ATOMIC AND MOLECULAR MASSES

Atoms are too light and small to be weighed individually. The mass of an atom, therefore, is expressed with respect to a standard or reference, when it is called the relative atomic mass or, simply, the atomic mass.

The relative atomic mass of an element is a number which shows how many times an atom of the element is heavier than an atom of a reference element.

Thus, the atomic mass should never be confused with the absolute mass of an atom, which can only be calculated but not directly determined.

Depending upon the reference element, different scales of atomic mass of an atom time to time. All of them contributed significantly to the development of chemistry.

The Hydrogen Scale

Hydrogen, being the lightest element, was first chosen as the reference, and the atomic mass was defined as follows.

The atomic mass of an element is the number of times an atom of element is heavier than an atom of hydrogen.

$$\text{Atomic mass} = \frac{\text{mass of 1 atom of the element}}{\text{mass of 1 atom of hydrogen}}$$

Thus, on the hydrogen scale, a hydrogen atom is assigned a mass of exactly 1 and the masses of the atoms of other elements are determined accordingly.

The Oxygen Scale

From Gay-Lussac's law of combining volumes, we know that one atom of O combines with two atoms of H to form a molecule of water (steam). We also know that 16 parts by mass of oxygen combine with two parts by mass of hydrogen to form 18 parts by mass of water. So that atomic mass of O may be treated as 16.

Considering the greater reactivity of oxygen than hydrogen, chemists shifted the reference from H = 1.000 to O = 16.000. On this scale, an atom of O is granted a mass of exactly 16 and the masses of the other atoms are determined accordingly.

However, when isotopes (atoms of the same element differing in mass number) were discovered, the oxygen scale became incontinent. It was discovered that a natural sample of oxygen has three isotopes = ^{16}O, ^{17}O and ^{18}O with abundances of 99.759, 0.037 and 0.204% respectively. Hence an atom of natural sample of oxygen as a reference has no significance.

Aston, therefore, proposed a scale based on oxygen – 16 (^{16}O = 16.000) as the standard. This scale required the revision of all atomic-mass data compiled earlier on the natural-O scale.

From Aston's work, it was clear that, instead of a natural sample of an element, only a definite isotope could be chosen as a standard. And, for convenience, the choice of the isotope should be such that minimal correction is required in the previously determined atomic masses.

The Carbon-12 Scale

Carbon has three isotopes - ^{12}C, ^{13}C and ^{14}C – of which ^{14}C is present in negligible amounts in natural samples. ^{12}C and ^{13}C have natural abundances of 98.89 and 1.11% respectively.

It was realised that carbon – 12 (^{12}C = 12.000) could also be chosen as a good standard. If this was done, the amount of correction required in the earlier data would be minimal. The atomic masses determined according to the natural-O scale would have to be reduced only by 0.004% to make the them consistent with the carbon-12 scale. So the carbon-12 scale was finally adopted.

On this scale, one-twelfth the mass of an atom of the isotope ^{12}C is treated as the atomic mass unit (amu), and relative atomic masses are determined accordingly. Relative atomic mass is defined as follows.

The relative atomic mass of an element is the ratio of the mass of an atom of the element to one-twelfth the mass of an atom of carbon-12.

In other words, it is a number that shows how many times an atom of an element is heavier than one-twelfth the mass of an atom of this isotope carbon-12.

$$\text{Relative atomic mass} = \frac{\text{mass of 1 atom of the element}}{\frac{1}{12} \times \text{mass of 1 atom of } ^{12}C}$$

$\frac{1}{12}$ of the mass of 1 atom ^{12}C, i.e., 1 amu = 1.66×10^{-24} g = 1.66×10^{-27} kg.

$\therefore$ mass of 1 atom of an element = relative atomic mass x 1.66×10^{-24} g.

The relative atomic mass of some important elements are given in the table.

Element	Symbol	Atomic number	Atomic mass	Element	Symbol	Atomic number	Atomic mass
Hydrogen	H	1	1.0079	Nickel	Ni	28	58.693
Helium	He	2	4.0026	Copper	Cu	29	63.546
Lithium	Li	3	6.941	Zinc	Zn	30	65.409
Beryllium	Be	4	9.0122	Gallium	Ga	31	69.723
Boron	B	5	10.811	Germanium	Ge	32	72.64
Carbon	C	6	12.011	Arsenic	As	33	74.922
Nitrogen	N	7	14.007	Selenium	Se	34	78.96
Oxygen	O	8	15.999	Bromine	Br	35	79.904
Fluorine	F	9	18.998	Krypton	Kr	36	83.798
Neon	Ne	10	20.180	Rubidium	Rb	37	85.468
Sodium	Na	11	22.990	Strontium	Sr	38	87.62

Magnesium	Mg	12	24.305	Palladium	Pd	46	106.42
Aluminium	Al	13	26.982	Silver	Ag	47	107.87
Silicon	Si	14	28.086	Cadmium	Cd	48	112.41
Phosphorus	P	15	30.974	Tin	Sn	50	118.71
Sulphur	S	16	32.065	Antimony	Sb	51	121.76
Chlorine	Cl	17	35.453	Tellurium	Te	52	127.60
Argon	Ar	18	39.948	Iodine	I	53	126.90
Potassium	K	19	39.098	Xenon	Xe	54	131.29
Calcium	Ca	20	40.078	Cesium	Cs	55	132.91
Scandium	Sc	21	44.956	Barium	Ba	56	137.33
Titanium	Ti	22	47.867	Gold	Au	79	196.97
Vanadium	V	23	50.942	Mercury	Hg	80	200.59
Chromium	Cr	24	51.996	Lead	Pb	82	207.20
Manganese	Mn	25	54.938	Bismuth	Bi	83	208.98
Iron	Fe	26	55.845	Radium	Ra	88	226
Cobalt	Co	27	58.933	Thorium	Th	90	232.04

The Gram–atomic mass or gram–atom

The gram–atomic mass or the gram–atom of an element is its relative atomic mass expressed in grams.

For example, the relative atomic mass of H is 1.008 and its gram–atomic mass 1.008 g.

Illustration 1: ***How many gram–atoms are there in 80.0 g of oxygen (A_r of O = 16.0)?***

Solution: The relative atomic mass of oxygen = 16.0

∴ the gram–atomic mass of oxygen = 16.0 g.

Given mass of oxygen = 80.0 g.

∴ the number of gram–atoms $= \frac{80.0g}{16.0g} = 5.$

Relative Molecular Mass

The relative molecular mass of a substance is the ratio of the mass of a molecule of the substance to one twelfth the mass of an atom of carbon–12.

$$\text{Relative molecular mass} = \frac{\text{mass of 1 molecule of the substance}}{\frac{1}{12} \text{ x mass of 1 atom of } ^{12}C}$$

Thus, the relative molecular mass of a substance is the number that shows how many times a molecule of the substance is heavier than an atom of ^{12}C.

The unit of molecular mass is the same as that of atomic mass (i.e., 1/12 x mass a ^{12}C atom). So the relative molecular mass of a substance – element or compound – can be easily calculated by adding the relative masses of all the individual atoms present in the molecule.

Gram–molecular mass

The gram–molecular mass of a substance is its relative molecular mass expressed in grams.

Examples

Substance	Molecular Formula	Relative molecular mass	Gram–molecular mass
1. Hydrogen	H_2	2 x 1 = 2	2 g
2. Oxygen	O_2	2 x 16 = 32	32 g
3. Ozone	O_3	3 x 16 = 48	48 g
4. Chlorine	Cl_2	2 x 35.5 = 71	71 g
5. Neon	Ne	1 x 20 = 20	20 g
6. Water	H_2O	2 x 1 + 16 = 18	18 g
7. Carbon dioxide	CO_2	12 + 2 x 16 = 44	44 g
8. Methane	CH_4	12 + 4 x 1 = 16	16 g
9. Nitric acid	HNO_3	1 + 14 + 3 x 16 = 63	63 g
10. Ethanol	C_2H_5OH	2 x 12 + 5 x 1 + 16 +1 = 46	46 g

Formula Mass

Covalent substances like HCl, CO_2 and CH_4 exist as discrete molecules, but ionic solids do not. For example, a crystal of sodium chlorine does not contain discrete molecules of NaCl; rather it contains Na^+ : Cl^- ratio is 1 : 1 and the formula is NaCl. So, the relative mass of NaCl (58.5; Na = 23, Cl = 35.5) should be called the formula mass rather than the molecular mass.

The same is the case with all other ionic solids. Several covalent substances also do not exist in the form represented by their molecular formulae. For example, water molecules to form

$(H_2O)_n$, both in the liquid and the solid state. However, the stoichiometry, i.e., the mass ratio of the elements in the compound, remains the same (1 : 8).

For all chemical calculations the formula mass is treated as the molecular mass as the stoichiometry is same in both cases. Also, the term 'molecular mass' is loosely used for formula mass.

Substance	Formula	Formula mass (amu)
1. Sodium chloride	NaCl	23 + 35.5 = 58.5
2. Calcium chloride	$CaCl_2$	40 + 2 x 35.5 = 111
3. Calcium oxide	CaO	40 + 16 = 56
4. Sodium hydroxide	NaOH	23 + 16 + 1 = 40
5. Sodium carbonate	Na_2CO_3	2 x 23 + 12 + 3 x 16 = 106
6. Calcium carbonate	$CaCO_3$	40 + 12 + 3 x 16 = 100