not there"
May god bless you

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 $H_2SO_4(aq)$ molecules contains 6.02×10^{23} molecules of $H_2SO_4(aq)$.

A **mole** is the amount of substance which contains 6.02×10^{23} particles.

Or

A mole is 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.

Do you still remember isotopes and isotopy???? Lets look at them precisely.

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

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	$^{1}_{1}H$ $^{2}_{1}H$ $^{3}_{1}H$	99.99% 0.01% Rare
Carbon	C	6	$^{12}_{6}C$ $^{13}_{6}C$ $^{14}_{6}C$	98.9% 1.1% Trace
Chlorine	Cl	17	$^{35}_{17}Cl$ $^{37}_{17}Cl$	75% 25%

NB;

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).

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 1/12 (a twelfth) of the mass of one atom of Carbon 12 isotope.

NB:

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

$$\text{R.A.M} = \frac{mass\ of\ an\ atom\ of\ an\ element}{\frac{1}{12} of\ the\ mass\ of\ one\ atom\ of\ C{-}12\ isotope}$$

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

Solution

$\left(\frac{90.5\times20}{100}\right)+\left(\frac{0.3\times21}{100}\right)+\left(\frac{9.2\times22}{100}\right)$

$\left(\frac{(90.5\times20)+(0.3\times21)+(9.2\times22)}{100}\right)$

$\left(\frac{2018.7}{100}\right)$ =20.187.

Examples 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 (M_r) or molecular mass/RMM (M_r)

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 ,H_2O (H=1,O=16)
=(1x2) + 16
=18

b) Oxygen molecule, O_2 (0=16)
=2x16
=32

c) Sodium sulphate, Na_2SO_4 (Na=23,S=32,0=16)
=(2x23)+32+(4x16)
=46+32+64
=142

d) Copper (II) sulphate crystals, $CuSO_4.5H_2O$ (Cu=64,S=32,O=16,H=1)
=64+32+(4x16)+5(2x1+16)
=250

Exercise

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

a) $FeSO_4.7H_2O$ (Fe=56,S=32,O=16,H=1)
b) $(NH_4)_2SO_4$ (N=14,H=1,S=32,O=16)
c) Al_2O_3 (Al=27,O=16)
d) $Ca(HCO_3)_2$ (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 H_20 = 6.02×10^{23} molecules

1mole of H_2O contains 2 moles of H atoms, therefore, number of H atoms in 1 mole of H_2O = $(2X6.02\times10^{23})$ =$12.04x10^{23}$ atoms.

Number of O atoms in 1 mole of H_2O =6.02×10^{23}.

b) 2 moles of oxygen molecules, ($2O_2$)

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

2 moles of oxygen molecules contains $(2x6.02\times10^{23})$= $12.04x10^{23}$ molecules

1 mole of oxygen molecule contains 2 moles of oxygen atoms

2 moles of oxygen molecules contains (2x2)= 4 moles of oxygen atoms

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

4 moles of oxygen atoms contains $(4x6.02\times10^{23})$= $24.08x10^{23}$ atoms.

c) Number of hydrogen ions in 2 moles of H_2SO_4

1 mole of H_2SO_4 contains 2 moles of H^+

2 moles of H_2SO_4 contains (2x2)= 4 moles of H^+

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

4 moles of H^+ contains $(4x6.02\times10^{23})$= $24.08x10^{23}$ H^+ ions

d) Total number of ions in 1 mole of $(NH_4)_2SO_4$

$(NH_4)_2SO_4(aq) \longrightarrow 2NH_4^+(aq) + SO_4^{2-}(aq)$

1 mole of $(NH_4)_2SO_4$ contains 3 moles of ions

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

3 moles of ions contain $(3x6.02\times10^{23})$

$= 18.06x10^{23}$ ions

2. Calculate the number of atoms in the following

a) 0.25 moles of calcium

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

0.25 moles of calcium contain $(0.25x6.02x10^{23}/1)$ atoms

$=1.5x10^{23}$ atoms

b) 8 moles of sulphur

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

8 moles of sulphur contain $(8x6.02x10^{23}/1)$ atoms

$=4.8x10^{23}$ atoms

c) O.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 $(1x6.02x10^{23}/16)$ atoms

0.4g of oxygen contains $(0.4x1x6.02x10^{23}/16)$ atoms

$=1.5 \times 10^{24}$ atoms

Exercise

1. 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 $Al_2(SO_4)_3$
   d) Number of hydroxyl ions in 2 moles of $Ca(OH)_2$
2. Calculate the number of particles in the following
   a) 0.1 moles of sodium atoms
   b) 0.5 moles of chlorine atoms
   c) O.3 moles of calcium atoms
3. Use the value of 6.02×10^{23} mol^{-1} for the Avogadro constant to find the number of atoms in;

a) $2.0x\ 10^{-3}$g of calcium

b) 16g of magnesium ribbon

c) 27g of sodium carbonate decahydrate or hydrated sodium carbonate.
d) $5.0x10^{-6}$g of argon
e) $1.00x10^{-10}$g of mercury

f) How many P_4 molecules and atoms are present in 30g of phosphorus? (P=31)

g) Sucrose is a compound commonly known as sugar. Its molecular formular is $C_{12}H_{22}O_{11}$. Calculate the number of atoms of each element in 4.32g of sucrose. (Hint; get the RMM first)

Converting number of particles to masses

Example

1. Calculate the mass of sodium with $1.5x10^{22}$ 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 ($1x23/6.02x10^{23}$)g of sodium
   1.5×10^{22} atoms are contained in ($1.5x10^{22}\ x1x23/6.02x10^{23}$)g of sodium
   =0.575g of sodium

2. How many grams of calcium contain?

a) $6.02x10^{23}$ 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 ($1x40/6.02x10^{23}$)g of calcium
   6.02×10^{23} atoms are contained in ($6.02x10^{23}\ x1x40/6.02x10^{23}$)g of calcium
   =40g of calcium

b) $1.5x10^{23}$ 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 ($1x40/6.02x10^{23}$)g of calcium

1.5×10^{23} atoms are contained in ($1.5x10^{23}\ x1x40/6.02x10^{23}$)g of calcium
=10g of calcium

c) $3.0x10^{23}$ 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 ($1x40/6.02x10^{23}$)g of calcium
3.0×10^{23} atoms are contained in ($3.0x10^{23}\ x1x40/6.02x10^{23}$)g of calcium
=20g of calcium

Self-test

1. convert $2.03\text{x}10^{23}$ particles of sodium into mass. (Na=23)
2. $6\text{x}10^{18}$ particles of magnesium
3. $3\ \text{x}10^{13}$ particles of lead(ii) nitrate
4. $2.4\ \text{x}10^{20}$ particles of glucose
5. $5.08\ \text{x}10^{3}$ 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 H_2O= 18g; 1 mole of $CaCO_3$= 100g; 1 mole of H_2SO_4= 98g.

Examples

1. Calculate the number of moles of in;

a) 4g of oxygen molecule (O_2)
(0=16, S=32)

RMM= (2X16) =32
32g is contained in 1 mole of O_2
1 g is contained in ($^1/_{32}$) moles of O_2
4g is contained in ($^{4x1x1}/_{32}$) moles of O_2
=0.125 moles of O_2

b) 160g of sulphur (S)

RAM= 32
32g is contained in 1 mole of S
1 g is contained in ($^1/_{32}$) moles of S
160g is contained in ($^{160x1x1}/_{32}$) moles of S
=5.0 moles of S

2. Calculate the number of moles in the following molecules

a) 2 g of Calcium oxide

RMM of CaO= 40+16 =56
56g is contained in 1 mole of CaO
1 g is contained in ($^1/_{56}$) moles of CaO
2g is contained in ($^{2x1}/_{56}$) moles of CaO
=0.036 moles of CaO

b) 4 g of sodium hydroxide

RMM of NaOH= 23+16+1 =40
40g is contained in 1 mole of NaOH
1 g is contained in ($^1/_{40}$) moles of NaOH
2g is contained in ($^{2x1}/_{40}$) moles of NaOH
=0.1 moles of NaOH

Summary

Number of moles = $\frac{given\ mass}{RAM}$, for atoms

Number of moles= $\frac{given\ mass}{RMM}$, for molecules

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 carbondioxide
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

0.23 moles of sodium weighs ($^{0.23x23}/_{1}$)g

=5.29g

b) 7.1 moles of chlorine molecules

RMM of Cl_2

=35.5x2

=71

1 mole of chlorine weighs 71g

7.1 moles of chlorine weighs($^{7.1x71}/_{1}$)g

=504.1g of chlorine

c) 0.1 moles of potassium carbonate

RMM of K_2CO_3
=(39X2) +12+(16X3)
=138
1 mole of potassium carbonate weighs 138g
0.1 moles of chlorine weighs($^{0.1x138}/_{1}$)g
=13.8g of chlorine

Exercise

Calculate the mass of

a) 0.1 moles of sodium atom
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 on percentage composition by mass

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

Percentage composition of an element $= \frac{Mass\ of\ an\ element}{Formula\ mass}$ x 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 $CaCO_3$ =40+12+(3x16) =100g
Mass of oxygen= (3x16) =48g

Percentage of oxygen= $\frac{48}{100}$ x100
=48%

b) Water in $CuSO_4.5H_2O$

Formula mass of $CuSO_4.5H_2O$ =64+32+(4x16) +5(1x2+16) =250g
Mass of water= 5(1x2+16) =90g

Percentage of oxygen= $\frac{90}{250}$ x100
=36%

c) Nitrogen in nitrogen dioxide

Formula mass of NO_2 = 14+(2x16) =46g
Mass of nitrogen= 14g

Percentage of oxygen= $\frac{14}{46}$ x100
=30.4%

d) Oxygen in baking powder, $NaHCO_3$

Formula mass of $NaHCO_3$ =23+1+12+(3x16) =84g
Mass of oxygen= (3x16) =48g

Percentage of oxygen= $\frac{48}{84}$ x100
=57.14%

Exercise

1. Calculate the percentage of nitrogen in each of the following

a) Ammonium chloride (NH_4Cl)
b) Sodium nitrate ($NaNO_3$)
c) Ammonium sulphate, $(NH_4)_2SO_4$

(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, $Na_2CO_3.10H_2O$
3. A metal sulphate, $X_2(SO_4)_3$ contains 28% by mass of metal X. Determine the Relative Atomic Mass of X and the Relative Molecular Mass of $X_2(SO_4)_3$

(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.

Example of empirical and molecular formulae

Compound	**Molecular formula**	**Empirical formula**
Water	H_2O	H_2O
Ammonia	NH_3	NH_3
Ethene	C_2H_4	CH_2
Benzene	C_6H_6	CH
Glucose	$C_6H_{12}O_6$	CH_2O

Steps in calculating empirical formulae

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

   Number of moles = $\frac{Mass}{RAM/RMM}$
4. Find the mole ratio of the elements by dividing the moles with the smallest number
5. Write down the empirical number.

If the mole ratio is in fractions;

1. Round off to the nearest whole number if it is very close to the whole number.

2. 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

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 n is number to be determined

Calculations on empirical and molecular formulae

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 present	C	H
Percentage composition	80	20
Number of moles	$^{80}/_{12}$	$^{20}/_{1}$
	6.7	20
Divide by the smallest	$^{6.7}/_{6.7}$	$^{20}/_{6.7}$
Mole ratio	1	3

The empirical formula is CH_3

b) (Empirical formula)n= Molar mass

$(CH_3)n=30$

$(12+3)n=30$

$15n=30$

$n=2$

The molecular formula is therefore $(CH_3)_2 = C_2H_6$

2. 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	$^{28}/_{56}$	$^{24}/_{32}$	$^{48}/_{16}$
	0.5	0.75	3
Divide by the smallest	$^{0.5}/_{0.5}$	$^{0.75}/_{0.5}$	$^{3}/_{0.5}$
Mole ratio	1	1.5	6
	2x (1	1.5	6)
	2	3	12

The empirical formula is $Fe_2S_3O_{12}$ or $Fe_2(SO_4)_3$

3. 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	$^{85.7}/_{12}$	$^{14.3}/_{1}$

	7.14	14.3
Divide by the smallest	$^{7.14}/_{7.14}$	$^{14.3}/_{7.14}$
Mole ratio	1	2

The empirical formula is CH_2

(Empirical formula)n= Molar mass

$(CH_2)n=28$

$(12+2)n=28$

$14n=28$

$n=2$

The molecular formula is therefore $(CH_2)_2 = C_2H_4$

4. Calculate the empirical formula of a compound that contains 52g of zinc, 9.6g of carbon and 38.4g of oxygen. (Zn=65,C=12,O=16)

Elements present	Zn	C	O
Composition by mass	52	9.6	38.4
Number of moles	$^{52}/_{65}$	$^{9.6}/_{12}$	$^{38.4}/_{16}$
	0.8	0.8	2.4
Divide by the smallest	$^{0.8}/_{0.8}$	$^{0.8}/_{0.8}$	$^{2.4}/_{0.8}$
Mole ratio	1	1	3

The empirical formula is $ZnCO_3$

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**)
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 $CuSO_4.5H_2O$**)
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 $Na_2C_2O_3.10H_2O$**)
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=$FeCl_3$**)

8. 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=$Na_2S_2O_3$**)
d) 32.4%Na, 22.5%S, 45.0%O (**Answer=Na_2SO_4**)

Calculation 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

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

   $Fe(s) + S(s) \longrightarrow FeS(s)$

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

   32g of S forms 88g of FeS

   1 g of S forms ($^{1x88}/_{32}$)g of FeS

   64g of S forms ($^{64x1x88}/_{32}$)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

   $2Mg(s) + O_2(s) \longrightarrow 2MgO(s)$

   2 moles of magnesium oxide is formed 2 mole of magnesium

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

   80g of MgO is formed by 48g of Mg

   1 g of MgO is formed by ($^{1x48}/_{80}$)g of Mg

   55g of MgO is formed by ($^{55x1x48}/_{80}$)g of Mg

   =33g 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

   a) $CuO(s) + H_2(g) \longrightarrow Cu(s) + H_2O(l)$

   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 ($^{1x64}/_{80}$)g of Cu

3.2g of CuO forms ($^{3.2x1x64}/_{80}$)g of Cu

=2.6g of Cu

b) 1 mole of CuO forms 1 mole of H_2O

(64+16)g of CuO forms (1x2+16)g of H_2O

80g of CuO forms 18g of H_2O

1 g of CuO forms ($^{1x18}/_{80}$)g of H_2O

3.2g of CuO forms ($^{3.2x1x18}/_{80}$)g of H_2O

=0.7g of H_2O

c) 1 mole of CuO reacts with 1 mole of H_2

(64+16)g of CuO reacts with (1x2)g of H_2

80g of CuO reacts with 2g of H_2

1 g of CuO reacts with ($^{1x2}/_{80}$)g of H_2

3.2g of CuO reacts with ($^{3.2x1x2}/_{80}$)g of H_2

=0.08g 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, 0=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

1. $2H_2(g) + O_2(g) \longrightarrow 2H_2O(g)$
   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

2. $H_2(g) + Cl_2(g) \longrightarrow 2HCl(g)$
   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 vise versa.

The relationship between vapor density and Relative Molecular Mass can be deduced from Avogadro's law and is expressed as
2 x Vapor Density= 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(stp)	22.4l or 22.4dm^3 or 22400cm^3
Room temperature and pressure (rtp)	24l or 24dm^3 or 24000cm^3

Example
1. Determine the number of moles in the following gaseous volumes at stp
   a) 1.2 dm^3 of nitrogen

b) $300cm^3$ of ammonia

Solution

a) $22.4\ dm^3$ is occupied by 1 mole of nitrogen

$1\ dm^3$ is occupied by $(^{1}/_{22.4})$ moles of nitrogen

$1.2\ dm^3$ is occupied by $(^{1.2x1}/_{22.4})$ moles of nitrogen

= 0.05 moles of nitrogen at stp

b) $22400\ cm^3$ is occupied by 1 mole of ammonia

$1\ cm^3$ is occupied by $(^{1}/_{22400})$ moles of ammonia

$300\ cm^3$ is occupied by $(^{300x1}/_{22400})$ moles of ammonia

= 0.013 moles of ammonia at stp

2. Determine at rtp the volume and mass of

a) 0.04 moles of hydrogen

b) 0.2 moles of carbondioxide

Solution

a) 1 mole of hydrogen occupies 24*l* at rtp

0.04 moles of hydrogen occupies $(^{0.04x24}/_{1})$ l at rtp

=0.96 l of hydrogen at rtp

RMM for H_2 =(1x2)= 2

1 mole of hydrogen weighs 2g

0.04 moles of hydrogen weighs $(^{0.04x2}/_{1})$g

=0.08g of hyrogen

b) 1 mole of carbondioxide occupies 24l at rtp

0.2 moles of carbondioxide occupies $(^{0.2x24}/_{1})$ l at rtp

=4.8 l of carbondioxide at rtp

RMM for CO_2 =12+(16x2)= 44

1 mole of carbondioxide weighs 44g

0.2 moles of carbondioxide weighs $(^{0.2x44}/_{1})$g

=8.8g of carbondioxide

3. Calculate the RMM of Y given that 0.8 g of Y occupies 560cm^3 at stp.

Solution

560 cm^3 of Y weighs 0.8g

1 cm^3 of Y weighs ($^{0.8}/_{560}$) moles of nitrogen

22400 cm^3 of Y weighs ($^{22400x1x0.8}/_{560}$) moles of nitrogen

=32g

The RMM of Y is 32.

Calculation of masses and volumes

Examples

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

$$2NH_3(g) + 3CuO(s) \longrightarrow 3Cu(s) + N_2(g) + 3H_2O(l)$$

Calculate

a) The number of moles of copper(II)oxide used

b) The mass of copper formed

c) The volume of ammonia used at stp

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

Solution

1.

a) From the equation,

2 moles of NH_3 reacts with 3 moles of CuO

1 mole of NH_3 reacts with ($^{3}/_{2}$) moles of CuO

0.2 moles of NH_3 reacts with ($^{3x0.2}/_{2}$) moles of CuO

=0.3 moles of CuO

b) From the equation

2 moles of NH_3 produce 3 moles of Cu

2 moles of NH_3 produce (3x64)g of Cu

1 mole of NH_3 produces ($^{192}/_{2}$) moles of Cu

0.2 moles of NH_3 produces ($^{192x0.2}/_{2}$) moles of Cu

=19.2g of Cu

c) 1 mole of NH_3 occupies 22.4l at stp

0.2 moles of NH_3 occupies ($^{22.4x0.2}/_{1}$)l at stp

=4.48 l of NH_3 at stp

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 stp

   (Cu=64,O=16,C=12, 1 mole of a gas occupies 24 dm^3 at stp)

Solution

$CuCO_3(s) \longrightarrow CuO(s) + CO_2(g)$

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 ($^{80}/_{124}$) g of CuO

2.5g of $CuCO_3$ produces ($^{2.5x80}/_{124}$) g of CuO

=1.61g 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 24dm^3 of CO_2

124g of $CuCO_3$ produces 24dm^3 of CO_2

1 g of $CuCO_3$ produces ($^{24}/_{124}$) dm^3 of CO_2

2.5 g of $CuCO_3$ produces ($^{2.5x24}/_{124}$) dm^3 of CO_2

=0.484 dm^3 of CO_2

3. From the equation

$2Fe(s) + 3Cl_2(g) \longrightarrow 2FeCl_3(s)$

Calculate the

   a) Volume of chlorine at stp required to react with 8 g of iron

   b) Mass of iron(III)chloride formed

Solution

a) From the equation

2 moles of Fe react with 3 moles of Cl_2

(2x56)g of Fe react with (3x22.4)dm^3 of Cl_2

112g of Fe react with 67.2dm^3 of Cl_2

1 g of Fe react with($^{67.2}/_{112}$)dm^3 of Cl_2

8 g of Fe react with($^{67.2x8}/_{112}$)dm^3 of Cl_2

=4.8 dm^3 of Cl_2 at stp

b) From the equation

2 mole of Fe produce 2 mole of $FeCl_2$

(2x56)g of Fe produces 2x(56+35.5x2)g of $FeCl_2$

112g of Fe produces 325g of $FeCl_2$

1 g of Fe produces ($^{325}/_{112}$) g of $FeCl_2$

8 g of Fe produces ($^{8x325}/_{112}$) g of $FeCl_2$

=23.01g of $FeCl_2$

Exercise

1. What volume of hydrogen at stp would be used if 40g of nitrogen combined with it to form ammonia.
2. 0.1 mole of lead nitrate is completely decomposed on heating according to the reaction

   $$2Pb(NO_3)_2(s) \longrightarrow 2PbO(s) + 4NO_2(g) + O_2(g)$$

   a) What volume at stp of nitrogen dioxide and oxygen were evolved
   b) What was the mass of the residue left
3. 1.4 litres of ammonia gas measured at stp where passed over hot copper(II)oxide. The equation for the reaction is

   $$3CuO(s) + 2NH_3(g) \longrightarrow 3Cu(s) + N_2(s) + 3H_2O(l)$$

   Calculate
   a) The number of moles of ammonia used
   b) The number of moles of nitrogen gas used
   c) The mass of copper formed
4. 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 stp needed to reduce the oxide
   c) Calculate the volume of carbondioxide at stp produced when the carbonate was completely decomposed.
5. Calculate the mass of sulphur deposited when 8.4 dm3 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

$A(aq) + B(aq) \longrightarrow C(aq) + D(aq)$

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

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 $1dm^3$ e.t.c.

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

A Molar solution

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/dm^3 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)=1cubic decimetre ($1dm^3$) = 1000 cubic centimetre ($1000cm^3$)

Calculations on molarity and masses

Examples

1. 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 1dm^3 of solution
b) 4.0g of sodium hydroxide in 400cm^3 of solution
c) 53g of anhydrous sodium carbonate in 2 dm^3 of solution

Solution

a) RMM of $CuCl_2$
=64+(35.5x2)
=135
135g is contained in 1 mole of $CuCl_2$
1 g is contained in ($^1/_{135}$) moles of $CuCl_2$
13.5g is contained in ($^{13.5x1}/_{135}$) moles of $CuCl_2$
=0.1M of $CuCl_2$

b) 400cm^3 of solution contains 4.0g of NaOH
1cm^3 of solution contains ($^{4.0}/_{400}$)g of NaOH
1000cm^3 of solution contains ($^{1000x4.0}/_{400}$)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 ($^1/_{40}$) moles of NaOH
10 g is contained in ($^{10x1}/_{40}$) moles of NaOH
=0.25M NaOH

c) 2dm^3 of solution contains 53g of Na_2CO_3
1dm^3 of solution contains ($^{53}/_2$)g of Na_2CO_3
=26.5g/dm^3 of Na_2CO_3 (concentration in grams/litre)

RMM of Na_2CO_3
=23x2+12+16x3
=106

106g is contained in 1 mole of Na_2CO_3
1 g is contained in ($^1/_{106}$) moles of Na_2CO_3
10 g is contained in ($^{10x1}/_{106}$) moles of Na_2CO_3

$=0.25M\ Na_2CO_3$

In general, *Molarity* $= \frac{concentration\ (g/l)}{molar\ mass}$

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

N.B. 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) 25cm^3 of 0.1M potassium carbonate solution
   d) 500cm^3 of 0.05M sodium carbonate solution

Solution

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

0.1dm^3 of solution contains ($^{0.1x2}/_{1}$) moles of Na_2SO_4

=0.2 moles of Na_2SO_4

RMM for Na_2SO_4

=23x2+32+16x4

=142

1 mole of Na_2SO_4 weighs 142g

0.2 moles of Na_2SO_4 weighs ($^{0.2x142}/_{1}$))g

=28.4g

b) 1*l* of solution contains 0.25 moles of NaOH

RMM for NaOH

=23+16+1

=40

1 mole of NaOH weighs 40g

0.25 moles of NaOH weighs ($^{0.25x40}/_{1}$))g

=10g

c) 1000cm^3 of solution contains 0.1moles of K_2CO_3

1 cm^3 of solution contains ($^{0.1}/_{1000}$) moles of K_2CO_3

25 cm^3 of solution contains ($^{25x0.1}/_{1000}$) moles of K_2CO_3

=0.0025 moles of K_2CO_3

RMM of K_2CO_3

=39x2+12+16x3

=138

1 mole of K_2CO_3 weighs 138g

0.0025 moles of K_2CO_3 weighs ($^{0.0025x138}/_{1}$))g

=0.345g

Calculating number of moles of ions in standard solutions

Examples

1. Calculate the number of moles of hydrogen ions in 25cm^3 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. 1000cm^3 of solution contains 0.2moles of H_2SO_4

1 cm^3 of solution contains ($0.2/1000$) moles of H_2SO_4

25 cm^3 of solution contains ($^{25x0.2}/_{1000}$) moles of H_2SO_4

=0.005 moles of H_2SO_4

From the equation of ionization of H_2SO_4

$$H_2SO_4(aq) \longrightarrow 2H^+(aq) + SO_4^{-2}(aq)$$

1 mole of H_2SO_4 produces 2 moles H^+

0.005 moles of H_2SO_4 produces ($^{0.005x2}/_{1}$) moles H^+

=0.01

moles H^+

2. 1000cm^3 of solution contains 0.12 moles of K_2CO_3

1 cm^3 of solution contains ($^{0.12}/_{1000}$) moles of K_2CO_3

35 cm^3 of solution contains ($^{35x0.12}/_{1000}$) moles of K_2CO_3

=0.0042 moles of K_2CO_3

From the equation of ionization of K_2CO_3

$K_2CO_3(aq) \longrightarrow 2K^+(aq) + CO_3^{-2}(aq)$

1 mole of K_2CO_3 produces 2 moles K^+

0.0042 moles of K_2CO_3 produces $\left({}^{0.0042x2}/_{1}\right)$ moles K^+

=0.0084moles K^+

ACID-BASE REACTION (Titration)

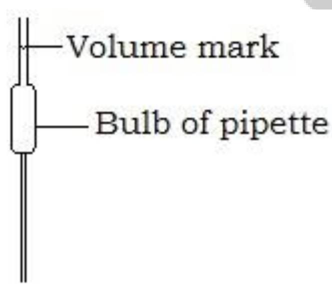
This is a method of volumetric analysis in which a solution (usually a standard solution) is added from a burette to another solution (usually whose concentration is unknown) until the reaction is complete.

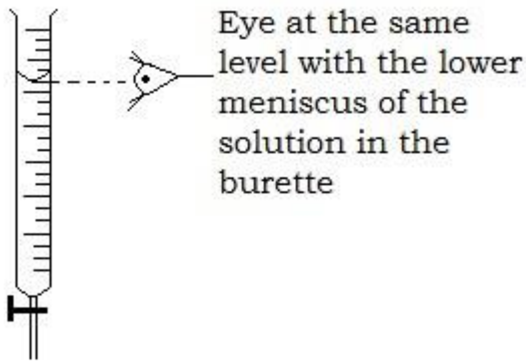
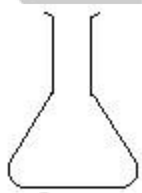
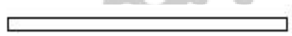
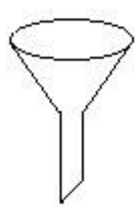
An indicator changes color immediately the reaction is complete or when the **end point** is reached. Most titrations at this level are acid-base titrations.

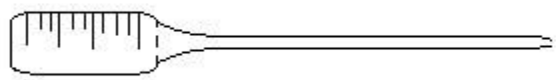
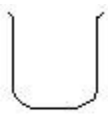
Common laboratory indicators and color changes

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

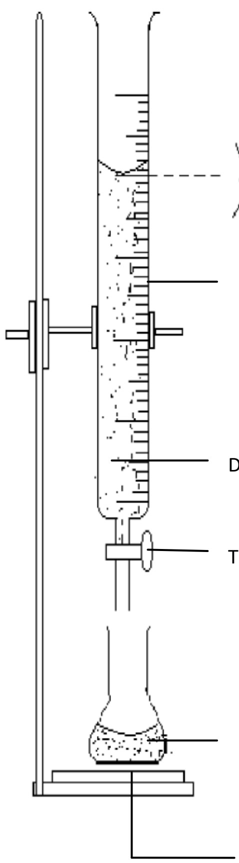
Apparatus commonly used in titration

Apparatus	**Use**
Pipette 	This is used for measuring and transferring fixed volumes of solutions. The volume of solution (usually whose concentration is not known) is measured accurately and transferred into a conical flask using a pipette. The solution transferred by the pipette is the **aliquot**. The common pipettes used are of volumes $20.0cm^3$ and $25.0cm^3$ which are usually written on them. The reading of a pipette is recorded to one decimal place e.g. 20.0 cm^3.

 Burette	This is used to transfer accurately a precise volume of solution. The **titrant** is filled into a burette using a filter funnel to prevent wastage of solutions. The reading of a burette is recorded to two decimal places e.g. $0.00cm^3$, $18.40cm^3$ e.t.c. For accurate reading of volumes of liquid in burettes and pipettes, the eye must be at the same level with the meniscus of the liquid as illustrated besides.
Conical flask 	It is a container in which the reaction between the acid and the alkali take place. The titrant is run from a burette into a conical flask containing a known volume of the solution (usually whose concentration is not known).
 White tile	It enables one to observe indicator color changes more clearly. A piece of white paper can be used in place of a white tile.
Filter funnel	Enables the solution to be poured directly into the burette without wastage
	

 Dropper	For addition of the indicator to the solution in the conical flask
Beaker 	For holding and pouring out solutions

Procedure for acid-base titration



1. Wash the pipette with distilled water then a little of the solution it is to measure. Use the pipette to deliver either $20.0cm^3$ or $25.0cm^3$ of the alkali into a clean conical flask. Add a few drops (2 or 3 drops) of indicator. 2. Wash the burette with distilled water then a little of the acid solution and run out the acid through the tap. Fill the burette above the 0 cm^3 mark and run a little of the acid out to bring the meniscus of the acid to the 0 cm^3 mark or slightly below it. Take the burette reading as V_1 cm^3.

3. Arrange the apparatus as shown on the left hand side. 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.

4. Take the burette reading again V_2 cm^3. Subtract (V_2- V_1) cm^3 to get the ―**titre**‖ (i.e. the volume of the acid needed to neutralize the known volume of alkali).

5. Repeat the titration. Obtain an average titre. From this volume you can calculate the unknown concentration.

N.B 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 =25.0cm^3

Number of titration	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 cm^3 and 15.00 cm^3

Average volume of acid used= $\frac{15.00+15.00}{2}$ =15.00 cm^3

Calculate

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

Method of calculation

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

   $HCl(aq) + NaOH(aq) \longrightarrow NaCl(aq) + H_2O(l)$

   1 mole of HCl neutralizes 1 mole of NaOH

2. 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).

a) Moles of NaOH that reacted

1000cm^3 of solution contains 0.1moles of NaOH

1 cm^3 of solution contains ($0.1/1000$) moles of NaOH

25 cm^3 of solution contains ($25x0.1/1000$) moles of NaOH

=0.0025 moles of NaOH

3. 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.

b) From the equation of reaction

1 mole of NaOH reacts with 1 mole of HCl

0.0025 moles of NaOH reacts with ($^{0.0025}/_{1}$) moles of HCl

=0.0025 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)

4. Now work out the molarity of the acid

c) 15cm^3 of solution contains 0.0025moles of HCl

1cm^3 of solution contains ($^{0.0025}/_{15}$) moles of HCl

1000cm^3 of solution contains ($^{1000x0.0025}/_{15}$) moles of HCl

=0.167 mol/*l* of HCl

The molarity of the HCl is 0.167M

5. 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 moles of HCl weighs ($^{0.167x36.5}/_{1}$)g

=6.1g/*l*

Other examples

1. 25cm^3 of sulphuric acid of concentration 0.15mol/dm^3 neutralized 31.2cm^3 of potassium hydroxide solution. Find the concentration of the KOH solution in mol/l and in grams/litre.

Solution

Write the equation

Equation for the reaction

H_2SO_4 (aq) + 2KOH(aq) $\longrightarrow$ K_2SO_4(aq) + H_2O(l)

1 mole of acid reacts with 2 moles of alkali

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

1000cm^3 of solution contains 0.15moles of H_2SO_4

1 cm^3 of solution contains ($^{0.15}/_{1000}$) moles of H_2SO_4

25 cm^3 of solution contains ($^{25x0.15}/_{1000}$)moles of H_2SO_4

=0.00375 moles of H_2SO_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 H_2SO_4 reacts with 2 moles of KOH

0.00375moles of H_2SO_4 reacts with ($0.00375^{x2}/_1$) moles of KOH

=0.0075 moles of KOH

Calculate the molarity (concentration in mol/dm^3) of the alkali

31.2cm^3 of solution contains 0.0075moles of KOH

1cm^3 of solution contains ($^{0.0075}/_{31.2}$) moles of KOH

1000cm^3 of solution contains ($^{1000x0.0075}/_{31.2}$) moles of KOH

=0.24 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 ($^{0.24x56}/_1$)g

=13.44g

The concentration of the KOH is 13.44g/*l*

2. Determine the molarity of acids in the following solutions.
   a) 16.0cm^3 of 1.5M sodium hydroxide neutralized by 20.0cm^3 of hydrochloric acid
   b) 25.0cm^3 of 0.2M ammonia solution neutralized by 20.0cm^3 of nitric acid

Solution

a) Equation of reaction

$HCl(aq) + NaOH(aq) \longrightarrow NaCl(aq) + H_2O(aq)$

1000cm^3 of solution contains 1.5moles of NaOH

1 cm^3 of solution contains ($^{1.5}/_{1000}$) moles of NaOH

16 cm^3 of solution contains ($^{16x1.5}/_{1000}$)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 ($^{0.024x1}/_{1}$)moles of HCl

=0.024 moles of HCl

20.0cm^3 of solution contains 0.024moles of HCl

1cm^3 of solution contains ($^{0.024}/_{20}$) moles of HCl

1000cm^3 of solution contains ($^{1000x0.024}/_{20}$) moles of HCl

=1.2 M HCl

The concentration of the HCl is 1.2M

Equation for reaction

b) $HNO_3(aq) + NH_4OH(aq) \longrightarrow NH_4NO_3(aq) + H_2O(aq)$

1000cm^3 of solution contains 0.2moles of NH_4OH

1 cm^3 of solution contains ($0.2/1000$) moles of NH_4OH

25cm^{3}of solution contains ($^{25x0.2}/_{1000}$)moles of NH_4OH

=0.005 moles of NH_4OH

From the equation

1 mole of NH_4OH reacts with 1 mole of HNO_3

0.005moles of NH_4OH reacts with ($^{0.005x1}/_{1}$)moles of HNO_3

=0.005 moles of HNO_3

20.0cm^3 of solution contains 0.005moles of HNO_3

1cm^3 of solution contains ($^{0.005}/_{20}$) moles of HNO_3

1000cm^3 of solution contains($^{1000x0.005}/_{20}$)moles of HNO_3

=0.25 M HNO_3

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)

End of Chapter Questions

1. Convert the following masses into moles and calculate the number of particles in those masses. (N=14, O=16, Cu=64, Zn=65).
   (a) 12.8 g of copper
   (b) 3.22 g of zinc nitrate.
2. Convert the following into mass. (N=14, O=16, Al=27, S=32, Pb=207).
   (a) 0.01 moles of lead(II) nitrate.
   (b) 0.04 moles of aluminium sulphate.
3. Calculate the;
   (a) mass of residue and
   (b) volume of gas evolved at STP. When 49.6 g of copper(II) carbonate were heated strongly. (C=12. O=16 and one mole of a gas occupies 22.4 dm^3 at STP)

4. Convert the following volumes into moles at standard conditions.
   (a) 1200 cm^3 of carbon dioxide at RTP.
   (b) 1.12 dm^3 of oxygen at STP.
5. Chlorine displaces bromine from sodium bromide solution.
   (a) Name the type of reaction.
   (b) Write the observation for the reaction.
   (c) Calculate the volume of bromine produced at room temperature when chlorine is bubbled through 20 cm^3 of a 0.5 M sodium bromide solution is: [one mole of a gas occupies 24 dm^3 at room temperature].
6. 18 cm^3 of 0.2 M sulphuric acid reacted with 25 cm^3 of sodium hydroxide.
   (a) Identify the standard solution
   (b) Write equation for the reaction
   (c) Deduce the mole ratio
   (d) Calculate
      (i) moles of sulphuric acid used.
      (ii) moles of sodium hydroxide used
      (iii) Molarity of sodium hydroxide
      (iv) concentration in *g/l* of sodium hydroxide.
7. 25.0 cm^3 of 0.1 M dilute acid H_nX reacted with 25.0 cm^3 of 0.2 M sodium hydroxide solution.
   (a) Calculate
      (i) the number of moles of sodium hydroxide solution which reacted.
      (ii) the number of moles of the acid that reacted.
   (b) Determine the value of n. Hence, the basicity of the acid.
   (c) Write the equation for the reaction.