

Second term Chemistry E-Lesson Note

SUBJECT: CHEMISTRY

CLASS: SS1

SCHEME OF WORK

WEEKS and TOPICS

- 1. Revision of last terms work/introduction to mole concept:** calculation of molar mass/relative molecular mass, calculation in term of formulae.
- 2. Mole concept cont'd:** Calculations in term of relative molecular mass and volume, percentage of an element in a compound, empirical formulae and molecular formulae
- 3. Writing and balancing of chemical equation,** calculation from chemical equation.
- 4/5. State, illustration and verification of the following chemical laws:** law of conservation of matter, law of constant composition or definite proportions, law of multiple proportion, calculations base on chemical laws
- 6. Chemical combinations or bonding:** types of bonds, strong bond: electrovalent or ionic bonding, covalent bonding, co-ordinate or dative bonding, metallic bonding, weak bond: hydrogen bonding and vander-waal force.
- 7. The kinetic theory of matter:** State of matter, postulate of kinetic theory of matter, applications of kinetic theory of gases, and phenomena to illustrate kinetic theory of matter, change of state.
- 8 Gas laws and calculations:** Boyle's laws, Charles' law, general gas law and ideal gas law
- 9. Gas law cont'd:** Gay- Lucca's Law and Avogadro's Law, Graham's law of diffusion, Molar volume of gases- Avogadro number and the mole, concept, Dalton's law of partial pressure
- 10. Experiment to illustrate gas laws**
- 11. Air:** constituent, percentage composition and properties. **Flame:** drawing, label and explain various zone of flame e.g hydrogen flame, bursen flame and candle flame.

REFERENCE TEXTS:

- 1 .Comprehensive Certificate Chemistry for Senior Secondary Schools by G N C Ohia et al

Second term Chemistry E-Lesson Note

2. New School Chemistry for Senior Secondary Schools by Osei Yaw Ababio
3. Chemistry for Senior Secondary Schools 1 by Magbagbeola O, et al; Melrose Books and Publishers
4. Revised edition understanding chemistry for schools and colleges by Godwin O. Ojokuku.

WEEK 1:

MOLE: The mole is the amount of a substance that contains the same number of particles as the number of atoms in exactly 12g of carbon-12. The elementary particles can be atoms, molecules, ions, electrons or any other. The number of atoms in a mole of any element is 6.02×10^{23} . It is called Avogadro's number. Hence, one mole of any substance contains 6.02×10^{23} elementary particle of that substance. Thus,

1 mole of an element contains 6.02×10^{23} atoms of the element

1 mole of a compound or molecule contains 6.02×10^{23} molecules

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

The relative atomic mass of an element = mass of 1 mole of the element

The relative molecular mass of a compound or molecule = mass of 1 mole of the compound or molecule

The number of ions present in an ionic compound is obtained by summing up the amount of individual ions present in the compound.

By definition:

$$\text{Mole} = \frac{\text{Reacting mass}}{\text{Molar mass}}$$

$$\text{Mole} = \frac{\text{Number of particles}}{6.02 \times 10^{23}}$$

$$\therefore \text{Reacting mass} = \frac{\text{Number of particles}}{\text{Molar mass} \times 6.02 \times 10^{23}}$$

Examples

1. What is the mass of 3 moles of oxygen gas? (O=16)
2. How many atoms are present in 5g sample of sodium (Na=23, N= 6.02×10^{23} .)
3. How many ions are there in 1 mole of calcium chloride?

Solution

1. Mass of 1 mole of $O_2 = (2 \times 16) = 32\text{g}$

Mass of 3 mole of $O_2 = (3 \times 32) = 96\text{g}$

$$2. \text{Mole} = \frac{\text{Reacting mass}}{\text{Molar mass}} = \frac{5}{23} = 0.217$$

But number of moles = number of particles (atoms)

Avogadro's number

$\therefore 0.217 = \text{number of atoms}$

6.02×10^{23} .

Number of atoms = $0.217 \times 6.02 \times 10^{23} = 1.3 \times 10^{23}$ atoms

3. 1 mole of $CaCl_2$ contains: 1 mole of Ca^{2+} and 2 moles of Cl^-

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

\therefore 1 mole of $CaCl_2$ contains $(1 \times 6.02 \times 10^{23}) + (2 \times 6.02 \times 10^{23})$ ions = 18.06×10^{23} ions

Molar mass: The molar mass of any substance is the mass of one mole of the substance expressed in grams. Its unit is g/mol. The molar mass of an element is its atomic mass expressed in grams and the molar mass of a compound is its relative molecular mass expressed in grams.

FORMULA MASS AND PERCENTAGE COMPOSITION

Second term Chemistry E-Lesson Note

Example: calculate the percentage by mass of each element in calcium hydroxide, $\text{Ca}(\text{OH})_2$
[Ca=40, O=16, H=1]

Solution

Rmm or formula mass of $\text{Ca}(\text{OH})_2 = 40 + (16 \times 2) + (1 \times 2) = 40 + 32 + 2 = 74$

%by mass of Ca in $\text{Ca}(\text{OH})_2 = 100 \times 40/74 = 54.1\%$

% by mass of O in $\text{Ca}(\text{OH})_2 = 100 \times 32/74 = 43.2\%$

%by mass of H in $\text{Ca}(\text{OH})_2 = 100 \times 2/74 = 2.7\%$

Assignment

1. What is the percentage by mass of sodium in sodium hydroxide? (Na = 23, O = 16, H=1)
2. How many atoms are there in one mole of hydrogen gas? [Avogadro's constant = $6.02 \times 10^{23} \text{mol}^{-1}$]

Tutorial questions

1. How many elementary particles are there in 4g of CO_2 ?
2. What is the relative molecular mass of MgSO_4 ? [Mg = 24, S = 32, O = 16]
3. A sample of chlorine contains 75% of $^{35}_{17}\text{Cl}$ and 25% of $^{37}_{17}\text{Cl}$. Calculate the relative molecular mass of Chlorine.
4. What is the percentage by mass of carbon in $\text{Ca}(\text{HCO}_3)_2$? [Ca = 40, H = 1, O = 16]

PERIOD 3: EMPIRICAL AND MOLECULAR FORMULAE

Empirical formula is the simplest formula which gives the ratio of the number of different atoms present in a compound. It does not give the exact number of each atom whereas the molecular formula gives the exact number of the atoms present in a molecule of a compound.

The molecular formula of a compound is a whole number multiple of its empirical formula.

Examples:

EMPIRICAL FORMULAR: is a formula that simply gives the relative number of atoms of each element present in a molecule. It is the simplest formula.

MOLECULAR FORMULA: is a formula that state the actual number of each kind of atom found in a molecule. A substance whose empirical formula is CH_2 have a molecular formula C_2H_4 , C_4H_8 and so on. The empirical formula of a compound can be found if the percentage composition and relative atomic masses of each elements in the compound are known. Also, the empirical formula can be related to the molar mass of a compound in order to obtain the molecular formular for the compound.

EXAMPLE1: An organic compound has the composition 55% of carbon, 9% hydrogen and 36% oxygen. Calculate the empirical formular for the organic compound. (C= 12, H= 1, O= 16)

NOTE: Always check that the addition of the percentage composition of all elements present in a compound equal to 100.

SOLUTION:

	CARBON	HYDROGEN	OXYGEN
Relative composition:	55	9	36
Divide by atomic mass:	$\frac{55}{12}$	$\frac{9}{1}$	$\frac{36}{16}$
	=4.58	= 9.00	= 2.25
Divide by the smallest:	$\frac{4.58}{2.25}$	$\frac{9}{2.25}$	$\frac{2.25}{2.25}$
	2	4	1

Hence the empirical formular = $\text{C}_2\text{H}_4\text{O}$.

Second term Chemistry E-Lesson Note

EXAMPLE 2: What is the empirical formula of an organic compound whose percentage composition is carbon, 52.2%, hydrogen, 13.1% and oxygen, 34.7% (C = 12, H = 1, O = 16).

SOLUTION:

	CARBON	HYDROGEN	OXYGEN
Relative composition:	52.2	13.1	34.7
Divide by atomic mass:	$\frac{52.2}{12}$	$\frac{13.1}{1}$	$\frac{34.7}{16}$
	=4.35	= 13.1	= 2.17
Divide by the smallest:	$\frac{4.35}{2.17}$	$\frac{13.1}{2.17}$	$\frac{2.17}{2.17}$
	= 2	= 6	= 1

Hence the empirical formula = C_2H_6O .

EXAMPLE 3: An organic compound contains carbon, 62.1%, hydrogen, 10.3% and oxygen, 27.6% by mass.

(i) Find the empirical formula of the compound

(ii) If the molar mass of the compound is 58.0g, find its molecular formula. (C = 12, H = 1, O = 16).

SOLUTION:

	CARBON	HYDROGEN	OXYGEN
Relative composition:	62.1	10.3	27.6
Divide by atomic mass:	$\frac{62.1}{12}$	$\frac{10.3}{1}$	$\frac{27.6}{16}$
	=5.1	= 10.3	= 1.8
Divide by the smallest:	$\frac{4.35}{1.8}$	$\frac{13.1}{1.8}$	$\frac{1.8}{1.8}$
	= 3	= 6	= 1

Hence the empirical formula = C_3H_6O .

(ii) To calculate the molecular formula, relate the empirical formula to the molar mass.

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

$$(C_3H_6O)n = 58$$

$$(3 \times 12 + 1 \times 6 + 16 \times 1)n = 58$$

$$(58)n = 58$$

$$n = 1$$

The molecular formula = $(C_3H_6O)n = C_3H_6O$.

EXAMPLE 4: A hydrocarbon contains 20.80% of hydrogen and has a relative molar mass of 30, what is the

(i) Empirical formula

(ii) molecular formula (C = 12, H = 1).

SOLUTION:

Hydrocarbon is known to contain carbon and hydrogen only. Since the percentage composition of all elements in a compound must be equal to 100. Therefore, the percentage composition of carbon which is the second element contained by a hydrocarbon equals 79.20%. i.e. 100 – percentage composition of hydrogen (20.80%).

	CARBON	HYDROGEN
Relative composition:	79.20	20.80
Divide by atomic mass:	$\frac{79.20}{12}$	$\frac{20.80}{1}$
	=6.60	= 20.80
Divide by the smaller:	$\frac{4.60}{6.60}$	$\frac{20.80}{6.60}$
	= 1	= 3

Hence the empirical formula = CH_3 .

(ii) To calculate the molecular formula, relate the empirical formula to the molar mass.

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

Second term Chemistry E-Lesson Note

$$\begin{aligned}(\text{CH}_3)_n &= 30 \\(12 + 1 \times 3)n &= 30 \\(15)n &= 30 \\n &= 2\end{aligned}$$

The molecular formula = $(\text{CH}_3)_n = \text{C}_2\text{H}_6$.

EXAMPLE 5: A carbohydrate contains 40% and hydrogen 6.72%, Calculate its empirical formula and the molecular formula, if the molar mass is 180 (C = 12, H = 1, O = 16).

SOLUTION:

Carbohydrate contains the elements carbon, hydrogen and oxygen, but from the question above oxygen is missing, hence the percentage composition of oxygen equals 100 – (percentage composition of carbon and hydrogen)

$$= 100 - (40 + 6.72)$$

$$= 100 - 46.72 = 53.30\%$$

	CARBON	HYDROGEN	OXYGEN
Relative composition:	40	6.72	53.3
Divide by atomic mass:	<u>40</u>	<u>6.72</u>	<u>53.3</u>
	12	1	16
	=3.33	= 6.72	= 3.33
Divide by the smallest:	<u>3.33</u>	<u>6.72</u>	<u>3.33</u>
	3.33	3.33	3.33
	= 1	= 2	= 1

Hence the empirical formula = CH_2O .

(ii) To calculate the molecular formula, relate the empirical formula to the molar mass.

(Empirical formula) n = Molar mass

$$(\text{CH}_2\text{O}) n = 180$$

$$(12 + 1 \times 2 + 16) n = 180$$

$$(30) n = 180$$

$$n = \frac{180}{30}$$

$$n = 6$$

The molecular formula = $(\text{CH}_2\text{O}) n = (\text{CH}_2\text{O})_6$
= $\text{C}_6\text{H}_{12}\text{O}_6$.

The molecular formula = $\text{C}_6\text{H}_{12}\text{O}_6$.

EXAMPLE 6: A hydrocarbon contains 92.40% of carbon. If the vapour density of the hydrocarbon is 39. Find

(i) Empirical formula

(ii) Molecular formula (C = 12, H = 1).

SOLUTION:

Hydrocarbon is known to contain carbon and hydrogen only. Since the percentage composition of all elements in a compound must be equal to 100. Therefore, the percentage composition of hydrogen which is the second element contained by a hydrocarbon equals 7.60%. i.e 100 – percentage composition of carbon (92.40%).

	CARBON	HYDROGEN
Relative composition:	92.40	7.60
Divide by atomic mass:	<u>92.40</u>	<u>7.60</u>
	12	1
	=7.70	= 7.60
Divide by the smaller:	<u>7.70</u>	<u>7.60</u>
	7.60	7.60
	= 1	= 1

Hence the empirical formula = CH.

(ii) To calculate the molecular formula, relate the empirical formula to the molar mass.

(Empirical formula) n = Molar mass

But molecular mass = 2 x vapour density

Second term Chemistry E-Lesson Note

$$\begin{aligned}(\text{CH}) n &= 2 \times 39 \\(12 + 1) n &= 78 \\(13) n &= 78 \\n &= 6\end{aligned}$$

The molecular formula = $(\text{CH}) n = \text{C}_6\text{H}_6$.

EXAMPLE 7: Calculate the empirical formula of an organic compound containing 81.8% carbon and 18.2% hydrogen (C = 12, H = 1).

SOLUTION:

	CARBON	HYDROGEN
Relative composition:	81.8	18.2
Divide by atomic mass:	$\frac{81.8}{12}$	$\frac{18.2}{1}$
	=6.82	= 18.2
Divide by the smaller :	$\frac{6.82}{6.82}$	$\frac{18.2}{6.82}$
	= 1	= 2.67

The carbon : hydrogen {C:H} ratio of 1:2.67 is too far from whole numbers and so the lowest multiple of this which gives a whole number ratio is the empirical formula, i.e 1:2.67, 2:5.34, **3:8.01**, 4:10.68 etc.

Hence the empirical formula = C_3H_8 .

EXAMPLE 8: What is the empirical formula of an oxide of phosphorous that contains 43.6% phosphorous and 56.4% oxygen (P = 31, O = 16).

SOLUTION:

	PHOSPHOROUS	OXYGEN
Relative composition:	43.6	56.4
Divide by atomic mass:	$\frac{43.6}{31}$	$\frac{56.4}{16}$
	=1.41	= 3.53
Divide by the smaller:	$\frac{1.41}{1.41}$	$\frac{3.53}{1.41}$
	= 1	= 2.5

The phosphorous : oxygen {P:O} ratio of 1:2.5 is not a whole number ratio and so the lowest multiple, which gives a whole number ratio is the empirical formulae, i.e 1:2.5, **2:5.0**, 3:7.5 etc.

Hence the empirical formula = P_2O_5 .

PERIOD 4: RELATIVE MOLECULAR MASS, MOLAR MASS AND PERCENTAGE COMPOSITION

If the formulae of a substance and the relative atomic Masses of each of the elements are known, then it is possible to determine the relative molecular mass of that substance.

The relative molecular mass refers to the number of times a mole is heavier than one-twelfth the mass of one atom of carbon -12. It has no unit.

The relative molecular mass of a compound is the sum of the masses of all the atoms present in one molecule of the compound .e.g.

For NaCl, the relative molecular mass= $(23 + 35.5) = 58.5$

For ethanol = $\text{C}_2\text{H}_5\text{OH}$ (carbon=12, H=1, O =16)

The relative molecular mass of ethanol

= $\text{C}_2\text{H}_5\text{OH}$

$$\begin{aligned}(12 \times 2) + (1 \times 5) + (16) + (1) \\ 24 + 5 + 16 + 1 = 46\end{aligned}$$

THE MOLAR MASS

This is the relative molecular mass expressed in grams. E.g. the molar mass of ethanol is 46g mol^{-1}

Second term Chemistry E-Lesson Note

In 12g of carbon-12, there are 6×10^{23} atoms of carbon. This is one mole of carbon -12.

A mole of any substance is the amount of that substance which contains 6×10^{23} particles of that substance e.g.

One mole of ethanol has a mass of 46g and contains 6×10^{23} ethanol molecules.

NOTE: The relative molecular mass has no units but the molar mass of any substance is expressed in grams.

PERCENTAGE COMPOSITION OF A COMPOUND

To calculate the percentage position of ethanol whose molecular formula is C_2H_5OH , given that the relative atomic masses of carbon, hydrogen and oxygen are 12, 1, and 16 respectively?

First calculate the molar mass of $C_2H_5OH = (12 \times 2) + (1 \times 5) + (16) + (1) = 46 \text{ g mol}^{-1}$

Then determine the masses of C H and O present;

Mass of carbon = $12 \times 2 = 24 \text{ g}$

Mass of hydrogen = $6 \times 1 = 6 \text{ g}$

Mass of oxygen = $16 \times 1 = 16 \text{ g}$

Molar mass of $C_2H_5OH = 46 \text{ g}$

Therefore, percentage of C = $\frac{24}{46} \times 100 = 52.17\%$

Percentage of hydrogen = $\frac{6}{46} \times 100 = 13.04\%$

Percentage of oxygen = $\frac{16}{46} \times 100 = 34.78\%$

CALCULATION OF THE CHEMICAL FORMULA FROM PERCENTAGE COMPOSITION BY MASS.

We can determine the simplest chemical formula of a compound, given its percentage composition e.g. If the formula for anhydrous disodium trioxocarbonate (iv) is not known, if its percentage composition by mass is known then its chemical formula could be calculated.

For example, the percentage composition of the compound was found to be Na=43.40%, C= 11.32% and O = 45.28%. This would mean that in every 100g of the compound, the masses of Na, C and O were 43.40g, 11.32g and 45.28g respectively.

∴ The amount in moles of Na, C and O would be.

$\frac{43.40}{23}$, $\frac{11.32}{12}$ and $\frac{45.28}{16}$ respectively.

Therefore, the amount in moles of Na = $\frac{43.40}{23} = 1.89$.

Amount in mole of C = $\frac{11.30}{12} = 0.94$.

Amount in mole of O = $\frac{45.28}{16} = 2.83$.

Molar ratio of H: C: O is 1.89: 0.94: 2.83.

2: 1 : 3

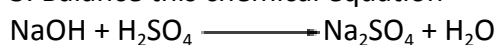
Na: C: O = $Na_2C O_3$

The simplest formula is therefore, Na_2CO_3

GENERAL EVALUATION

ESSAY QUESTION

1. Differentiate between valency and oxidation number.
2. Determine the empirical formula of an oxide of nitrogen containing 70% oxygen, if the relative molecular mass of the oxide is 92, deduce its molecular formula.
3. Balance this chemical equation



WEEK 3

PERIOD 1: CHEMICAL EQUATIONS

Chemical equations is the representation of chemical reaction in terms of the symbols and formulars of the elements and compound involved. The substances that participate in the reaction, i.e. reactants, are written on

Second term Chemistry E-Lesson Note

the left hand side, while the product of the reaction are written on the right hand side. Reactants and products are linked together by an arrow symbol (\longrightarrow) whose head points towards the products. However, reversible reaction is one which can be made to proceed in either direction under suitable conditions. The reactants can form products, so also the products can form reactants. Reversible reactions is denoted by the symbol (\rightleftharpoons). Chemical reactions are represented in form of equations which show the reactants and products in any given chemical reaction. For example, the reaction of aqueous hydrogen chloride and aqueous sodium hydroxide is represented by the equation: $\text{HCl}_{(aq)} + \text{NaOH}_{(aq)} \longrightarrow \text{NaCl}_{(aq)} + \text{H}_2\text{O}_{(l)}$. While a reversible reaction is $\text{NH}_4\text{Cl}_{(s)} \rightleftharpoons \text{NH}_3_{(g)} + \text{HCl}_{(g)}$

Balancing Chemical Equations

All chemical equations must be balanced in order to comply with the law of conservation of mass. For example, to balance the equation for the reaction involving the combustion of ammonia gas in air, the following steps should be followed:

Step 1: The reactants (LHS) are NH_3 and O_2 , while the products (RHS) are NO and H_2O



Step 2: Place a 4 in front of NH_3 and a 6 in front of H_2O . There are now 12 hydrogen atoms on both sides of the equation.



Step 3: Place a 4 in front of NO to balance with the 4 nitrogen atoms of the LHS of the equation.



Step 4: Place a 5 in front of O_2 . An "atom count" shows that the equation is balanced.



An equation must be balanced. A balanced equation contains the same number of atoms of the elements of the compounds on both sides of the equation.

The equation must also show the physical states of the reactants and the products i.e. whether in aqueous solution or gaseous or solid state.

When Balancing a chemical reaction, the following must be remember

1. Equations must be balanced through the use of co-efficient in front of the formular and not by changing the subscript numbers within the formulae
2. Common gases such as Hydrogen, Oxygen, Nitrogen and Chlorine in the free state are diatomic e.g. H_2 , O_2 , N_2 , and Cl_2 .
3. Other elements such as Sodium, Potassium, Copper and Iron in the Free State are represented by their atomic symbols e.g. Na, K, Cu and Fe
4. In simple chemical reactions, radicals remain unchanged during the reactions.

Information Provided by a chemical equation

1. The stoichiometry of the reaction (i.e. the relationship between the amounts of the reactants and the products) in terms of mole ratio and the mass ratio of the reactants and products.
2. The molar mass of each compound from the relative atomic masses of elements forming the compound
3. The relative volume of reactants and products if gaseous (1 mole occupies 22.4dm^3 at s.t.p)
4. The reactants and products involved
5. Which reactants are in excess in a mixture of known amount of reactant and how much.
6. The state of matter of the reactants and products as indicated by the state symbols (s) for solid, (l) for liquid, (g) for gas and (aq) for aqueous solution.

Information Not provided by a chemical equation

1. The physical state, such as colours, of the reactants and products
2. The conditions of the reaction i.e. whether the solutions are concentrated or diluted, heat change.
3. The speed of the reaction.

Evaluation

Balance the equations below:

Second term Chemistry E-Lesson Note

- $\text{NaOH}_{(aq)} + \text{HCl}_{(aq)} \longrightarrow \text{NaCl}_{(aq)} + \text{H}_2\text{O}_{(l)}$
- $\text{KClO}_{3(s)} \longrightarrow \text{KCl}_{(s)} + \text{O}_{2(g)}$
- $\text{NaOH}_{(aq)} + \text{CO}_{2(g)} \longrightarrow \text{Na}_2\text{CO}_{3(s)} + \text{H}_2\text{O}_{(l)}$
- $\text{C}_{(s)} + \text{H}_2\text{O}_{(g)} \longrightarrow \text{CO}_{(s)} + \text{H}_{2(g)}$
- $\text{BaCl}_{2(aq)} + \text{H}_2\text{SO}_{4(aq)} \longrightarrow \text{BaSO}_{4(s)} + \text{HCl}_{(aq)}$
- $\text{NH}_{3(g)} + \text{O}_{2(g)} \longrightarrow \text{H}_2\text{O}_{(g)} + \text{N}_{2(g)}$
- $\text{Al}(\text{OH})_{3(s)} + \text{HNO}_{3(aq)} \longrightarrow \text{Al}(\text{NO}_3)_{3(aq)} + \text{H}_2\text{O}_{(l)}$
- $\text{H}_2\text{SO}_{4(aq)} + \text{KOH}_{(aq)} \longrightarrow \text{K}_2\text{SO}_{4(aq)} + \text{H}_2\text{O}_{(l)}$
- $\text{NH}_{3(g)} + \text{O}_{2(g)} \longrightarrow \text{NO}_{(g)} + \text{H}_2\text{O}_{(g)}$
- $\text{ZnCO}_{3(s)} + \text{HCl}_{(aq)} \longrightarrow \text{ZnCl}_{2(aq)} + \text{H}_2\text{O}_{(l)} + \text{CO}_{2(g)}$
- $\text{CaCO}_{3(s)} + \text{HCl}_{(aq)} \longrightarrow \text{CaCl}_{2(s)} + \text{H}_2\text{O}_{(l)} + \text{CO}_{2(s)}$
- $\text{CO}_{2(g)} + \text{NaOH}_{(aq)} \longrightarrow \text{Na}_2\text{CO}_{3(aq)} + \text{H}_2\text{O}_{(l)}$

WEEK 4 AND 5

TOPIC: LAWS OF CHEMICAL COMBINATION

CONTENT:

- Law of conservation of matters
- Law of constant composition
- Law of multiple proportions.

PERIOD 1: CHEMICAL LAWS OF COMBINATIONS

There are four laws of chemical combination which describe the general features of a chemical change.

Law of conservation of mass: This law was established by Lavoisier, a French chemist. The law of conservation of mass states that matter is neither created nor destroyed during chemical reaction, but changes from one form to another.

Experiment to verify the law of conservation of matter (mass)

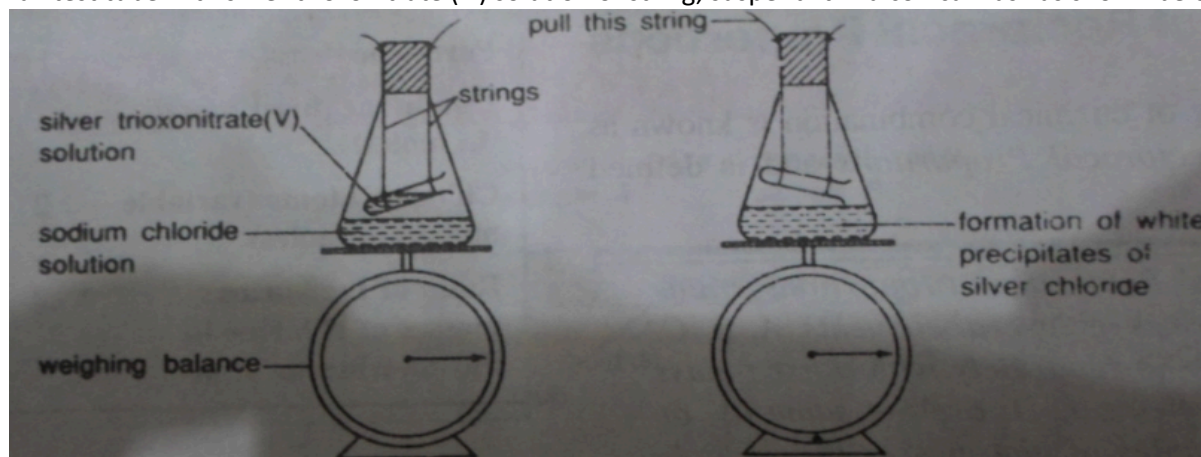
Theory:

The equation of the chemical reaction chosen for study is as follows;

Silver nitrate + sodium chloride \rightarrow Silver chloride + Sodium trioxonitrate(v)
(White precipitate)

Method:

- Put some sodium chloride solution in a conical flask
- Fill a small test tube with silver trioxonitrate (iv) solution of string, suspend it in a conical flask as shown below:



Insert the stopper and weight the whole apparatus on a balance, note the mass of the whole system.

Second term Chemistry E-Lesson Note

Mix the two liquids by pulling the string attached to the bottom end of the small test tube.

Weigh the whole apparatus again.

Result: When the two reactants are mixed together, a white precipitate is formed indicating that a chemical reaction has taken place. The new substances formed are known as the products of the chemical reaction. The masses of the system taken before and after the reaction are found to be the same, indicating that the mass of the reactants equals that of the products.

CONCLUSION: Since there is no overall change in mass when the products are formed, we can infer that matter is neither created nor destroyed during the chemical reaction. The law is, hence valid.

EVALUATION:

1. Mention another compound that could be used instead of silvertrioxo-nitrate (v) with sodium chloride
2. State the law of conservation of mass/matter.

PERIOD 2: LAW OF DEFINITE PROPORTION OR LAW OF CONSTANT COMPOSITION

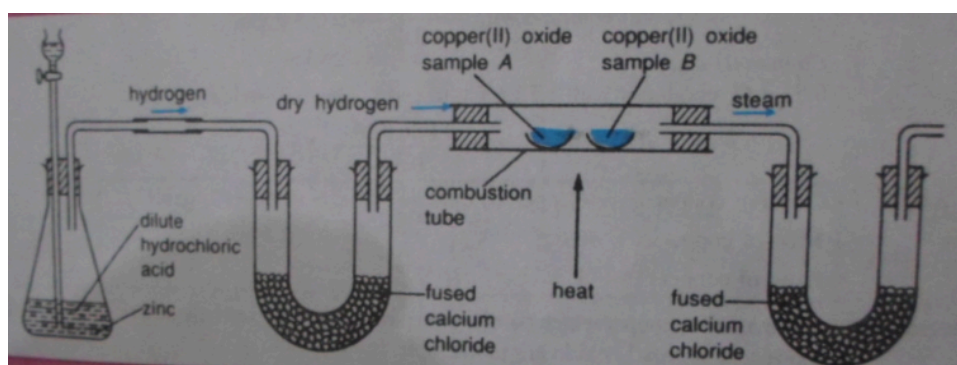
The second law of chemical combination which is supported by the Atomic theory was proposed by provost (1755-1826) known as the Law of definite proportions or constant composition.

The law of definite proportions states that all pure samples of a particular chemical compound contain similar elements combined in the same proportion by mass. It is based on the fact that when elements combine to form a given compound, they do so in fixed proportions by mass, so that all pure samples of that compound are identical in composition by mass.

Water for example: chemical analyses showed that as long as it is pure, its composition is always in the ratio of one mole of oxygen to two moles of hydrogen. i.e. 32g of O to 4g of H. Irrespective of whether the water comes from river, sea, rain or anywhere.

Experiment to verify the law of definite proportion

Method: Prepare two samples of black copper (ii) oxide, each by a different method as given below:



Sample A: Place some coppers turning in a crucible and add some concentrated trioxonitrate (v) acid, a little at a time, until the copper dissolves completely. Evaporate the resulting green solution of copper II oxide trioxonitrate (v) to dryness; continue to heat the residue until it decomposes to give a black solid which is copper II oxide. Keep the black residue dry in desiccator.

Sample B: Place some copper (i) trioxocarbonate (iv) in a crucible and decompose it into copper (ii) oxide and carbon (iv) oxide store the residue in a desiccator.

ANALYSES:

Determine the amount of copper present in the two samples of copper oxide by reducing the oxide in a stream of hydrogen or carbon II oxide as follows.

1. Weigh two clean metal boats.
2. Add a reasonable amount of sample A to one and sample B to the other
3. Reweigh and determine the mass of each sample. Place the boats inside a hard glass tube as shown. Heat the samples strongly while passing a stream of dry hydrogen gas through the tube. After some time, a reddish-brown copper residue is left in each boat. Remove the flame, but continue passing the hydrogen as the copper residues cool down. This prevents the re-oxidation of the hot copper residue by atmospheric oxygen. Any water formed during the reaction is absorbed by the fused calcium chloride in the adjacent U-tube.

Result:

Sample	A	B
Mass of copper II oxide	3.55g	3.02g
Mass of copper residue	2.81g	2.42g

Second term Chemistry E-Lesson Note

Percentage of copper present in copper (ii) oxide	$\frac{2.81}{3.55} \times 100$	$\frac{2.42}{3.02} \times 100$
	79.2%	80.1%

The percentage of copper residue in the two samples is approximately 80.0, irrespective of the method of preparation of the copper (II) oxide samples.

CONCLUSION: In the pure copper (II) oxide copper and oxygen are always present in a definite proportion by mass of approximately 4 to 1 i.e.

Copper (II)oxide = copper + oxygen

100% 80% 20%
Ratio 4 : 1

EVALUATION:

1. State the Law of Definite proportion.

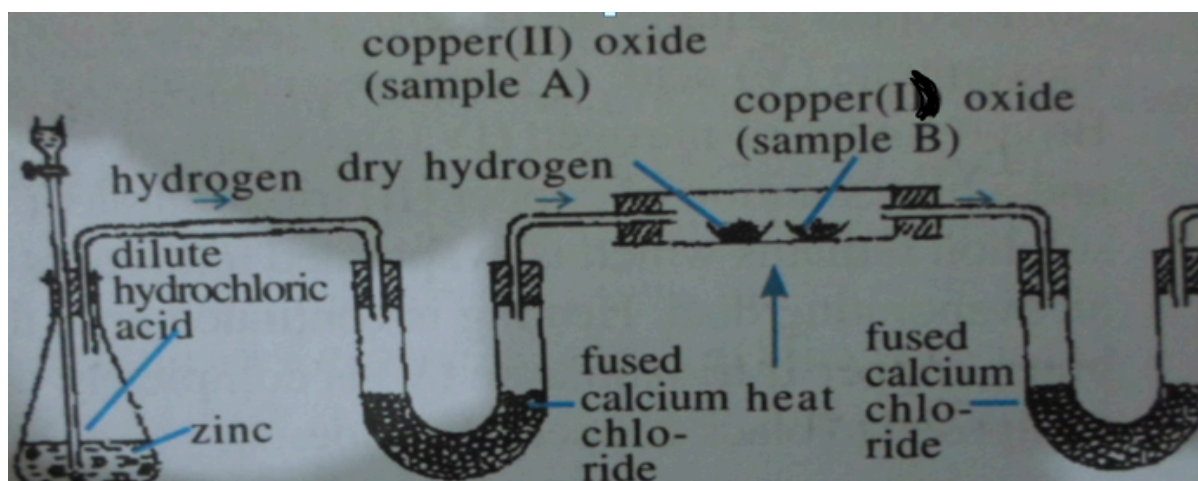
PERIOD 3: LAW OF MULTIPLE PROPORTIONS

This law states that if two elements combine to form more than one compound, the masses of one of the elements which separately combine with a fixed mass of the other element are in simple ratio.

VERIFICATION OF THE LAW OF MULTIPLE PROPORTIONS

Some elements form more than one compound, depending on the conditions of the reaction and the valency copper forms. Copper (I) and copper (II) with oxygen. Also in an insufficient supply of air, carbon burns to form carbon (II) oxide and when the supply of air is sufficient, carbon (iv) oxide is obtained.

The sample of the copper (I) oxide and copper (II) are placed in porcelain, boats and placed in a combination tube as in the diagram below.



A current of dry hydrogen is passed through the combustion tube until the oxides are reduced to metallic coppers. They are now cooled and weighed and the masses of copper and oxygen are determined in the two samples.

Calculations	Sample A	Sample B
(i) Mass of porcelain boat	4.55g	5.38g
(ii) Mass of porcelain boat + copper oxide	6.44g	8.21g
(iii) Mass of copper oxide	1.89g	2.83g
(iv) Mass of porcelain boat + copper	6.05g	7.90g
(v) Mass of copper (iv) – (i)	1.50g	2.52g
(vi) Mass of oxygen (iii) –(v)	0.39g	0.31g

For example A **1.50g of copper** combines with **0.39 of oxygen**.

∴ 100g of copper combines with $\frac{0.39}{1.50} \times 100 = 26\text{g}$

For sample (b) 2.52g of copper combines with 0.31g of oxygen

∴ 100g of copper combines with $\frac{0.31}{2.52} \times 100 = 12.3\text{g}$

From these calculations, the masses of oxygen (26g and 12.3g) which combine with a fixed mass (100g) of copper are in simple ratio 2:1

PERIOD 4: LAW OF RECIPROCAL PROPORTION

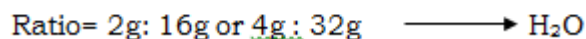
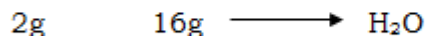
Second term Chemistry E-Lesson Note

This is the fourth law of chemical combination. **This law states that the masses of several elements, A, B, C, which combine separately with a fixed mass of another element, D, are the same as, or simple multiples of, the masses in which A, B, C, themselves combine with one another.** For example C, H, O (12, 1, 16) respectively. Carbon and hydrogen combine to form methane (CH₄). Carbon and oxygen combine to form carbon (iv) oxide, (CO₂) and hydrogen and oxygen combine to form water (H₂O).

In water,

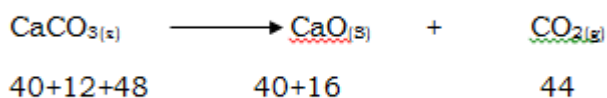


2 moles + 1 mole



This is the prediction of the law of reciprocal proportions. For example, 23g of calcium trioxocarbonate (iv) on heating decomposes to give calcium oxide (CaO) and carbon (iv) oxide. Calculate the masses of calcium oxide and carbon (iv) oxide produced [C= 40, O = 16, C= 12]

Solution



100g of CaCO₃ yield 56g of CaO

$$\therefore 23\text{g of CaCO}_3 \text{ will yield } \frac{56 \times 23\text{g}}{100} = \mathbf{12.88\text{g}}$$

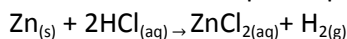
100g of CaCO₃ yields 44g of CO₂

$$\therefore 23\text{g of CaCO}_3 \text{ will yield } \frac{44 \times 23\text{g}}{100} = \mathbf{10.12\text{g}}$$

EVALUATION:

State the law of multiple proportion

State the law of reciprocal proportions.



GENERAL EVALUATION:

ESSAY QUESTIONS

	Copper(I)oxide	Copper(II)oxide
Mass of sample oxide	30.4g	1.91g
Mass of copper residue	2.55g	1.38g
Mass of oxygen removed from oxide	0.49g	0.53g

From the above table, calculate the various masses of copper which would combine separately with a fixed mass of 1 g of oxygen.

2. What mass of copper will be produced from the reduction of 7.95g of copper (II) oxide? (C= 63.5, O= 16)

3. Write down the names of these chemical compounds :(i) HNO₃ (ii) CuCl₂(iii) CaCO₃(iv) Fe₂O₃

4. Write the symbol and the valency of the following.(i) Boron (ii) Carbon (iii) Sulphur (iv) Argon

5. Calculate the formula of a compound with 31.9% potassium 28.93%, chlorine and the rest oxygen. K=39, Cl =35.5, O=16

OBJECTIVE TEST

1. Which of the following relative molecular mass has empirical formula CH₂O (H=1 C=12, O=16). (a) 42 (b) 80 (c) 4 (d) 60

2. The relative molecular mass of tetraoxosulphate(VI) acid is? (a) 98 (b) 49 (b) 49 (c) 96 (d) 106

3. Chemical equations will provide all these except. (a) State of chemicals is solved (b) Direction of reaction (c) Mass of products (d) Reactants

4. All pure samples of a particular compound contain the same elements combined in the same proportion by mass. The statement is the law of (a) Definite proportion (b) Multiple proportion (c) Conservation of mass or matter (d) Atomic proportion

Second term Chemistry E-Lesson Note

WEEK 6

TOPIC: CHEMICAL COMBINATION

CONTENTS

- (a) Types of chemical bonds
- (b) IUPAC system of naming inorganic compounds.

PERIOD 1: PERIODIC TABLE (FIRST 20 ELEMENTS)

Chemical combination: The atoms of noble gases e.g. Helium He, Neon Ne are very stable; because their outermost shells are completely filled with electrons. The tendency of the other elements is to attain this stable structure possessed by the noble gases. This is achieved during **chemical combination**. Thus, **chemical bonding** is the coming together of atoms of the same or different elements, in order to form a stable structure.

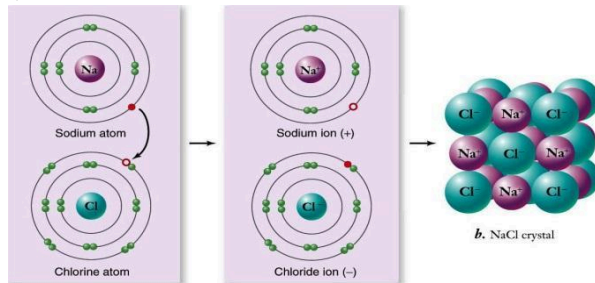
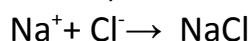
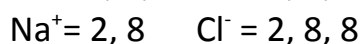
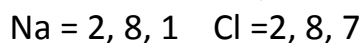
PERIODS 2 and 3: TYPES OF CHEMICAL BONDS

The attractive force between atoms when they combine chemically is called a **chemical bond**. There are two main types of chemical bonds namely (i) Strong bonds (ii) Weak bonds.

(i) Strong bonds are: (a). Electrovalent (or Ionic) (b). Covalent (c). Co-ordinate (or Dative) (d). Metallic

(a) Electrovalent (Ionic) bond is defined as the electrostatic force of attraction between oppositely charged ions. It involves the transfer of electrons from one atom, donor atom, (usually metallic) to another atom, acceptor atom, (usually non-metallic). The electrons involved reside in the outermost shells of the atoms and are called Valence electron(s).

IONIC COMPOUNDS are crystal lattices consisting of aggregates of oppositely charged ions. Examples of such compounds are Sodium chlorides, calcium oxides, ammonium chlorides etc. The formation of sodium chloride is illustrated below:



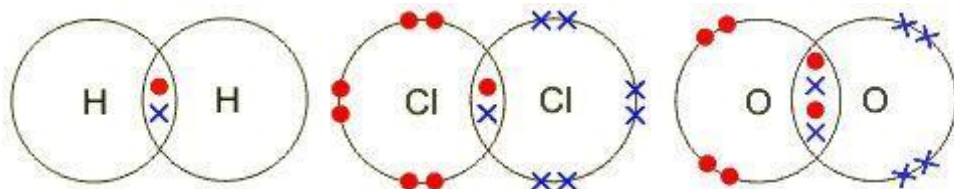
Characteristics of Electrovalent bonds

- i. They have high melting and boiling points
- ii. They are generally soluble in water
- iii. They are good conductors of electricity when molten or in solution
- iv. They do not conduct electricity when solid
- v. The energy needed to separate them is relatively high.

(b) Covalent bonds: This is defined as the bonds formed when two atoms donate equal numbers of electrons and share the donated electrons to attain stable octet structure. In covalent bonding, electrons are shared between atoms of the same or different elements such that each atom contributes the shared electrons so as to attain stable noble gas configuration. During the process, discrete or separate molecules are formed with covalent bonds between the atoms.

If electrons are shared between similar atoms, the donated electrons are equally shared which give rise to a **non-polar covalent bond** e.g. F_2 , H_2 , O_2 , Cl_2 but if the shared pair of electrons involves two different elements with difference in electronegativity, the electrons are then not equally shared, this results in the formation of **polar covalent bond**, e.g. the formation of HCl , H_2O , HF .

Second term Chemistry E-Lesson Note

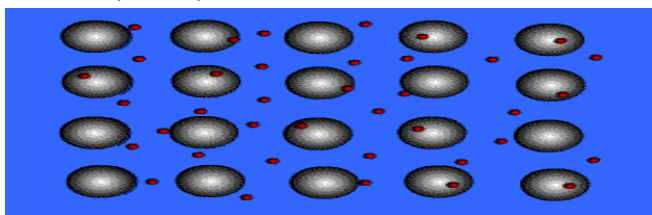


Characteristics of Covalent bonds

- Low melting and boiling points.
- Energy required for separation is low
- Do not conduct electricity in the solid or molten state, or in solution
- They generally have a strong, easily noticeable smell
- They are not easily soluble in water, but are usually soluble in organic solvents

(c) Co-ordinate covalent or Dative bond involves sharing of electrons as in the normal covalent bonding, but the shared pair is donated by only one of the participating atoms. For instance, Ammonia and water molecules possess lone pairs and so readily enter into coordinate covalent bonding.

(d) Metallic bond: The electrostatic force of attraction between the positive nuclei and the sea of mobile electrons is called metallic bond. Metallic bonding, therefore, is the process whereby the positively charged nuclei of metal atoms are simultaneously attracted to the sea (or cloud) of mobile electrons. Metallic bond increase with increase in valence electrons of the metal. For example, in period 3, metallic bond increases from sodium to aluminum.



(ii) Weak bonds: These are intermolecular forces of attraction that hold atoms and covalent molecules together in gases, liquids and solids. The most common ones are: (a). Van der Waals forces (b). Hydrogen bond (c) Dipole-Dipole (Dipolar)

(e) Van der Waals forces: They were first described by J.D. van der Waals, and are known as van der Waals forces. They are weak short-ranged attractive forces formed between covalent molecules. They are the only attractive forces between the atoms of the noble gases and non-polar covalent molecules, and are responsible for the low melting and boiling point of covalent compounds. Due to increase in van der Waals forces, there is gradation in the physical properties of the Halogens: Fluorine and chlorine are gases; bromine is a liquid; and iodine is a solid.

Van der Waal force becomes stronger as the relative atomic mass increases among the non-metals.

(f) Hydrogen bond: This is an intermolecular force which arises when hydrogen is covalently linked to highly electronegative elements like nitrogen, oxygen and fluorine.

The presence of hydrogen bonds between H₂O molecules is responsible for water being a liquid at room temperature and with a high boiling point; if not, it would have been a gas, like hydrogen H₂S. HF is a liquid at room temperature, while HCl is a gas.

EVALUATION

What is a chemical bond?

List three (3) general properties of electrovalent compounds.

Define covalent bond and state its characteristics

State two (2) differences between covalent and electrovalent compounds.

Explain the term *Metallic bonding*

WEEK 7

CONTENTS:

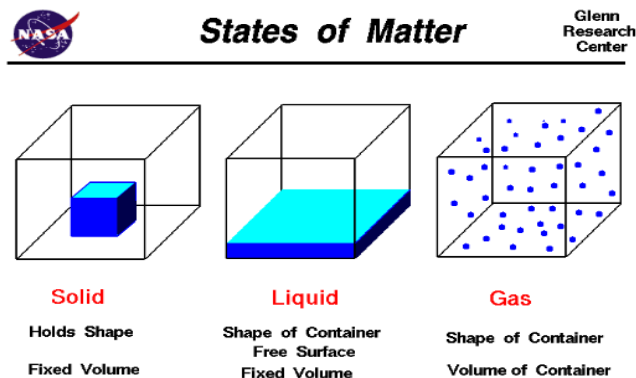
1. State of matter,
2. Kinetic theory of gases,
3. Applications of kinetic theory of gases

Second term Chemistry E-Lesson Note

4. Phenomena to illustrate kinetic theory of matter

PERIOD 1: STATE OF MATTER

Matter is made up of very tiny particles such as atoms, molecules and ions. Matter exists in three physical states namely; solid, liquid and gas. Matter has mass and occupies space.



(i) Solid: The tiny particles in solid are packed very closely together hence they cannot move about. The presence of forces of attraction among the tiny particles of solid made them to be very closely packed together and allowing only vibration among them instead of moving about. This is what makes the solid to have a fixed shape and volume with high density. However, at high temperature, the molecules gain more energy and break the forces of attraction among the particles thereby causing the particles to move faster. As the temperature increases more and more, the solid melts and turns to liquid. Therefore, solid have fixed shape and volume, incompressible and very dense with the least kinetic energy when compared with those of liquid and gasses.

(ii) Liquid: The particles in the liquid are slightly further apart than those in a solid and there exist among them weaker forces of attraction (weaker force of cohesion); thus allowing them to vibrate rotate and move about. Liquid have fixed volume and a fixed shape but takes the shape of its container. When the liquid is heated, its temperature increases, the forces of cohesion / attraction becomes progressively weaker. Eventually, a stage is reached when the molecules acquire enough energy to escape as a gas. The temperature at which this occurs is called the boiling point of the liquid. Therefore, the liquid have fixed volume, no fixed shape, less dense and incompressible with kinetic energy relatively higher than those of solid state.

(iii) Gas: The particles in the gaseous state are very freely with large distances between them. The particles of gases have no fixed volume and shape, compressible and least dense with the highest kinetic energy when compared with particles of other states of matter.

EVLUATION

1. Define the term matter.
2. Enumerate the three main state of matter.
3. Explain each of the states.

PERIOD 2: KINETIC THOERY OF GASES

EXPLANATION OF KINETIC THEORY

The kinetic theory of matter postulates that the tiny Particles of matter are continually moving and so possess kinetic energy. An increase in temperature causes an increase in the average kinetic energy of the particle.

Dalton's experimental evidence shows that chemical compounds consists of molecules, which are groups of atoms of various elements. The gas laws which explained the physical behaviour of gases can be explained by kinetic theory of gases. This theory describes the behaviour of an ideal or perfect gas. This is to say that the kinetic theory of gases explains quantitatively the properties of gas molecules and in so doing put up the following assumptions.

1. A gas consists of very tiny particles (usually molecules and atoms).
2. The cohesive forces of the gas molecules are negligible.
3. These particles are in constant random and rapid motion in straight lines.
4. As a result of these movements, collisions occur between the molecules and also with the walls of the container, hence the molecules exert pressure. No energy is lost when collision occupy. This means that the collisions are perfectly elastic.

Second term Chemistry E-Lesson Note

5. The space between the molecules is very large compared to the size of the molecules. The molecules therefore have negligible (almost zero) volume compared to the volume of the container.
6. The average kinetic energy per molecule is the same for all gas samples at any given temperature. The absolute temperature of the gas is a measure of the average kinetic energy of the gas particles.

EVALUATION

The kinetic theory of matter postulates

When the temperature of the particles of the molecules increases, what happens to kinetic energy?

Enumerate the six assumptions of the kinetic theory of matter.

PERIOD 3: THE APPLICATION OF KINETIC THEORY

The kinetic theory is useful in several respects and such aspects include;

1. It provides reasonable explanations for the behaviour of gas.
2. It accounts for the gas laws
3. It explains important phenomenon such as diffusion
4. It provides a fundamental equation for gases.

$$PV = \frac{1}{3} NMC^2$$

EVALUATION

Mention four applications of kinetic theory of matter.

PERIOD 4: PHENOMENA TO ILLUSTRATE THE KINETIC THEORY OF MATTER

The following phenomena illustrate the kinetic theory of matter

1. Dispersion
2. Evaporation
3. Diffusion:
4. Brownian motion
5. Osmosis
6. Tyndall effect
7. Effusion

GENERAL EVALUATION

OBJECTIVE TEST:

1. The kinetic theory of matter states _____
2. Water (H₂O) exists as a solid, liquid and gas respectively because: (a) water is colorless. (b) water is electrovalent (c) Water in any state possesses a certain degree of motion in the molecules (d) water is molecular
3. Which of the three states of matter has no fixed volume and least dense? (a) Gas (b) Solid (c) Liquid
4. The presence of sodium chloride in ice will. (a) Decrease or lower the boiling point of sodium chloride (b) Increase the melting point of sodium chloride (c) Make sodium chloride impure (d) Lower the freezing point of sodium chloride.
5. The escape of molecules with more than average kinetic energy of the molecules is called _____ (A) Melting (b) Freezing (C) Evaporation (d) Efflorescence

ESSAY QUESTION:

1. Define the term matter.
2. State the three states of matter.
3. Explain two out of the three main states of matter.
4. List four importance of kinetic theory.
5. Give assumptions of kinetic theory of matter.

ASSIGNMENT:

Read New School Chemistry for Senior Secondary Schools, by Osei Yaw Ababio pages 584-594

PRE- READING ASSIGNMENT

Read the meaning of chemical industry and its development.

Second term Chemistry E-Lesson Note

WEEKEND ACTIVITY

Explain with at least three points what you understand by chemical industry and also mention specifically five types of chemical industry.

REFERENCE TEXT:

1. New School Chemistry for senior Secondary Schools by Osei Yaw Ababio 6th edition
2. Comprehensive certificate chemistry by G.N.C Ohia, G.I, Amasiatiu, J.O Ajagbe =, G.O. Ojokuku and U Mohammed. 2nd Edition

WEEK 8

TOPIC: GASEOUS STATE AND GAS LAWS

CONTENTS:

1. Boyle's law
2. Charles' law
3. General gas law
4. Ideal gas law

PERIOD 1: BOYLE'S LAW

The relationship between volume and pressure of a gas was first started by Robert Boyle in 1662.

Boyle's law states that the volume of a given mass of gas is inversely proportional to its pressure, provided that the temperature remains constant.

According to Boyle's law, volume of a gas increases as the pressure decreases and vice versa.

This relationship is independent of the nature of the gas and it can be expressed mathematically as:

$$V \propto \frac{1}{P}$$

$$\therefore V = \frac{k}{p}$$

$$\text{Or } PV = k$$

Where V= volume at pressure P

K = a mathematical constant

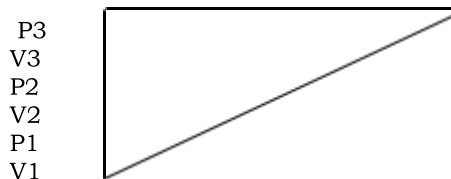
For a given mass of a gas, the product of its pressure and its volume is always a constant. If the pressure of a given mass of gas increases, its volume will decrease by a similar proportion and vice versa, as long as the temperature remain constant. This relationship can also be expressed mathematically as:

$$P_1 V_1 = P_2 V_2$$

Where V_1 = volume at pressure P_1

V_2 = Volume at pressure P_2

Boyle's law can still be re-stated as: 'The pressure of a given mass of gas is inversely proportional to its volume, provided the temperature remains constant. Boyle's law can further be illustrated with the diagram below, showing that when P is increasing, V is decreasing and when P is decreasing, V is increasing.



P_1 = Initial pressure

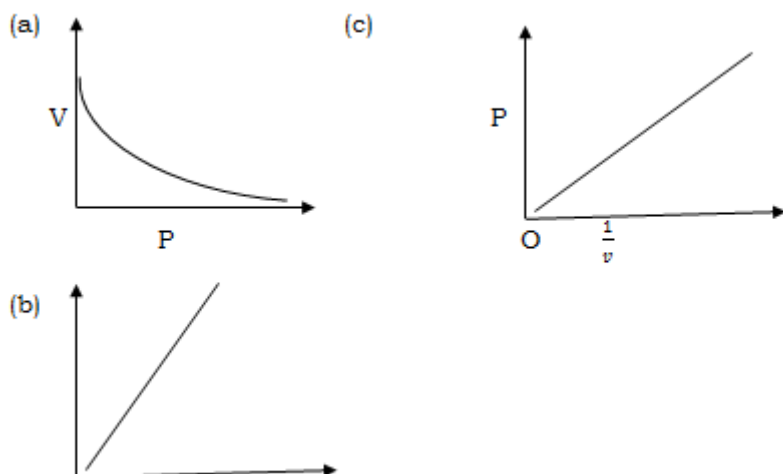
P_2 = Final pressure

V_1 = Initial volume

V_2 = final volume

Second term Chemistry E-Lesson Note

Graphical representation of Boyle's law



Worked examples

1. 375cm^3 of a gas has a pressure of 770mmHg . Find its volume if the pressure is reduced to 750mmHg .

$$P_1V_1 = P_2V_2 \text{ (Boyle's law)}$$

$$P_1 = 770\text{mmHg}$$

$$P_2 = 750\text{mmHg}$$

$$V_1 = 375\text{cm}^3$$

$$V_2 = ? \text{ (New volume of gas)}$$

$$P_1V_1 = P_2V_2$$

$$V_2 = \frac{P_1V_1}{P_2} = \frac{770 \times 375}{750} = 385\text{cm}^3$$

The new volume will be 385cm^3

2. 100cm^3 of a gas has pressure of 1 atmosphere. Determine the volume of the gas at 5 atmospheres keeping the temperature constant.

Solution: since T is constant, we are to use Boyle's law.

$$P_1 \rightarrow \text{Initial pressure} = 1 \text{ atmosphere}$$

$$P_2 \rightarrow \text{Final pressure} = 5$$

$$V_1 \rightarrow \text{Initial volume} = 100\text{cm}^3$$

$$V_2 \rightarrow \text{(required quantity)}$$

$$\text{Recall: } V_2 = \frac{P_1V_1}{P_2} = P_2V_2$$

$$V_2 = \frac{P_1V_1}{P_2} = \frac{100 \times 1}{5} = 20\text{cm}^3$$

EVALUATION:

1. A given quantity of gas occupies a volume of 228cm^3 at a pressure of 750mmHg . What will be its volume at atmospheric pressure if temperature is kept constant?

2. A given mass of gas at 55°C has a pressure of $3.6 \times 10^4 \text{Nm}^{-2}$ and occupies a volume of 1.8dm^3 . What volume will it occupy if its pressure is increased to $4.8 \times 10^4 \text{Nm}^{-2}$ if the temperature is kept constant?

PERIOD 2: CHARLES' LAW

The effect of temperature changes on the volume of a given mass of a gas at a constant pressure is described by Charles. Charles' law states that *the volume of a given mass of gas is directly proportional to its temperature in Kelvin, provided that pressure remains constant.*

The volume of the gas decreases as the temperature decreases, and increases as the temperature increases.

Mathematically, the law can be expressed as:

Second term Chemistry E-Lesson Note

$$V \propto T$$

$$\therefore V = kT$$

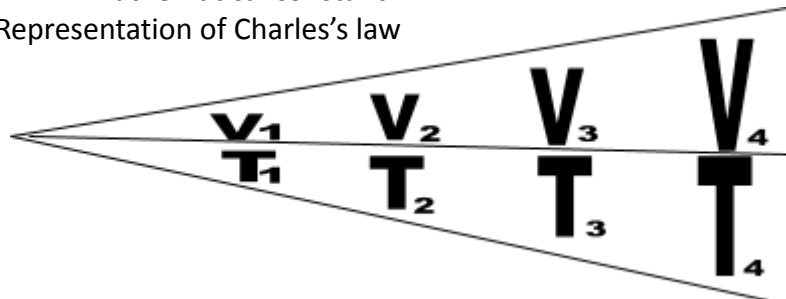
$$\text{Or } \frac{V}{T} = k$$

Where v= volume

T= Kelvin Temperature

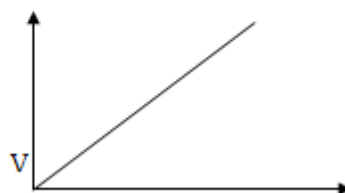
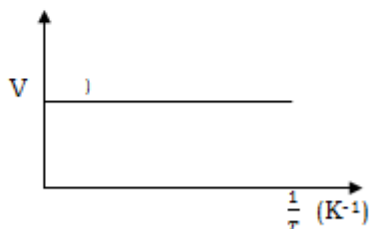
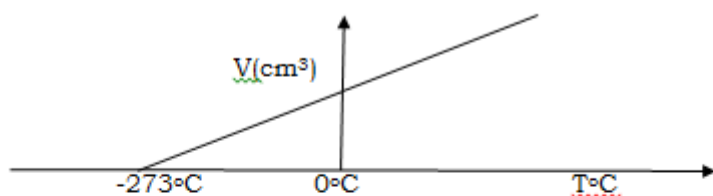
K= mathematical constant

A Representation of Charles's law



For a direct relationship, when the temperature increases, the volume will also increase at the same rate and vice versa, at constant pressure. The diagram above shows that when V is decreasing, T is also decreasing and when V is increasing, T is also increasing thus, making the quotient constant.

Charles's law can be represented graphically as shown below.



If we divide the varying gas volumes by the corresponding temperature in Kelvin, the result would always be a constant. This relationship can also be expressed in another form.

$$\frac{V_1}{T_1} = \frac{V_2}{T_2} \therefore V_2 = \frac{T_2 V_1}{T_1}$$

Where V_1 is the volume at temperature T_1

V_2 is the volume at temperature T_2

ABSOLUTE ZERO

This is the temperature at which the volume of a gas is theoretically zero. At this temperature there is no motion of any form and all gases have been liquefied or solidified. The value of the temperature is -273°C .

TEMPERATURE CONVERSION

1. To convert from Celsius scale to Kelvin scale, add 273 i.e. $T = ^\circ\text{C} + 273$. This is because $0^\circ\text{C} = 273\text{K}$.

2. To convert from Kelvin scale to Celsius scale, subtract 273. i.e.

$$^\circ\text{C} = T - 273.$$

Where T= Temperature in Kelvin

$^\circ\text{C}$ = Temperature in Celsius.

Examples:

1. Convert the following Celsius temperature to Kelvin temperature.

Second term Chemistry E-Lesson Note

100°C (b) 0°C (c) -57°C

Solution

Recall: $T = 0C + 273$

$100^{\circ}C = (100 + 273) = 373k$

$0C = (0 + 273) = (0 + 273) = 373k$

$-57^{\circ}C = (-57 + 273)k = (273 - 57) = 216k$

2. Convert the following Kelvin temperatures to Celsius temperature.

298k (b) 405k (b) 285k (d) 0k

Solution

Recall $0^{\circ}C = k - 273$

$298k = (298 - 273)^{\circ}C = 25^{\circ}C$

$405k = (405 - 273)^{\circ}C = 12^{\circ}C$

$0k = (0 - 273)^{\circ}C = -273^{\circ}C$

Worked examples on Charles's law

1. A gas occupies a volume of $20.0dm^3$ at 373k. Its volume at 746k at that pressure will be?

Here pressure is constant. Charles's law will apply.

$V_1 = 20.0dm^3$

$T_1 = 273k$

$T_2 = 746$

$V_2 = ?$

Recall Charles's law = $\frac{V_1}{T_1} = \frac{V_2}{T_2}$ $V_2 = \frac{V_1 T_2}{T_1} = \frac{20 \times 746}{273} = 40.0dm^3$

EVALUATION:

1. State Charles's law
2. Express the two laws mathematically
3. Draw two graphs to illustrate Charles' law.

PERIOD 3: GENERAL GAS LAW

From the gas laws, we know that the volume of a gas depends on both its temperature and pressure. The relationship between the three variable; i.e. volume,, temperature and pressure can be summarized up as follows:

If $V \propto \frac{1}{P}$ (Boyle's law at constant temperature) and $V \propto T$ (Charle's law at constant pressure)

$V \propto \frac{1}{P} \times T$ (both temperature and pressure may vary) or $\frac{PV}{T} = K$ (a mathematical constant for a fixed mass of gas)

$\frac{PV}{T} = k$ is often known as the general gas equation.

GENERAL GAS EQUATION

General gas equation states that for fixed mass of a gas under any set of conditions of V , P and T , the value of $\frac{PV}{T}$ must remain constant. If for a fixed mass of gas V_1 is the volume at pressure

P_1 and absolute temperature T_1 and V_2 is the volume at pressure P_2 and absolute temperature T_2 it follows that.

$$\frac{P_1 V_1}{T_1} = \frac{P_2 V_2}{T_2}$$

The general gas equation can be used to find the volume of a gas when both its pressure and temperature change. Thus;

$$V_2 = \frac{P_1 V_1 T_2}{P_2 T_1}$$

The standard temperature and pressure

The value of gases are sometimes given in standard temperature and pressure (S. T. P). These values are standard temperature = 273k and standard pressure = 760mmHg. The S.I unit of standard pressure when used is $1.01 \times 10^3 Nm^{-2}$

Examples

Second term Chemistry E-Lesson Note

1. At S. T. P a certain mass of gas occupies a volume of 790cm^3 , find the temperature at which the gas occupies 1000cm^3 and has a pressure of 720mmHg

$$\frac{P_1 V_1}{T_1} = \frac{P_2 V_2}{T_2}$$

$$P_1 = 760\text{mmHg (at stp)},$$

$$V_1 = 790\text{cm}^3$$

$$T_1 = 273\text{k (at stp)},$$

$$V_2 = 1000\text{cm}^3$$

$$P_2 = 726\text{mmHg}$$

$$T_2 = \text{New Temperature}$$

$$\therefore T_2 = \frac{P_2 V_2 T_1}{P_1 V_1} = \frac{720 \times 1000 \times 273}{760 \times 790} = 330.1\text{k}$$

The new temperature of the gas is 330.1k

2. A given mass of gas occupies 850cm^3 at 320k and $0.92 \times 10^3\text{Nm}^{-2}$ of pressure. Calculate the volume of the gas at S.T.P.

$$\frac{P_1 V_1}{T_1} = \frac{P_2 V_2}{T_2}$$

$$P_1 = 0.92 \times 10^3\text{Nm}^{-2}$$

$$T_1 = 320\text{k}$$

$$V_1 = 850\text{cm}^3$$

$$P_2 = \text{SP} + 1.01 \times 10^3\text{Nm}^{-2}$$

$$T_2 = 273\text{k (at stp)}$$

$$V_2 = \text{new volume of gas.}$$

$$\therefore V_2 = \frac{P_1 V_1 T_2}{P_2 T_1} = \frac{0.92 \times 850 \times 273}{1.01 \times 10^3 \times 320} = 660.5\text{cm}^3$$

EVALUATION

1. Explain the general gas equation.

2. If the volume of a given mass of a gas at 298k and a pressure of $205.2 \times 10^5\text{Nm}^{-2}$ is 2.12dm^3 . What is the volume of the gas S.T.P (standard pressure = $1013 \times 10^5\text{Nm}^{-2}$, standard temperature = 273)

PERIOD 4: IDEAL GAS LAW

The ideal gas: This is a gas sample whose properties correspond, within experimental error, to the relationship $PV = nRT$. An ideal gas must obey all the rules guiding Boyle's and Charles's laws. Ideal gas conforms to the kinetic theory of gases. Four quantities are important in all experimental work, measurements or calculations involving gases. They are: (i) volume (ii) pressure (iii) temperature and (iv) numbers of moles

Ideal gas equation is given by $PV = nRT$

The value of R for one mole of a gas at 273K , 1atm and volume 22.4dm^3 is $0.0821\text{atmdm}^3\text{K}^{-1}\text{mol}^{-1}$ or $8.314\text{JK}^{-1}\text{mol}^{-1}$

Examples:

1. Calculate the volume occupied by 2.5 moles of an ideal gas at -23°C and 4.0atm . ($R = 0.0821\text{atmdm}^3\text{K}^{-1}\text{mol}^{-1}$)

Solution:

$$\text{Using } PV = nRT \text{ where } P = 4.0\text{atm} \quad n = 2.5 \text{ mole} \quad T = -23 + 273 = 250\text{K}$$

$$\text{Hence, } V = \frac{nRT}{P} = \frac{2.5 \times 0.0821 \times 250}{4} = 12.8\text{dm}^3 \\ = 12.8\text{dm}^3$$

NOTE: Pressure can also be measured in other units. $760\text{mmHg} = 1\text{atm} = 101325\text{Nm}^{-2}$

Ideal gases only exist at experimental conditions of high pressure and low temperature. Basically all gases are real

REASONS WHY REAL GASES DEVIATE FROM IDEAL GAS BEHAVIOUR

1. The forces of attraction in real gases are not negligible.
2. The volume of real gases are not negligible. Hence, real gases have their own volume called excluded volume.
3. Real gases undergo inelastic collision

EVALUATION:

1. What is an ideal gas?
2. Write down the ideal gas equation for n -mole of a gas.

GENERAL EVALUATION

OBJECTIVE TEST

Second term Chemistry E-Lesson Note

1. A gas occupies $30.0dm^3$ at S.T.P. What volume will occupy at $91^\circ C$ and $52662.5Nm^{-2}$. (a) $20.0dm^3$ (b) $40.0dm^3$ (c) $60.0dm^3$ (d) $76.96dm^3$
2. Gases can be easily compressed because. (a) The molecule are relatively far apart (b) the molecule are quite close together (c) the molecule are very soft (d) the molecules are in constant, rapid motion
3. A give mass of gas occupies X_1cm^3 at Y_1K . When the temperature is changed to Y_2K , the volume becomes X_2cm^3 , the pressure remaining constant. Which of the following equations correctly express the relationship between X_1, X_2, Y_1 and Y_2 ? (a) $X_1 Y_1 = X_2 Y_2$ (b) $\frac{X_1}{Y_1} = \frac{X_2}{Y_2}$ (c) $X_1 X_2 = Y_1 Y_2$ (d) $X_1 = X_2 Y_1 Y_2$
4. Kelvin temperature can be converted into Celsius temperature by. (a) $^\circ C = K - 273$ (b) $k + 273$ (c) $\frac{0C+273}{k}$ (d) $\frac{k+273}{0C}$
5. What will be the new volume (v) if the new pressure is halved and the initial pressure remain the same. (a) $2p_1 V_1 = p_2 V_2$ (b) $p_1 V_1 = 2p_2 V_2$ (c) $\frac{p_1 V_1}{2} = \frac{p_2 V_2}{2}$ (d) $p_1 V_1 = \frac{p_2 V_2}{2}$

ESSAY QUESTIONS

1. $130cm^3$ of a gas at $20^\circ C$ exert a pressure of $750mmHg$. Calculate its volume is increased to $150cm^3$ at $35^\circ C$
2. Draw the graphical representation of both Boyle's and Charles' law respectively.
3. Convert the following temperature to K. (a) $15^\circ C$ (b) $275^\circ C$ (c) $88^\circ C$
4. The volume of gas at $25^\circ C$ ($298k$) is $100cm^3$. What will be the volume at (a.) $75^\circ C$ ($348k$) (b). $50^\circ C$ ($223k$), pressure remaining constant?

WEEKEND ASSIGNMENT:

Read about Graham's law, Avogadro's number and the mole concept.

WEEK 9

TOPIC: GASEOUS STATE AND GAS LAWS

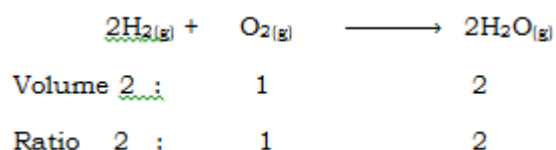
CONTENTS:

1. Gay- Lussac's Law and Avogadro's Law.
2. Graham'slaw of diffusion.
3. Molar volume of gases- Avogadro number and the mole concept.
4. Dalton's law of partial pressure.

PERIOD 1: GAY- LUSSAC'S LAW AND AVOGADRO'S LAW

Gay- Lussac's law describes the combining volumes of gases that react together. In his experiment, all temperatures and pressures were kept constant:

A. STEAM: Gay- Lussac's observed that two volumes of hydrogen reacted with one volume of oxygen to yield two volumes of steam



B. HYDROGEN CHLORIDE GAS: One volume of hydrogen combined with one volume of chloride to yield two volumes of hydrogen.

Hydrogen + Chlorine \rightarrow Hydrogen Chloride

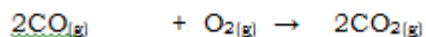
Second term Chemistry E-Lesson Note



Volume = 1 : 1 : 2

Ratio = 1 : 1 : 2

C. Carbon (ii) oxide + Oxygen + Carbon (IV) oxide



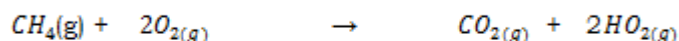
Ratio 2 : 1 : 2

Gay-Lussac's noticed that the combining volumes as well as the volumes of the products, if gaseous, were related by simple ratios of whole numbers. He proposed the law of combining volume or gaseous volumes. Hence; **Gay-Lussac's law combining volumes states that** *when gases react, they do so in volumes which are in simple ratios to one another and to the volumes of the products, if gaseous provide that the temperature and the pressure remain constant.*

EXAMPLES

1. What is the volume of oxygen required to burn completely $45cm^3$ of methane (CH_4)?

Equation of reaction:



Vol: 1 2 : 1 2

Ratio: 1 2 : 1 2

By Gay-Lussac's Law:

1 volume of methane required 2 volumes of oxygen i.e.

$1cm^3$ of methane requires $2cm^3$ of oxygen

$\therefore 45cm^3$ of methane require $90cm^3$ of oxygen

2. $20cm^3$ of carbon (I) oxide are sparked with $20cm^3$ of oxygen. If all the volumes of gases are measured at a S.T.P, calculate the volume of the residual gases after sparking?

Equation of reaction	$2CO_g + O_{2g}$	$2CO_{2g}$
Combining volume	2 : 1	: 2
Volumes before sparking	$20cm^3$ $10cm^3$	$20cm^3$
Volumes after sparking	-10	20

Residual gases = un-reacted oxygen + carbon (IV) oxide formed

Volume of residual gas = $10cm^3 + 20cm^3 = 30cm^3$

AVOGADRO'S LAW

Avogadro's Law states that *equal volumes of all gases at the same temperature and pressure contain the same number of molecules.*

This law means that for all of gases e.g. oxygen, hydrogen, Chlorine etc. if their volumes are the same, they will have the same number of molecules.

Avogadro's Law is easily applied to convert volume of gases to the number of molecules. Avogadro's Law can be used to solve problem under Gay-Lussac's law of combining volumes.

The formation of steam from reaction of Hydrogen and Oxygen is given below:

Reaction: Hydrogen + Oxygen \rightarrow Steam

Volume: 2 1 : 2

Gay-Lussac's: 2 : 1 : 2

Avogadro's Law: 2 : 1 : 2

Second term Chemistry E-Lesson Note

This agrees with the equation below:



i.e. 2 molecules of hydrogen combine with 1 molecule of oxygen to produce 2 molecules of steam

Example:

1. $60cm^3$ of hydrogen are sparked with $20cm^3$ of oxygen at $100^\circ C$ and 1 atmosphere. What is the volume of the steam produced?

Solution



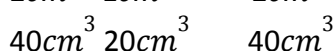
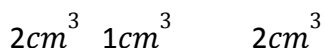
From the equation, 2 molecules of hydrogen react with 1 molecule of oxygen to produce 2 molecules of steam.



From the above information, when $2cm^3$ (2 vol) of H_2 react, $1cm^3$ (1 vol) of O_2 will react i.e. half of H_2 vol, to give $2cm^3$ (2 vol) of H_2O .

Thus, $10cm^3$ of H_2 will react with $5cm^3$ of O_2 to produce $10cm^3$ of H_2O and so on.

From the question, we have $60cm^3$ of H_2 and $20cm^3$ of O_2 , thus, when all the $20cm^3$ of O_2 react, only $40cm^3$ of H_2 will react to give $40cm^3$ of H_2O , because the volume of H_2 is the same as that of H_2O i.e.



Thus, the volume of steam (H_2O) formed is $40cm^3$

2. What volume of propane is left unreacted when $80cm^3$ of oxygen and $20cm^3$ of propane react according to the equation below?



Solution



Volume of the propane before the reaction = $20cm^3$

The volume that reacted = $4cm^3$

Volume that did not react = volume before

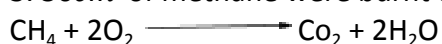
Reaction – volume that reacted i.e. $20 - 4 = 16cm^3$

EVALUATION

1. State Gay –Lussac's law

2. State Avogadro's law.

3. $50cm^3$ of methane were burnt completely in oxygen according to the equation below.



Calculate: (a) volume of oxygen used (b) Volume of carbon(IV) oxide produced

(c) Volume of steam produced.

WEEKEND ACTIVITY:

(a) Define Graham's law of diffusion. (b) What is mole and mole concept?

PERIOD 2: GRAHAM'S LAW OF DIFFUSION.

Second term Chemistry E-Lesson Note

This law states that, at constant temperature and pressure, the rate of diffusion of a gas is inversely proportional to the square root of its relative molecular mass or square root of its vapour density.

Mathematically, Graham's law of diffusion can be represented as:

$$\frac{R_1}{R_2} \propto \sqrt{\frac{p_2}{p_1}} \text{ Where } R_1 \text{ and } R_2 \text{ are the rates of diffusion and } P_1 \text{ and } P_2, \text{ the densities of the two gases.}$$

The density is directly proportional to its molecular mass.

EXAMPLES

1. 100cm^3 of oxygen diffuse through an office in 60 seconds while it takes 120seconds for the same office. Calculate the molecular mass of the unknown gas [O=16]

Solution

$$\frac{R}{R_x} O_2 = \frac{M_x}{MO_2}$$

Since the rate of diffusion is inversely proportional to the time taken:

$$\frac{R}{R_x} O_2 = \frac{tx}{tO_2} = \sqrt{\frac{M_x}{MO_2}}$$

$$\left(\frac{tx}{tO_2}\right) = \frac{M_x}{MO_2}$$

$$M_x = MO_2 \times \left(\frac{tx}{tO_2}\right)^2 = 32 \times \left(\frac{120}{60}\right)^2 = 32 \times 2^2$$

$$M_x = 32 \times 4 = 128\text{g}$$

2. 200cm^3 of hydrogen diffused through a porous pot in 40 seconds. How long will it take 300cm^3 of chlorine to diffuse through the same pot?

Solution

200cm^3 of hydrogen diffused in 40secs

$\therefore 300\text{cm}^3$ of chlorine will diffuse in

$$\frac{300\text{cm}^3}{200\text{cm}^3} \times 40 = 60$$

$$(3 \times 20) = 60\text{seconds}$$

Now, using the equation

$$\frac{t_1}{t_2} = \sqrt{\frac{M_1}{M_2}}$$

Where $t = 60\text{s}$, $M_1 =$ molecular mass of hydrogen

i.e $H_2 = (2 \times 1) = 2$

$M_2 =$ molecular mass of chlorine = $Cl_2 = 2 \times 35.5 = 71$

$$T_2 = t_1 \sqrt{\frac{M_1}{M_2}} = 60 \sqrt{\frac{2}{71}} = 60 \sqrt{35.5} = 60 \times 5.96$$

$$= 357.5\text{sec}$$

Time of diffusion of chlorine = 358s.

3. How many times the rate of diffusion of hydrogen is faster than that of oxygen and what law do you use to get the answer? [vapour density] of [H=1, O=16]

Solution

Rate (R_+) of diffusion of $H_2 =$

$$\sqrt{\frac{\text{Density of } O_2}{\text{Density of } H_2}}$$

$$\frac{R_1}{R_2} = \sqrt{\frac{16}{1}} = \frac{R_1}{R_2} = \frac{4}{1}$$

\therefore Hydrogen diffuses four times faster. The law used is Graham's law of diffusion.

RELATIVE VAPOUR DENSITY OF A GASE

The vapour density of a gas or vapour is the number of times a given volume of gas (or vapour) is heavier than the same volume of hydrogen measured and weighed under the same temperature and pressure

Second term Chemistry E-Lesson Note

$$\text{Vapour density} = \frac{\text{mass of 1 vol of a gas or vapour}}{\text{mass of equal volume of hydrogen}}$$

Applying Avogadro's law, it is possible to show that the vapour density of a gas is related to the relative molecular mass of the gas.

$$\text{V.D} = \frac{\text{mass of 1 mole of a gas or vapour}}{\text{mass of 1 molecule of hydrogen}}$$

$$\text{V.D} = \frac{\text{mass of 1 vol of a gas}}{\text{mass of 2 atoms of hydrogen}}$$

$$\therefore 2 \times \text{V.D} = \text{relative molecular mass}$$

The density of hydrogen at S.T.P is 0.09 dm^3

Example

Calculate the vapour densities of the following gases from the given data.

560 cm^3 of oxygen at S.T.P weighs 0.8g

$1,400 \text{ cm}^3$ of sulphur (iv) oxide weighs 4g

Solution

1000 cm^3 of hydrogen at S.T.P weighs 0.09g

$$\therefore 560 \text{ cm}^3 \text{ of hydrogen at } \frac{560 \text{ cm}^3}{1000 \text{ cm}^3} \times 0.09$$

$$= 0.05 \text{ g}$$

$$\text{V.D} = \frac{\text{mass of given volume of gas}}{\text{mass of equal volume of hydrogen}}$$

\therefore Vapour density of oxygen =

$$\frac{\text{mass of } 560 \text{ of oxygen}}{\text{mass of } 560 \text{ of hydrogen}}$$

1000 cm^3 of hydrogen at S.T.P weighs 0.09g.

\therefore 1400 of hydrogen will weigh

$$\frac{1400 \times 0.09}{1000} = 0.126 \text{ g}$$

$$\text{Vapour density} = \frac{\text{mass of given volume of gas}}{\text{mass of equal volume of hydrogen}}$$

$$\therefore \text{Vapour density of } \text{SO}_2 = \frac{\text{mass of } 1400 \text{ cm}^3 \text{ of } \text{SO}_2}{\text{mass of } 1400 \text{ cm}^3 \text{ of } \text{H}_2}$$

$$= \frac{4 \text{ g}}{0.126} = 31.74 = 32$$

EVALUATION

Deduce the relationship between relative molecular mass and vapour density of a substance.

Define vapour density of a gas.

PERIOD 3: MOLAR VOLUME OF GASES- AVOGADRO NUMBER AND THE MOLE CONCEPT

The molar volume of any gas is the volume occupied by one mole of that gas at s.t.p. and is numerically equal to 22.4 dm^3 i.e. one mole of any gas at s.t.p. occupies the same volume the value of which is 22.4 dm^3 . This value is called molecular mass or molar mass.

From the Avogadro's law, the molar volume for all gases contains the same number of molecules. This number is called the Avogadro's number or constant and the value is 6.02×10^{23} at s.t.p

MOLE: The mole can be defined as the amount of substance which contains as many elementary particles or entities e.g. ions, molecules, atoms, electrons as the number of atoms in exactly 12 grams of carbon -12.

The mole of any substance represents 6.02×10^{23} particles of any substance. Therefore, a mole refers to Avogadro's number of particles of any substance.

In summary, the molar mass of a gas contains Avogadro's number of molecules which is 6.02×10^{23} and occupies a volume of 22.4 dm^3 at s.t.p.

The atomic mass of every element also contains Avogadro's number of atoms.

Second term Chemistry E-Lesson Note

The mole concept- This says that one mole of any substance contains the same number of particles; which can be atoms, molecules or ions. This number is $6.023 \times 10^{23} dm^3$ (the Avogadro's number)

Examples

1. 158g of a gas at s.t.p. occupies a volume of $5000 dm^3$. What is the relative molecular mass of the gas? (Molar volume at s.t.p = $22.4 dm^3 mol^{-1}$)

Solution

Volume of gas: $V = 5000 dm^3$

Molar volume of gas; $V = 22.4 dm^3 mol^{-1}$

N = amount in moles

$$= \frac{V}{V}$$

$$N = \frac{50}{22.4 dm^3 mol^{-1}} = 2.23 mol$$

$$\text{Molar mass } M \text{ of the gas} = \frac{M}{n} = \frac{158g}{2.23 mol} = 70.8$$

Molar mass = $71 g mol^{-1}$

2. What is the mass of 3 moles of oxygen gas O_2 ? ($O = 16$)

Mass of 1 mole of $O_2 = (2 \times 16)g = 32g$

Mass of 3 moles of $O_2 = (3 \times 32)g = 96g$

3. How Many moles are there in 20g of $CaCO_3$? [$CaCO_3 = 100$]

Molar mass of $CaCO_3 = 100g$

100g of $CaCO_3 = 1 \text{ mole}$

20g of $CaCO_3 = \frac{20}{100} \times 1 \text{ mole} = 0.2 \text{ moles}$

EVALUATION

1. Using the relationship between mole and Avogadro's number. Define mole in six ways.

PERIOD 4: DALTON'S LAW OF PARTIAL PRESSURE

Dalton's law of partial pressure states that for a mixture of gases that do not react chemically, the total pressure exerted by the mixture of gases is equal to the sum of the partial pressures of the individual gases.

Mathematically, Dalton's law of partial pressure for a mixture of n gases can be expressed as:

$P_{\text{total}} = P_1 + P_2 + P_3 + \dots + P_n$ where P_{total} is the total pressure exerted by the mixture of gases that do not react, $P_1, P_2, P_3, \dots, P_n$ are partial pressure of the individual gases.

Example:

If $20.0 dm^3$ of hydrogen were collected over water at $17^\circ C$ and $79.7 kNm^{-2}$ pressure; Calculate the

(a) Pressure of dry hydrogen at this temperature.

(b) Volume of dry hydrogen at s.t.p.

(vapour pressure of water is $1.90 kNm^{-2}$ at $17^\circ C$)

Solution:

$$(a) P_{H_2} = P_{\text{total}} - P_{\text{water vapour}}$$

$$= 79.7 - 1.90$$

$$= 77.8 kNm^{-2}$$

$$(b) \frac{P_1 V_1}{T_1} = \frac{P_2 V_2}{T_2}$$

$$\frac{77.9 \times 20}{290} = \frac{101.3 V_2}{273}$$

$$V_2 = 14.5 dm^3$$

GENERAL EVALUATION

OBJECTIVE TEST

Second term Chemistry E-Lesson Note

1. A liquid begins to boil when (a) Its vapour pressure is equal to the vapour pressure of its solid at the given temperature (b) Molecules start escaping from its surface (c) Its vapour pressure equals the atmosphere pressure (d) Its volume is slightly increased.
2. Hydrogen diffuses through a porous plug (a) At the same rate as oxygen (b) Twice as fast as oxygen (c) Three times as fast as oxygen. (d) Four times as fast as oxygen
3. When pollen grains are suspended in water and viewed through a microscope, they appear to be in a state of constant but erratic motion. This is due to: (a) Convection current (b) small change in temperature (c) a chemical reaction between the pollen grains and the water (d) the bombardment of the pollen grain by molecules of water.
4. If the quantity of oxygen occupying a 2.76litre container at pressure of 0.825 atmosphere and 300k is reduced by one-half, what is the pressure exerted by the remaining gas? (a) 1.650atm (b) 0.825atm (c) 0.413atm (d) 0.275atm
5. 200cm^3 of oxygen diffused through a porous plug in 50secs. How long will 80cm^3 of methane (CH_4) take to diffuse through the same porous plug under the same conditions (C= 12, O= 16, H=1) (a) 40sec (b) 20sec (c) 14sec (d) 7sec

ESSAY QUESTIONS

1(a) State Graham's law of diffusion

Arrange the following gases in decreasing order of diffusion rate: Chlorine, hydrogen chloride, hydrogen sulphide and Carbon (IV) oxide

[H=1, C= 12, O=16, S= 32, Cl=35.5]

2. (a) What do you understand by s.t.p?

(b) If the volume of a given mass of gas at 298k and pressure of $205.2 \times 10^3 \text{ Nm}^{-2}$ is 2.12dm^3 , what is the volume at S.T.P? Standard pressure= $101.3 \times 10^3 \text{ Nm}$. Standard temperature= 273k

3. (a) Calculate the number of moles of the following at s.t.p

i. 16g of oxygen

ii. 67.2dm^3 of nitrogen gas, and

iii. 1.14dm^3 of hydrogen chloride gas.

O=16, H=14, N=1. Molar volume of gas at S.T.P = 22.4dm^3

(b) (i) Convert 33°C and -41°C to Kelvin scale

(ii) Convert 270k and 315k to 0°C

PRE READING ASIGNMENT:

Read about standard separation techniques.

WEEKEND ACTIVITY:

List all the separation techniques that you know.

WEEK 10

Experiment to illustrate the laws

WEEK 11

ATMOSPHERE AND ENVIRONMENT

Air: Air is mixture of gases in the atmosphere and it consists of nitrogen, oxygen, carbon (iv)oxide, noble gases, water vapour and gaseous impurities. The most active constituent of air is oxygen. The atmosphere is a layer of air containing mixture of several gases. This mixture composition varies according to time and place. The composition of water vapour varies from 0 - 5%, depending on the humidity of air.

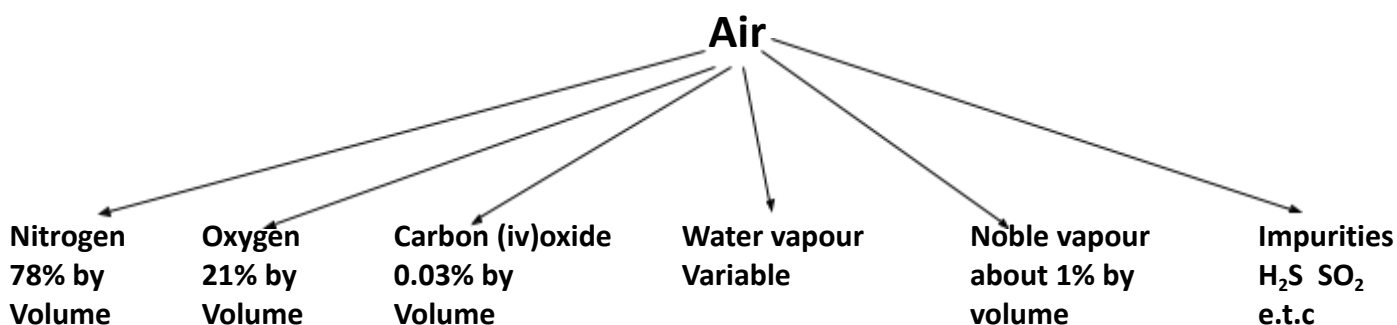
Air is a mixture of gases composed mainly of nitrogen and oxygen with small amount of carbon (IV) Oxide, noble gases and water vapour. The following are the percentage composition of the air;

-Nitrogen – 78%

-Oxygen – 21%

Second term Chemistry E-Lesson Note

- Carbon (IV) oxide – 0.03%
- Noble gases – about 1%
- Water vapour – variable
- Dust and others – variable



Properties of air

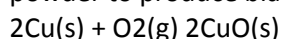
- It is colourless and odourless
- It fills the space in which it is placed
- It can be compressed
- Its density is about 1.2kgm^{-3} .
- It exerts a pressure known as atmospheric pressure
- It supports burning and respiration because of the presence of oxygen

Evidence that air is a mixture

1. The constituents of air can be separated easily by physical methods.
2. The constituents of air still remain their individual properties.
3. The composition of air cannot be represented by a chemical formulae as in the case of a compound.
4. If the appropriate proportions of each of the different constituents of air are mixed together under ordinary condition, there will be no evidence of any chemical combination such as heat production or volume change.

Percentage Composition of Oxygen in Air

A known volume of air is passed through tube with burning copper powder and oxygen in air will react with hot copper powder to produce black copper oxide

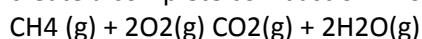


If oxygen is depleted, the readings on both syringes will be steady and the reaction has completed. Hence, to find the volume of oxygen in air collected in syringe Volume of O_2 = Initial volume of air – Final volume of air Example, the initial volume of air in one syringe is 80cm^3 and the final volume is 64cm^3 . Hence, the percentage volume of O_2 in air is

$$\begin{aligned} \% \text{ Volume of } \text{O}_2 &= 80 - 64 = 16\text{cm}^3 \\ &= 16/80 \times 100\% = 20\% \end{aligned}$$

Liquefaction of Air: Firstly, CO_2 is removed by passing air through NaOH . Then, the air is cooled to -25°C to freeze water vapour to be removed and the remaining air is cooled and compressed to become liquid which is then separated into its constituents by fractional distillation as each constituent has different boiling point.

Combustion MOST substances react with O_2 in an exothermic reaction, which is called combustion. If flames are produced during combustion, it is called burning. ALL carbon compounds burn in O_2 to produce CO_2 while ALL hydrogen containing compounds burn in O_2 to produce H_2O . When adequate supply of oxygen is available during burning; it will create a complete combustion. If otherwise, the combustion is incomplete



Typical Combustion Reaction

When air hole is closed, oxygen cannot enter reaction chamber, and hence soot (unburnt carbon) and CO is produced from incomplete hydrocarbon gas combustion. As a result, flame is yellow due to glowing specks of hot soot in heat and the flame is not hot. When air hole is opened, air supplies plenty of oxygen, allowing complete combustion.

Oxygen in combustion Reaction used as

Second term Chemistry E-Lesson Note

1. As rocket fuel
2. In steel making, to burn off impurities
3. In oxy-acetylene cutting and welding
4. In oxygen tanks for deep sea divers and mountain climbers to provide oxygen
5. For respiration for most animals
6. Used as oxygen tents in hospital to aid patients with respiratory problems

FLAMES

A flame is a region where gases combine chemically, with the production of heat and light. Flames are produced when substances burn and the structure varies according to the chemical composition of the gas which is burning. Flames are not homogeneous, but are composed of several defined zones. A flame may be luminous or non-luminous.

The burning or combustion of substances in air gives out light and heat. The reaction often produces Flame. A flame consists of several zones. It may be luminous or non-luminous.

(a) Non-luminous flame (e.g. Bunsen burner flame)

The flame consists of 3 zones:

- Dark base cone which is cold and contains unburnt gas and air.
- The blue inner cone which is the zone of incomplete combustion containing unburnt gas, Carbon (II) oxide and carbon
- Almost colourless outer cone near the tip of the inner cone which is the zone of complete Combustion and the hottest part of the flame.

Fig. Luminous and non-luminous flame

(b) Luminous flame: it is produced when there is limited supply of air e.g. when the hole of the Bunsen Burner is closed. There are four zones;

- The dark base zone of unburnt gas
- Zone of incomplete combustion found in the orange inner cone
- The almost colourless outer zone where complete combustion occur. The tip of this cone is the Hottest part of the flame.
- The blue zone of partial combustion at the base of the flame.

Types of Flame

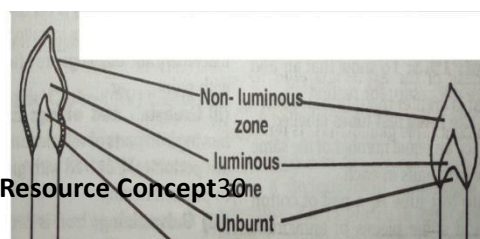
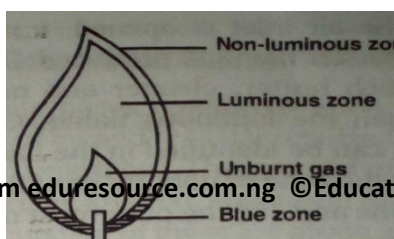
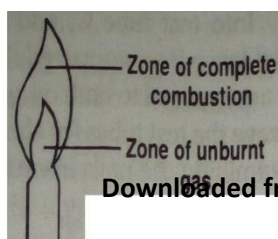
1. Hydrogen Flame
2. Candle Flame
3. Bunsen Flame

HYDROGEN FLAME: Hydrogen flame burns with a pale – blue, non-luminous flame. The structure of the hydrogen flame is simple and consists of only two zones: zone of unburnt gas and zone of complete combustion.

CANDLE FLAME: A candle flame burns with a luminous flame. Four zone can be identified in the flame

- (i) The zone of unburnt gas around the wick
- (ii) The bright yellow luminous zone where there is incomplete burning of the hydrocarbon, due to insufficient supply of air.
- (iii) The luminous zone on the outside, where there is complete combustion of the carbon particle due to the plentiful supply of air from the atmosphere.
- (iv) The blue zone at the base of the flame which is also a region of complete combustion, i.e. region of rapid burning.

BUNSEN FLAME: A Bunsen burner is a device with air inlet at the base of the burner tube to ensure more complete combustion of the fuel gas, by introducing a supply of stream of air to the flame together with the fuel gas. When the air inlet at the base of the burner tube is closed, a luminous Bunsen flame is produced.



Second term Chemistry E-Lesson Note

HYDROGEN FLAME

CANDLE FLAME

BUNSEN FLAME

Air pollution

Pollution is the release of toxic or harmful substances into the environment by the natural forces, man and other animals to the extent that causes damage to living things and their resources.

Air pollution is the contamination of the atmosphere by substances in concentrations that are harmful to human beings, plants and animals.

The unwanted substances that cause pollution are called **pollutants**.

Causes of air pollution

The main cause of air pollution is the combustion of fuels such as coal, petrol, gasoline and gases. Other causes include the release of freons from aerosol cans and poisonous gas from chemical warfare.

Air pollutants

The main air pollutants include;

- Particulate matters (tiny solids)
- Oxides of carbon
- Oxides of sulphur and nitrogen
- Gaseous hydrocarbons and
- Chlorofluorocarbons

a) Particulate matters: they are solid pollutants emitted by industrial processes, burning of coal and wood fires. They include smoke, soot and dust. They cause damage to the respiratory system if breathed in as air.

b) Oxides of carbon: carbon (II) oxide produced by the incomplete combustion of automobile fuel is Poisonous when inhaled. It competes with oxygen for the haemoglobin in the blood to form carboxyhaemoglobin.

Carbon (IV) oxide, when in excess in the atmosphere results in a greater retention of infra

– Red ray (heat ray) radiated from the earth. This causes the greenhouse effect – a gradual warming of The planet.

c) Oxides of sulphur and nitrogen: sulphur (IV) oxide is produced through the combustion of sulphur Containing coal and cracking of petroleum from industrial plant. Oxides of nitrogen are generated by Automobiles and electricity generating stations. Both oxides of sulphur and nitrogen dissolve in water to Produce acids. This results in the fall of acid rain which is harmful to plants and animals, building and Metal structures. When breathed in, these oxides cause irritation of the nose, throat and respiratory Tissues.

d) Gaseous hydrocarbons: Hydrocarbons are released into the air through deforestation, decomposition of organic matters, evaporation of organic solvents and incomplete combustion of coal, oil and wood. Hydrocarbons cause cancer and other hazards. It also reacts with oxides of nitrogen to produce smog Which reduces visibility.

e) Chlorofluorocarbons (CFCs): they are group of man – made chemicals such as Freon which are used as Propellants for spray cans, coolants in refrigerators and air – conditioners and for making plastic foams. CFCs cause the depletion of ozone layer through the following process; in the upper atmosphere, the ultraviolet light breaks off a free chlorine atom from CFCs molecule. The free chlorine atom is very Reactive and attracts a molecule of ozone to produce an oxygen molecule and an unstable oxide of Chlorine, ClO. This oxide reacts with atomic oxygen to form oxygen molecule and release the chlorine Atom which then attacks another ozone molecule.

Second term Chemistry E-Lesson Note

Ozone layer protects the earth surface from excessive ultraviolet rays which causes an increase in Cataracts, skin cancer and sun burns.

Control of air pollution

1. Strict laws should be passed to make factories and users of motor vehicles comply with antipollution regulations.
2. People should be educated on the causes and dangers of pollution.
3. Machines should be improved so that efficient combustion of fuels can be ensured.

Corrosion of metals

Corrosion occurs from the combined actions of atmospheric oxygen and water. It is aided by the Presence of carbon (IV) oxide and other gaseous pollutants in the air. The corrosion of iron is known as Rusting. Iron rusts by reacting with oxygen in the presence of water to form brown hydrated iron (III) Oxide, $\text{Fe}_2\text{O}_3 \cdot \text{xH}_2\text{O}$ which is soft and breaks off easily. Prolong exposure of iron to water results in complete rusting.

Prevention of rusting

- Painting: the use of paints or tars to protect the iron e.g. iron installed in machineries, vehicles and buildings.
- Greasing and oiling: application of grease or oil on exposed or moving parts of a machine or vehicle.
- Galvanizing: by coating the iron with zinc.
- Electroplating: the use of electrolytic process to coat iron with other metals like copper, chromium, tin, gold or silver.
- Enamelling: enamel is a glass – like substance which is put on metals (e.g. cooking pot) to protect and decorate it.

Rusting and burning

1. Both are oxidation processes i.e. they require oxygen. However, while water is needed for rusting, burning is difficult in the presence of water.
2. Rusting is a very slow process while burning is rapid.
3. Both processes release heat but while burning releases detectable, large amount of heat within a short time, rusting produced undetectable small amount of heat over a long period of time.

Atmospheric carbon(IV) oxide

Carbon(IV) oxide constitutes about 0.03% of atmospheric air. Its proportion in the atmosphere is kept fairly constant by a balance between the processes which remove carbon(IV) oxide from the atmosphere (photosynthesis) and those which return it to the atmosphere (respiration) in living organism. Respiration is a catabolic process (breaking down) while photosynthesis is an anabolic (building up) process.

Assignment

Answer question 5 on page 246 of the Essential Chemistry

Tutorial questions

1. Differentiate between luminous and non-luminous flame
2. Compare and contrast; rusting and burning
3. Discuss the ways by which rusting can be prevented/controlled
4. Define air pollution
5. Describe any five air pollutants
6. Name three air pollutants that can be generated by burning coal
7. State three reasons why air is regarded as a mixture