**ATOMIC STRUCTURE**.

ATOMIC NUMBER, MASS NUMBER, ISOTOPES AND RELATIVE ATOMIC MASS

The nucleus: The nucleus is at the centre of the atom and contains the proton and neutrons. Protons and neutrons are collectively known as **nucleons**. Virtually, all the mass of the atom is concentrated in the nucleus, because the electrons weigh so littles.

**Atomic number (z):** This is the number of protons in the nucleus of an atom. In an atom, the number of protons is equal to the number of electrons. The atomic number is also given a more descriptive name of proton number.

**Mass number (A):** is the sum of the number of protons and neutrons in the nucleus of an atom. Number of protons + Number of neutrons = mass number of the atom. The mass number is also called the **nucleon number**. This information can be given simply as illustration below;

A particular element, X can be depicted as **A Z X** where the superscript **A** is the nucleon number and the subscript is **z**, which is the atomic number.

Mathematically, **A = Z+N**

Example fluorine atom, **199 F** (how many protons, and neutrons has this atom got?) Atomic number (z)=9; mass number(A)=19 A=Z+N A-Z=N  **19-9=10 Neutrons.**

**DIFFERENCES BETWEENPROTONS AND ELECTRONS.**

|  |  |
| --- | --- |
| **PROTONS** | **ELECTRONS** |
| Have positive charge | have negative charges |
| Protons reside inside nucleus  | Electrons reside outside the nucleus |
| Protons do not take part in bonding  | Electrons are used in bonding  |
| the relative mass is 1 unit | the mass is negligible |

**Worked examples**

1. The atomic number of an atom Y is 17 and the mass number is 35. Indicate the number of (i)protons ( ii) electrons (iii) neutrons in the atom.

(i)The number of protons is 17 (because the number of protons is equal to the same as the atomic number.

(ii)The number of the electrons is 17 (because the number of protons is equal to the number of electrons in a neutral atom.

(iii)The number of Neutron(N)=mass number(A)- Atomic number (2)

 N=A-Z

 N=35-17=18

1. An atom has 17 protons and 11 neutrons in its nucleus. Find the ff. for the atom (i) mass number 28 (ii)number of electrons 17.

**ISOTOPES**: are defined as atoms of the same element having different mass numbers due to differences in number of the neutrons of the atoms. **OR** atom of the same element with the same number of protons but different number of neutrons.

**Examples;** chlorine exists naturally with 2 isotopes namely chlorine -35 or **3517 Cl** and chlorine -37 **3717 Cl**.

In **3517 Cl** atom there are 17 protons and 18 neutrons so that the atom has a mass number of 35. In **3717 Cl** atom, there are 17 protons and 20 neutrons so that this atom has a mass number of 37. Other common elements which have isotopes are:

Oxygen [168O, 178O, 188O]

Carbon [126C, 136C, 146C]

Hydrogen [11H, 21H ,31H]

Magnesium [2412Mg, 2512Mg, 2612Mg]

**THE ELECTRONIC CONFIGURATION.**

The arrangement of electrons in the shells of an atom is called electronic configuration. A shell is the path through which an electron moves or energy level of electrons. Every nucleus has shell around it and every shell has a name assigned to it. The shell with the lowest energies is nearest to the nucleus and those with largest energy are further away from the nucleus.

 **The shell and their capacities**

* K shell is the first shell. It can take a maximum of 2 electrons.
* L shell is the second shell. It can take maximum of 8 electrons.
* M shell is the third shell can take a maximum of 18 electrons.

For stability, the M shell can contain a maximum of 8 electrons instead of 18. This exception is according to a rule called **octet rule.**

* N shell, is the fourth shell and takes a maximum of 32 electrons.

**Configuration of the first 20 elements of the periodic table.**

|  |  |  |  |  |
| --- | --- | --- | --- | --- |
| Element/symbol | Atomic number | Mass number | No. of P N E | Electronic configuration |
| Hydrogen, H | 1 | 1 | 1 0 1 | 1 |
| Helium, He | 2 | 4 | 2 2 2 | 2 |
| Lithium, Li | 3 | 7 | 3 4 3 | 2, 1 |
| Beryllium, Be | 4 | 9 | 4 5 4 | 2, 2 |
| Boron, B | 5 | 11 | 5 6 5  | 2, 3 |
| Carbon, C  | 6 | 12 | 6 6 6 | 2, 4 |
| Nitrogen, N | 7 | 14 | 7 7 7 | 2, 5 |
| Oxygen, O | 8 | 16 | 8 8 8  | 2, 6 |
| Fluorine. F | 9 | 19 | 9 10 9 | 2, 7 |
| Neon, Ne | 10 | 20 | 10 10 10 | 2, 8 |
| Sodium, Na | 11 | 23 | 11 12 11 | 2, 8, 1 |
| Magnesium, Mg | 12 | 24 | 12 12 12 | 2, 8, 2 |
| Aluminum, AL | 13 | 27 | 13 14 13 | 2, 8, 3 |
| Silicon, Si | 14 | 28 | 14 14 14 | 2, 8, 4 |
| Phosphorus, P  | 15 | 31 | 15 16 15 | 2, 8, 5 |
| Sulphur, S  | 16 | 32 | 16 16 16 | 2, 8, 6 |
| Chlorine, Cl | 17 | 35 | 17 18 17 | 2, 8,7 |
| Argon, Ar | 18 | 40 | 18 22 18 | 2, 8, 8 |
| Potassium, K | 19 | 39 | 19 20 19 | 2, 8, 1 |
| Calcium, Ca | 20 | 40 | 20 20 20 | 2, 8, 2 |

 **Relative Atomic Mass (Ar)**

Is **ratio** of the average mass of atoms of an element (from a single given sample or source) to 1/12 of the mass of an atom carbon – 12 (known as the unified atomic mass unit). It is also known as atomic weight. **OR Ar** of an element is the number of times an atom is heavier than one twelfth the mass of one atom of carbon-12.

Thus, Relative atomic mass (Ar) = $\frac{mass of one atom of an element}{\frac{1}{12}of the mass of the carbon-12 atom}$

Owing to the existence of isotopes, atomic masses are usually not whole numbers. Actually, the atomic mass of an element is the weighted average of the atomic masses of all isotopes of the element. Hence, we may define **relative atomic mass (Ar)**, of an element as the weighted average mass of the various isotopes of an element compared with one twelfth the mass of an atom of carbon-12 isotopes. Some elements have relative masses which are whole numbers. Eg; Ar of oxygen is 16.0, AL 27.0. The **Ar** of these two elements are the same as their mass numbers.

**Illustration**

Carbon has 2 main isotopes, carbon 13 $$ and carbon 12 $$. Their relative abundance are 11.11% and 98.89% R.P.T Hence, Ar of carbon s expressed as

 $\frac{13×1.11}{100}$ × $\frac{12×98.89}{100}$ = 12.01

Chlorine has 2 isotopes, chlorine – 35 $$ and chlorine-37 $$. Hence, Ar = $\frac{35×3}{4}$ + $\frac{37×1}{4}$ =$\frac{142}{4}$ = 35.5

**NB;** The Ar mass of an element is a pure number and it does not have a unit.

**MOLE, MOLAR MASS AND FORMULA MASS.**

**The mole concept.**

**Mole as a quantity**. The mole is defined as the quantity of given substance that contains as ֒many elementary entities as there are atoms n exactly 12g of carbon – 12 isotopes.

**Mole as a unit;** is a unit that is used to describe a certain quantity of particles. The mole = 6.02 × $10^{23}$ particles or entities. The term particles refers to either atoms, protons, electrons, ions, molecules or formula unit. The number 6.02 × $10^{23}$ is called **Avogadro’s constant** or **Avogadro’s number**( $N\_{A}$ or L).

**MOLAR MASS(M):** The mass of one mole of a substance is the molar mass. The unit of molar mass is g/mol.

Eg; 1 mole of carbon has a mass of 12g, hence the molar mass carbon is 12g/mol.

Eg; calculate the molar mass of

i.H2SO4  ii. Nacl iii. CuSO4 iv. C12H22O11  v.Ammonia(NH3)

Relative atomic masses[H=1, S=32, O=16, Na=23, Cl=35.5, Cu=63, C=12, N=14]

1. H2SO4= (1×2)+(32×1)+(16×4)

 = 2+32+64= 98g/mol

**Formular mass;** defined as the mass of 1 mole of a formula unit, especially of ionic compound.

Eg;

Formular mass of KNO3= (39×1)+(1×14)+(3×16)= 101g

**؞** This means that 1 mole of KNO3 has a mass of 101g.

Formular mass of CaSO4 = (1×40)+ (1×32)+ (4×16) = 136g

This means 1mole of CaSO4  has a mass of 136g or the mass of 1mole of CaSO4 = 136g

**Relation between Grains and Mole.**

Mathematically, $n=\_{M}^{m}$

Where, n= amount of substance

m = mass

M = molar mass

Eg1, calculate the mass of 0.5 mole of ion [Ar of Fe = 56.0]

n= 0.5

M = 56.0

But $n=\_{M}^{m}$

m = n×M = 0.5 mol × 56g/mol = 28g

2. How many moles are there in 50g 0f NaOH? (1.25mol)

 [Na = 23, O = 16, H=13

**PREPARATION OF SOLUTIONS OF GIVEN CONCENTRATONS.**

**Solution;** is a homogenous mixture of 2 or more substances. A solution contains a solute and solvent.

**Solvent;** is the substance of the solution that dissolves the other component. It is often present in greater amount.

**Solute;** is the substance of the solution that is being dissolved in the solvent. It is generally in small amount.

Eg; brass, aqueous ethanoic acid solution, Air.

**CONCENTRATION OF SOLUTION.**

Concentration expresses the quantity of solute in a given quantity of solvent or solution. There are a number of ways to express the relative amounts of solute and solvent in solution. These are; Molarity/amount of substance concentration, Mass concentration, percent composition(by mass), parts per million.

**MOLARITY;** is the amount of solute in exactly one litre of a solution. The symbol of molarity is “C”. The unit is mol/dm3 or ‘M’.

Molarity(C)= $\frac{amount of solute(n)}{volume of solute is litres(dm^{3})(V)}$

Eg1; what is the concentration of a solution prepared by dissolving 6.0g of Nacl to make 500cm3 of solution

C= $\frac{n}{v}$

n= $\frac{m}{M}$ = $\frac{6.0g}{58.5g/mol}$ = 0.103mol

= c = $\frac{0.103mol}{0.5dm3 }$ = 0.206mol/dm3

2. What is the mass of NaOH contained in 1dm3 of 0.2M of NaOH . solution [Na= 23, 0= 16, H=1] 8g

3. What is the volume of 0.5mol HCL solution with a concentration of 2M? V= 0.25dm3.

**Mass concentration ;** Mass concentration of a solution is expressed as the mass of a solute in grams(g) per unit volume (in dm3)of a solution.

Mass concentration, P = $\frac{moles of solute(n)}{volume of solution in litres(v)}$

= $\frac{g}{dm^{3}}$

Eg; If 100g of sucrose is dissolved in 100dm3 of solution, the mass concentration of the solution is given by

Mass concentration, p, = $\frac{moles of solute}{volume of solution}$ = $\frac{100g}{100dm^{3}}$

**Percentage by mass;** This is considered in 2 ways;

* Parts of solute per 100 parts of solution
* Fraction of a solute in solution multiplied by 100.

Percentage by mass = $\frac{moles of solute}{volume of solution}$ × 100%

Eg1; If 100g of salt is present in 400g of solution, then the conc. of the solution is expressed by

C = = $\frac{moles of solute}{volume of solution}$ × 100% = = $\frac{100g}{400g}$ × 100%

 2. A solution is prepared by dissolving 50g of glucose in 450g of H20. Cal the conc. of the solution 10%

**Concentration of a solution in parts per Million(ppm);** The concentration of a solution can be expressed in terms of parts per million(ppm). The conc. is parts per million is used when the solution is highly diluted.

**PREPARATION OF A STANDARD SOLUTION.**

A standard solution is a solution whose conc. is accurately known

m = c x M x V

**Steps involved in preparation standard solutions**

1. Calculate the molar mass of the substance whose solution you want to prepare.
2. Work out the mass of the substance whose standard solution you need to prepare.
3. Weigh the calculated mass on the balance.
4. Transfer the solution into the required standard volumetric flask.
5. Rinse the beaker with enough distilled water and add to the flask and its content.
6. Top it up with distilled water to the mark.
7. Insert the stopper and shake it well to obtain a uniform solution.
8. Label the solution you have prepared.

**DILUTION OF SOLUTION.**

To dilute a solution means to add more solvent without the addition of more solute.

The rules of solute before dilution = moles of solute after dilution

 n concentration of solution = n dilution of solution

but n = c x v

c1 v1 = c2 v2

examples,

1. 53.4cm3 of a 1.50M solution of conc. HCL is on hand, but you need some 0.800M solution. How many mL of 0.800M can you make.

Solution

Using the dilution equation C1 V1 = C2 V2

1.50mol/dm3 x 53.4cm3 = 0.800mol/dm3

1. 100.0cm3 of 2.500mol/dm3 KBr solution is on hand. You need 0.5500mol/dm3. What is the find volume of solution which results? (454.5cm3)