

TOPIC 7.

CHEMICAL CALCULATIONS I - atomic and formula weights.

Atomic structure revisited.

In Topic 2, atoms were described as ranging from the simplest atom, H, containing a single proton and usually no neutrons in its nucleus with one electron orbiting outside that nucleus, through to very large atoms such as uranium for example which contains 92 protons and even more neutrons in its nucleus. The point was also made that the mass of the atom is almost entirely located in its nucleus, being attributable to protons and neutrons, while the mass of electrons is negligible.

Amount in chemistry.

Basic principle: atoms form compounds by combining in simple numerical ratios but not in simple mass ratios.

Because the masses of atoms of different elements are different, in quantitative calculations it is not possible to say, for example, that if one atom of element A reacts with one atom of element B, then 1 gram of A reacts exactly with 1 gram of B. Thus while atoms do actually combine in simple integer ratios (like 1:1 in NaCl or 2:1 in H₂O), **they do not combine in simple mass ratios**. In the example of NaCl, the actual ratio by mass of Na : Cl in this compound is 22.99 g : 35.45 g. How can this ratio be deduced? To calculate the amounts of substances that are needed to react exactly with a given amount of other substances, a method of counting in chemistry is essential. Because a single atom or molecule is such a small amount of material, we need a more appropriate way of measuring amount than counting atoms or molecules.

To illustrate this need by analogy, consider our method of counting money. Use of cents as the basic unit of amount of money would be limited to only very cheap purchases - indeed the smallest coin now in use is the 5 cent piece. Instead, most of our transactions are in units of amount of dollars which is the basic unit used in most cases. The cent is just too small an amount to be the base unit of money. As another example, when our government engineered the change to metric units for weights and measures, they decreed the unit of length to be the millimetre. The result is very large numbers appear on drawn plans and in specifications. The latter is a good example of how the unit selected as the base unit can be too small to be convenient. Similarly, as atoms and molecules individually are exceedingly small, a very large number of atoms or molecules of elements or compounds is required as the basic unit of amount.

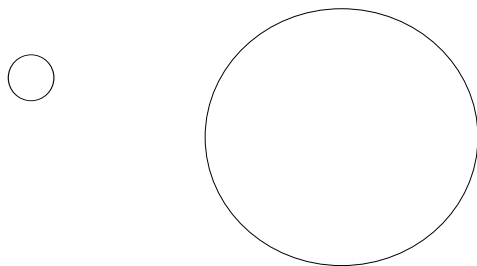
Atomic masses.

The mass of a single atom of hydrogen, ¹H, is 1.66×10^{-24} g while the mass of a single atom of the most common isotope of uranium, ²³⁸U, is 3.95×10^{-22} g. These masses are not obtained by weighing any individual H or U atoms but are deduced from the mass of a very large number of these atoms. In calculations, these very small numbers are inconvenient and, in any case, they refer to invisibly small amounts of materials. Instead, rather than the actual masses of atoms, a scale of **RELATIVE ATOMIC MASSES** is used. The most obvious way to set up a scale of relative atomic masses is to adopt the mass of the lightest atom, ¹H, as exactly 1 and express the masses of all other atoms relative to the ¹H atom.

Using this basis, the relative atomic mass of a uranium atom, ^{238}U , would be

$$\frac{\text{mass of 1 } ^{238}\text{U atom}}{\text{mass of 1 } ^1\text{H atom}} = \frac{3.95 \times 10^{-22} \text{ g}}{1.66 \times 10^{-24} \text{ g}} = \frac{238}{1} = 238$$

Comparing these two atoms,



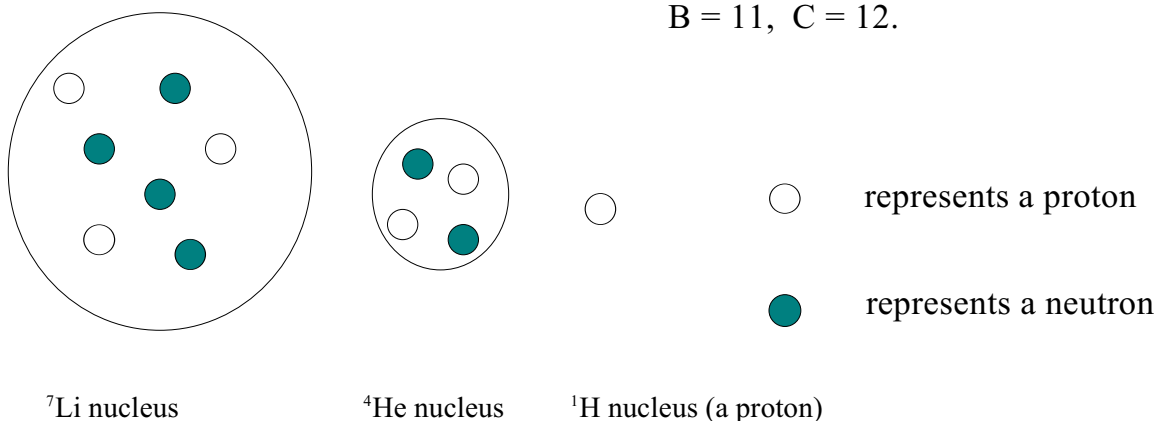
^1H atom	^{238}U atom
1 proton	92 protons + 146 neutrons
1 electron	92 electrons
1.66×10^{-24} g	3.95×10^{-22} g

Thus a single ^{238}U atom is 238 times heavier than a single ^1H atom. Similarly, the relative atomic masses or weights of atoms of all elements can be obtained and tabulated. Note that this number has no units as it is a ratio only.

As another example, consider a helium atom containing 2 protons + 2 neutrons. If one makes the approximation that a proton has the same mass as a neutron, it could be predicted that the relative atomic mass of helium would be 4. Similarly, carbon atoms with 6 protons and 6 neutrons would have a relative atomic mass = 12.

On this basis, the relative atomic weights of the most common isotopes of the first 6 elements would be

H = 1, He = 4, Li = 7, Be = 9,
B = 11, C = 12.



However, the tables of relative atomic masses, called **ATOMIC WEIGHTS** or **ATOMIC MASSES** for short, (the terms “mass” and “weight” are typically used interchangeably in this topic) apply to the **weighted average for all isotopes** for each element as it occurs in nature, not to just the most common isotope of the element. Most elements occur naturally as two or more isotopes, and therefore their relative atomic weights are not always integer quantities. [See the box below.]

**CALCULATION OF THE AVERAGE ATOMIC WEIGHT OF SILVER
from isotopic abundance ratios.**

The element silver, Ag, atomic number = 47, occurs naturally as only two isotopes: ^{107}Ag (abundance = 51.84 %) and ^{109}Ag (abundance = 48.16 %). The weighted average of the relative atomic masses of these two isotopes is calculated as follows:

$$\begin{aligned}\text{Weighted average} &= \text{isotopic mass 1} \times \text{fractional abundance 1} \\ &+ \text{isotopic mass 2} \times \text{fractional abundance 2} \\ &= 107.0 \times 0.5184 + 109.0 \times 0.4816 \\ &= 55.46 + 52.49 \\ &= 107.9\end{aligned}$$

In this calculation, the isotopic masses have been taken as 107.0 and 109.0 on the basis of the number of protons + neutrons present in each isotope, and taking the mass of a proton and a neutron as being identical. A more precise calculation and using the unit of atomic weight scale to be 1/12 the mass of a ^{12}C atom would give a value of 107.87 for the relative atomic weight of silver.

For various reasons, the basis taken for the atomic weight scale has been changed a number of times over the past century. The currently used basis is to define the unit of atomic weight scales as 1/12 of the mass of the atom of the carbon isotope with 6 protons and 6 neutrons in its nucleus, ^{12}C . The values which result for atomic weights of the elements using this basis are almost the same as that obtained by taking the atomic weight of hydrogen as 1. The following table lists the (relative) atomic weights to 1 decimal place for the first 20 elements.

Element::	H	He	Li	Be	B	C	N	O	F	Ne
Atomic mass:	1.0	4.0	6.9	9.0	10.8	12.0	14.0	16.0	19.0	20.2
Element:			Na	Mg	Al	Si	P	S	Cl	Ar
Atomic mass:			23.0	24.3	27.0	28.1	31.0	32.1	35.5	40.0

For a complete table of more precise atomic weights, see the Periodic Table at the front cover or the tear-out page at the back of this book. There is no need to rote learn values of atomic weights as precise tabulations are always available.

From the table above, it is seen for example that each He atom has four times the mass of one H atom. Consequently, it follows that in 4.0 g of helium atoms there would be the same number of He atoms present as there are atoms of hydrogen in 1.0 g of H atoms. Similarly, 6.9 g of Li atoms, 9.0 g of Be atoms or 40.0 g of Ar atoms would all have the same number of atoms present as there are in 1.0 g of H atoms. By several experimental methods, this number of atoms has been determined to be 6.022×10^{23} atoms.

Thus,	1.0 g of H atoms contains	6.022×10^{23}	H atoms
	4.0 g of He atoms contains	6.022×10^{23}	He atoms
	6.9 g of Li atoms contains	6.022×10^{23}	Li atoms

Likewise, 9.0 g of Be atoms contains 6.022×10^{23} Be atoms
 .
 40.0 g of Ar atoms contains 6.022×10^{23} Ar atoms

and for **any element**, the relative atomic weight expressed in grams (**GRAM ATOMIC WEIGHT**) is the mass of that element which contains the same number of its atoms as there are in 1.0 g of hydrogen atoms, and that number in each case is 6.022×10^{23} **atoms** of the element.

This number is called the **AVOGADRO NUMBER**, and the relative atomic mass expressed in grams for any element contains an Avogadro number of atoms of that element. The following examples illustrate these concepts.

Example 1. What mass of copper atoms contains 6.022×10^{23} Cu atoms?

As 6.022×10^{23} atoms is the number of Cu atoms in 1 gram atomic weight of Cu, and the atomic weight tables list Cu = 63.55, then 63.55 g of copper contains 6.022×10^{23} Cu atoms.

Example 2. How many neon atoms are present in 10.1 g of Ne?

From the atomic weight tables, the atomic weight of neon = 20.2.

Therefore 20.2 g of neon contains 6.022×10^{23} Ne atoms.

By proportion, 10.1 g of neon contains $\frac{10.1}{20.2} \times 6.022 \times 10^{23}$ Ne atoms
 = 3.011×10^{23} Ne atoms.

Example 3. The mass of one atom of an element Y is found to be 2.00×10^{-23} g. What is the atomic weight of element Y?

If 1 atom of Y weighs 2.00×10^{-23} g,

then an Avogadro number of atoms weighs $6.022 \times 10^{23} \times 2.00 \times 10^{-23}$ g = 12.0 g.

Therefore, the gram atomic weight of Y = 12.0 g and its atomic weight = 12.0.

The (chemical) mole.

It is convenient to have a shorter name for "an Avogadro number of atoms" and the "relative atomic mass expressed in grams" or "gram atomic weight or mass". The name adopted by chemists is the **MOLE**. When it is necessary to specify the mass of a mole of atoms in grams, the term used is **MOLAR MASS**. One mole of atoms of any element is an amount equal to the atomic weight of that element expressed in grams **and also** it is an Avogadro number of atoms of that element.

Thus 1 mole of hydrogen atoms = 6.022×10^{23} H atoms = 1.0 g of H atoms

and 1 mole of helium atoms = 6.022×10^{23} He atoms = 4.0 g of He atoms

and 1 mole of lithium atoms = 6.022×10^{23} Li atoms = 6.9 g of Li atoms

and 1 mole of atoms of any element = 6.022×10^{23} atoms of that element
 = the atomic weight of that element expressed in grams.

Thus,

$\text{moles of atoms of any element} = \frac{\text{mass of element}}{\text{atomic weight}} = \frac{\text{number of atoms of element}}{6.022 \times 10^{23}}$

Comparison of one mole of some atoms:

^1H atom	^4He atom	^{16}O atom
1 mole of ^1H atoms contains 6.022×10^{23} atoms	1 mole of ^4He atoms contains 6.022×10^{23} atoms	1 mole of ^{16}O atoms contains 6.022×10^{23} atoms
Mass per mole: 1.0 g	4.0 g (as a ^4He atom is $4 \times$ heavier than an ^1H atom)	16 g (as an ^{16}O atom is $16 \times$ heavier than an ^1H atom)
Mass per atom: 1.7×10^{-24} g	6.6×10^{-24} g	27×10^{-24} g

Relative molecular weights.

It has already been seen that most elements occur in larger aggregates rather than as single atoms. Some occur as molecules, for example hydrogen and oxygen both normally occur as the diatomic molecules H_2 and O_2 rather than as individual H or O atoms. These two elements react to form the compound water, H_2O , another molecule. The H_2 molecule must have twice the mass of a single H atom and likewise an O_2 molecule must have twice the mass of a single O atom.

The relative masses of any molecules such as these can be calculated by simply adding all the component atoms' atomic weights.

Thus the H_2 molecule has a **RELATIVE MOLECULAR WEIGHT** (usually shortened to "molecular weight") = $2 \times 1.0 = 2.0$ (using atomic weights data to one decimal place.)

The O_2 molecule has a (relative) molecular weight = $2 \times 16.0 = 32.0$

In the case of the water molecule, each H_2O molecule has a (relative) molecular weight = $2 \times$ atomic weight of H + $1 \times$ atomic weight of O

$$= 2 \times 1.0 + 16.0 = 18.0$$

Similarly, the molecular weight of carbon dioxide, CO_2 , = $12.0 + (2 \times 16.0)$

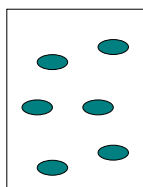
$$= 44.0$$

Because molecular weights are based on addition of atomic weights, it necessarily follows that **the molecular weight expressed in grams (gram molecular weight) for any molecular species would contain 6.022×10^{23} molecules of that species.** This is also an Avogadro number of molecules and the terms "mole" and "molar mass" are equally applicable to molecules as to atoms. Thus 1 mole of any molecular species contains 6.022×10^{23} molecules and has a mass equal to the molecular weight expressed in grams, its molar mass.

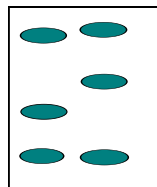
For example,

2.0 g of hydrogen contains 6.022×10^{23} H₂ molecules = 1 mole H₂ molecules
 32.0 g of oxygen contains 6.022×10^{23} O₂ molecules = 1 mole O₂ molecules
 18.0 g of water contains 6.022×10^{23} H₂O molecules = 1 mole H₂O molecules
 i.e. the molar mass of any molecular species contains the Avogadro number of molecules and is one mole of that species.

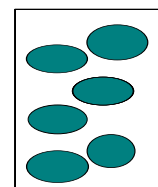
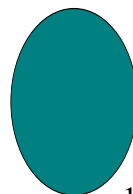
Comparison of one mole of some molecules:



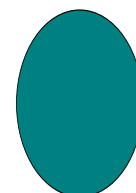
1 mole of H₂ molecules
 6.022×10^{23} H₂ molecules
 mass = 2.0 g per mole
 = 3.3×10^{-24} g per molecule



1 mole of O₂ molecules
 6.022×10^{23} O₂ molecules
 mass = 32.0 g per mole
 = 53×10^{-24} g per molecule



1 mole of CO₂ molecules
 6.022×10^{23} CO₂ molecules
 mass = 44.0 g per mole
 = 70×10^{-24} g per molecule



It was shown in earlier Topics that ionic species don't form molecules as such, and that the empirical formula of these compounds is the only formula possible.

Consequently as there are no molecules, the term "molecular weight" strictly has no meaning when applied to ionic compounds. Instead, for ionic species the term "molar mass" refers to the **GRAM FORMULA WEIGHT** where the formula is the empirical formula of the ionic compound. The gram formula weight is the sum of the constituent atomic weights of all the atoms in the empirical formula, expressed in grams. Again, the term "molar mass" is more commonly used than "gram formula weight" for these species, but the term "gram formula weight" is used in sections of these notes as it has the advantage of being self-explanatory. Gram formula weight can, like molar mass, be applied to any species regardless of whether they be individual atoms or molecules or ionic compounds or aggregates of atoms as in the case of all metals and most non-metals which are represented by just their atomic formulas.

For example, using the atomic formula for the metal magnesium, 1 mole of magnesium represented by the formula Mg, would contain 6.022×10^{23} Mg atoms and would weigh 24.3 g.

Similarly, 1 mole of the ionic compound sodium chloride, represented by the empirical formula NaCl, would have a mass = 23.0 + 35.5 = 58.5 g. This mass of sodium chloride would contain 1 mole of Na⁺ ions and 1 mole of Cl⁻ ions, i.e. 6.022×10^{23} of each ion.

This familiar container of table salt shown on the right holds 125 g of sodium chloride.

One mole of sodium chloride has a mass equal to the sum of the gram atomic weights of Na and Cl. i.e. the gram formula weight or molar mass of NaCl
 $= 23.0 + 35.5 = 58.5 \text{ g.}$

Therefore the container holds $\frac{125}{58.5}$ moles of NaCl
 $= 2.1$ moles of salt.



The following expression summarises the relationships between moles, mass and number of molecules.

$$\text{moles of substance (of specified formula)} = \frac{\text{mass of substance}}{\text{molar mass}} = \frac{\text{number of entities}}{6.022 \times 10^{23}}$$

Note that while 1 mole of hydrogen **molecules** weighs 2.02 g and it necessarily contains 6.022×10^{23} H₂ molecules, the same mass of hydrogen if converted to individual H atoms, would contain $2 \times 6.022 \times 10^{23}$ H **atoms** = 12.044×10^{23} H atoms. Thus, 1 mole of H₂ molecules contains 2 moles of H atoms bonded to form 6.022×10^{23} H₂ molecules. Similarly, there are 2 moles of O **atoms** in each mole of O₂ molecules.

This poses the question as to what is meant when referring to a mole of a given element - is it atoms or molecules of that element?

The answer is that it must be made clear in some way in the question as to which species is involved. For those few elements that occur normally as individual atoms, (the noble gases), the mole refers to an Avogadro number of atoms of each element and would have a mass equal to its gram atomic weight. For those elements that occur normally as diatomic molecules, (H₂, N₂, O₂, F₂, Cl₂, Br₂ and I₂), the mole refers to an Avogadro number of diatomic molecules and would have a mass equal to the gram molecular weight. All other elements exist normally as

larger, often indeterminate, aggregates of atoms and the mole is usually taken as referring to an Avogadro number of atoms of each element even though the element is not made up of individual free atoms. Thus, a mole of carbon is taken to mean the gram atomic weight of that element, 12.0 g, even though carbon in all its naturally occurring forms has enormous numbers of C atoms joined by chemical bonds. Similarly, all metals consist of extremely large numbers of atoms joined by metallic bonds at normal conditions, but the mole is taken to mean an Avogadro number of the atoms and the molar mass of any metal is its gram atomic weight.

Check your understanding of this section.

Why is it that when atoms of elements form compounds, they don't combine in simple integer ratios by mass?

Why are relative atomic weights of the elements usually not integers?

Give two definitions for a mole of helium.

Define the term "relative atomic weight".

How many helium atoms are present in 8.0 g of helium?

What mass of sodium chloride constitutes two moles of this compound?

Define the term "molar mass".

Calculate the mass of one mole of (i) carbon (ii) oxygen (iii) carbon dioxide.

How many moles of (i) C atoms (ii) O atoms are present in one mole of carbon dioxide?

The mole in review.

From the above, it is apparent that there are two definitions of the mole:

(i) as the formula weight of any chemical species expressed in grams.

(ii) as simply a number of entities, 6.022×10^{23} , atoms or molecules of a given chemical species.

This second definition can be extended to include things other than specific chemical species - in particular, electrons. Use will be made of this extended definition later when moles of electrons transferred in chemical reactions will be encountered in reactions where oxidation and reduction occur.

The following table summarises how to interconvert mass, moles and number of constituent formula units (atoms/molecules/empirical formulas).

mass in grams	\div molar mass	\rightleftharpoons	moles	\rightleftharpoons	number of atoms, molecules or formula units
	\times molar mass			$\times N_A$	
				$\div N_A$	

where molar mass is the gram formula weight and
 N_A is the Avogadro constant, $6.022 \times 10^{23} \text{ mole}^{-1}$

Abbreviation used for “mole”.

Units attached to numbers are typically abbreviated, such as “g” for grams as in 10.0 g or “L” for litres as in 15.0 L. For moles, the unit is abbreviated to “mol” as for example in 2.0 g = 1.0 mol of H₂. When first encountered, this might wrongly be confused as an abbreviation for molecules. There is no shorthand for “molecules”.

The following examples further illustrate the concepts discussed in this topic.

Example 1. What is the mass of 1.00 mole of carbon dioxide?

The gram formula weight (molar mass) of CO₂ = 12.0 + 2 × 16.0 = 44.0 g.

Therefore 1.00 mole would be 44.0 g of CO₂ molecules.

Example 2. What is the mass of 1.00 mole of the salt lithium fluoride?

Lithium fluoride, an ionic compound, has the (empirical) formula LiF.

The gram formula weight (molar mass) of LiF = (6.9 + 19.0) g = 25.9 g.

Therefore 1.00 mole of LiF weighs 25.9 g.

Example 3. What is the mass of 1.00 mole of C₁₂H₂₂O₁₁?

The gram formula weight (molar mass) = (12 × 12.0 + 22 × 1.0 + 11 × 16.0) g
 = 342 g

Therefore, 1.00 mole weighs 342 g.

Example 4. How many molecules are present in 0.500 mole of water?

As 1 mole of water contains 6.022×10^{23} H₂O molecules, then 0.500 mole would contain $0.500 \times 6.022 \times 10^{23}$ molecules = 3.01×10^{23} molecules.

Example 5. How many molecules are present in 8.0 g of oxygen?

Oxygen is diatomic normally so the molecular formula O_2 is assumed.

The mass of 1 mole of $O_2 = 2 \times 16.0 = 32.0$ g.

Therefore, 8.0 g of O_2 would be $8.0/32.0$ mole = 0.25 mole.

As 1 mole contains 6.022×10^{23} molecules of O_2 , then 0.25 mole contains

$0.25 \times 6.022 \times 10^{23}$ molecules of O_2

= 1.5×10^{23} molecules.

Example 6. How many molecules are present in a 2.00 g cube of table sugar having the molecular formula $C_{12}H_{22}O_{11}$.

The mass of 1 mole would be $(12 \times 12.0 + 22 \times 1.0 + 11 \times 16.0)$ g = 342 g

which contains 6.022×10^{23} molecules of $C_{12}H_{22}O_{11}$.

Therefore, 2.00 g contains $(2.00/342) \times 6.022 \times 10^{23}$ molecules of $C_{12}H_{22}O_{11}$

= 3.5×10^{21} molecules.

Example 7. How many moles of each of the constituent ions are there in one mole of sodium sulfate?

The formula of sodium sulfate, an ionic compound, is Na_2SO_4 . Thus each mole of Na_2SO_4 would contain two moles of sodium ions, Na^+ , plus one mole of sulfate ions, SO_4^{2-} .

Example 8. How many moles of each of the constituent ions are there in one mole of barium chloride and what mass of each species would be present?

The formula of barium chloride, also an ionic compound, is $BaCl_2$. Thus each mole of $BaCl_2$ would contain one mole of barium ions, Ba^{2+} , plus two moles of chloride ions, Cl^- .

To obtain the mass of each of these species present, use the atomic weights of Ba and Cl. [Note that the difference in the mass of an atom such as Ba and its cation, Ba^{2+} , is negligible, so atomic weights are used even for cations and anions.]

\therefore mass of one mole of $Ba^{2+} = 137.3$ g and

the mass of two moles of $Cl^- = 2 \times 35.5$ g = 71.0 g.

Objectives of this Topic.

When you have completed this Topic, including the tutorial questions, you should have achieved the following goals:

1. Know why atoms of different elements have different masses.
2. Understand the terms (relative) atomic mass (weight); (relative) molecular mass (weight); gram formula weight.
3. Know how to obtain the atomic weight of any element from the tables.
4. Know that the atomic weight of any element when expressed in grams contains 6.022×10^{23} atoms of that element, and that this number is called the Avogadro number.
5. Be able to calculate the number of atoms in a given mass of an element, and be able to calculate the mass of a given number of atoms of an element.
6. Understand the meaning of the terms "mole" and "molar mass".
7. Be able to calculate the molar mass of a compound.
8. Be able to interconvert a given mass of a compound, moles of the compound and number of constituent entities.
9. Recognise when the molar quantity of a compound or element refers to an empirical formula or to a molecular formula.
10. Understand that the mole when regarded as an Avogadro number of entities can apply to species other than elements or compounds, in particular, to electrons.

SUMMARY

Atoms of each element have their own characteristic masses due to the different numbers of protons and neutrons in their nuclei. Consequently, while atoms combine in simple integer ratios when forming compounds, they do not combine in simple integer ratios by mass.

Masses of individual atoms and molecules are exceedingly small, so calculating amounts that combine in reactions on the basis of actual masses of each constituent atom would involve inconveniently small quantities for the basic unit and extremely large numbers of each atom. Instead, a relative scale of atomic masses was devised wherein the mass of the atom of each element is expressed relative to that of the lightest atom, hydrogen. This basis has been altered to replace the mass of the H atom with 1/12 the mass of the carbon isotope, ^{12}C , although the resulting values are almost the same as for the hydrogen-based scale. The atomic mass of each element is calculated as the weighted average of all the naturally occurring isotopes of that element and consequently for most elements, the atomic weight is not an integer.

Because of the relative nature of the atomic weight scale, it follows that the atomic weight of any element expressed as grams must contain the same number of atoms as the atomic weight of any other element expressed as grams. This quantity is called the "gram atomic weight" and the number of constituent atoms is known as the Avogadro number, determined by experiment to be 6.022×10^{23} atoms.

It is more convenient to use the name "mole" to refer to the gram atomic weight of an element. Thus one mole of atoms of any element has a mass equal to its atomic weight expressed as grams and contains an Avogadro number of atoms.

On the same basis, compounds consisting of molecules have a characteristic relative molecular weight which is the sum of the constituent atoms' atomic weights. A mole of such a compound contains an Avogadro number of molecules and has a mass equal to the relative molecular weight expressed as grams.

A number of synonymous terms are to be found in texts. Relative atomic weight and relative atomic mass, both usually shortened to atomic weight and atomic mass are identical. Similarly, relative molecular weight and relative molecular mass, again shortened to molecular weight and molecular mass, are identical. Gram atomic weight, gram molecular weight, gram formula weight and molar mass all refer to the mass of one mole of an element or compound, i.e. the mass containing an Avogadro number of the constituent entities as defined by its formula.

The formula weight definition of the mole cannot be applied to species other than those having a defined chemical formula but the Avogadro number definition of the mole can be applied to any species, for example, electrons.

The mole of the few non-metals that usually occur as diatomic molecules such as O_2 , H_2 etc refers to their molecular formulas. However, for elements such as metals and most other non-metals which consist of extremely large aggregates of atoms, the term mole is applied on the basis of a single atom of that element even though the element occurs as enormous numbers of bonded atoms. Thus for example, a mole of carbon is taken as the amount in 12.0 grams of that element and containing 6.022×10^{23} C atoms. For non-molecular compounds such as all ionic compounds, the mole applies to the empirical formula such as NaCl for sodium chloride.

Interconversion between mass of a substance, moles of substance and number of constituent entities requires the molar mass (from atomic weight tables) and the Avogadro number (6.022×10^{23}) respectively.

TUTORIAL QUESTIONS - TOPIC 7.

1. Explain the meaning of each of the following terms:

(i) relative atomic mass.

(ii) gram atomic weight.

(iii) gram formula weight.

(iv) the mole.

(v) molar mass

2. Calculate the number of moles in each of the following:

(i) carbon (18.0 g)

(ii) argon (12.5 g)

(iii) bromine (23.6 g)

(iv) aluminium (36.4 g)

(v) water (36.0 g)

(vi) silicon tetrachloride (200 g)

(vii) sodium hydrogencarbonate (20.0 g)

(viii) iron(II) phosphate (50.0 g)

(ix) sulfur (2.80 g)

(x) chlorine (18.3 g)

(xi) ammonium sulfate (42.8 g)

(xii) potassium dichromate (94.6 g)

(xiii) copper(II) sulfate (20.4 g)

(xiv) sulfur dioxide (15.0 g)

3. Calculate the mass of each of the following:

- (i) silver (0.150 mol)
- (ii) silicon (2.00 mol)
- (iii) hydrogen (2.00 mol)
- (iv) iodine (0.100 mol)
- (v) water (2.00 mol)
- (vi) sodium chloride (3.50 mol)
- (vii) barium chloride (0.370 mol)
- (viii) sulfur trioxide (1.87 mol)
- (ix) iron(II) phosphate (0.550 mol)
- (x) potassium dichromate (1.70 mol)
- (xi) ammonium sulfate (1.70 mol)
- (xii) copper(II) nitrate (0.480 mol)
- (xiii) manganese(II) hydroxide (1.50 mol)
- (xiv) nickel(II) iodide (2.90 mol)
- (xv) lithium carbonate (5.30 mol)

4. Write the formulas for the constituent ions of the compounds below. Calculate how many moles of each ion is present in the given mass of each compound.

- (i) sodium chloride (65.0 g)
- (ii) potassium sulfide (23.5 g)
- (iii) cobalt(II) chloride (20.0 g)
- (iv) potassium oxide (50.0 g)
- (v) lithium carbonate (32.5 g)
- (vi) iron(III) phosphate (100 g)
- (vii) aluminium sulfate (50.0 g)

5. (a) (i) How many moles of carbon dioxide would be present in a sample containing 4.53×10^{24} molecules?

(ii) What would be the mass of this amount of carbon dioxide?

(b) (i) How many moles of water are present in a glass containing 7.34×10^{23} molecules?

(ii) What would be the mass of this amount of water?

(c) The mass of a crystal of rock salt, sodium chloride is 5.15 g. Calculate the number of sodium ions and chloride ions in that crystal.

(d) A sugar cube has a mass = 6.50 g. If the cube consists of pure sucrose of formula $C_{12}H_{22}O_{11}$. How many sucrose molecules are present in the sugar cube?

ANSWERS TO TUTORIAL TOPIC 7.

1. (i) Relative atomic mass: The mass of an atom relative to that of, initially, the H atom, but currently taken as relative to 1/12 of the mass of a carbon-12 atom. Being a ratio of two masses, it has no units.

(ii) Gram atomic weight: The relative atomic weight of any atom expressed as grams. Units are grams.

(iii) Gram formula weight: The sum of the atomic weights of all the component atoms in a compound, expressed as grams. Can be applied to atoms, molecular formulas or, in the case of ionic compounds, to the empirical formula. Units are grams.

(iv) Mole: The amount of a substance in an Avogadro number (6.022×10^{23}) of its constituent entities. For an element or compound, the mole is also the amount present in a mass of that substance equal to its gram formula weight, i.e. the sum of the atomic weights of all of its constituent atoms expressed as grams. Units are moles, abbreviated "mol".

(v) Molar mass: Same as gram formula weight. Can be used to refer to atoms, molecules or formula units such as the empirical formulas of ionic compounds. Units are grams.

- | | | | | | | |
|----|--------|---------------------------|--------|---------------------------|-------|---------------------------|
| 2. | (i) | 1.50 mol | (ii) | 3.13×10^{-1} mol | (iii) | 1.48×10^{-1} mol |
| | (iv) | 1.35 mol | (v) | 2.00 mol | (vi) | 1.18 mol |
| | (vii) | 2.38×10^{-1} mol | (viii) | 1.40×10^{-1} mol | (ix) | 8.73×10^{-2} mol |
| | (x) | 2.58×10^{-1} mol | (xi) | 3.24×10^{-1} mol | (xii) | 3.22×10^{-1} mol |
| | (xiii) | 1.28×10^{-1} mol | (xiv) | 2.34×10^{-1} mol | | |

Explanations and Worked Solutions.

All parts of the question involve converting mass to moles. Before this can be done, it is **essential that the correct formula for the element or compound must be used**. Then in each part, calculate the molar mass (gram formula weight) of the element or compound and divide the specified mass of the substance by its molar mass.

(i) The element carbon is represented by the symbol for its atom, C. The atomic weight of carbon is 12.01 or, its molar mass = 12.01 g mol^{-1} .
 \therefore moles of carbon in 18.0 g = $18.0 \div 12.01 = 1.50 \text{ mol}$.

(ii) The element argon is represented by the symbol for its atom, Ar. The atomic weight of argon is 39.95 or, its molar mass = 39.95 g mol^{-1} .
 \therefore moles of argon in 12.5 g = $12.5 \div 39.95 = 0.313 \text{ mol}$ or $3.13 \times 10^{-1} \text{ mol}$.

(iii) The element bromine consists of diatomic molecules and so it is represented by the molecular formula, Br_2 and not by the symbol for its atom, Br. The atomic weight of bromine is 79.90 so its molecular weight = $2 \times 79.90 = 159.8$ or its molar mass = 159.8 g mol^{-1} .
 \therefore moles of bromine in 23.6 g = $23.6 \div 159.8 = 0.148 \text{ mol}$ or $1.48 \times 10^{-1} \text{ mol}$.

(iv) The element aluminium is represented by the symbol for its atom, Al. The atomic weight of aluminium is 26.98 or, its molar mass = 26.98 g mol^{-1} .
 \therefore moles of aluminium in 36.4 g = $36.4 \div 26.98 = 1.35 \text{ mol}$.

(v) The compound water consists of molecules having molecular formula H_2O . The molecular weight of water = sum of the atomic weights of its component atoms = $2 \times 1.008 + 16.00 = 18.02$ or its molar mass = 18.02 g mol^{-1} .
 \therefore moles of water in 36.0 g = $36.0 \div 18.02 = 2.00 \text{ mol}$.

(vi) The compound silicon tetrachloride consists of molecules having molecular formula SiCl_4 . The molecular weight of silicon tetrachloride = sum of the atomic weights of its component atoms = $28.09 + 4 \times 35.45 = 169.9$ or its molar mass = 169.9 g mol^{-1} .
 \therefore moles of silicon tetrachloride in 200 g = $200 \div 169.9 = 1.18 \text{ mol}$.

(vii) The compound sodium hydrogencarbonate is ionic, consisting of Na^+ and HCO_3^- ions. Thus the formula for sodium hydrogencarbonate is NaHCO_3 . For an ionic compound, its empirical formula (i.e. the simplest whole number ratio of the constituent ions) is used as the basis for calculating the molar mass. The molar mass of sodium hydrogencarbonate = sum of the atomic weights of its component atoms = $22.99 + 1.008 + 12.01 + 3 \times 16.00 = 84.01 \text{ g mol}^{-1}$.
 \therefore moles of sodium hydrogencarbonate in 20.0 g = $20.0 \div 84.01 = 0.238 \text{ mol}$ or $2.38 \times 10^{-1} \text{ mol}$.

(viii) The compound iron(II) phosphate is ionic, consisting of Fe^{2+} and PO_4^{3-} ions. Thus the formula for iron(II) phosphate is $\text{Fe}_3(\text{PO}_4)_2$. For an ionic compound, its empirical formula is used as the basis for calculating the molar mass. The molar mass of iron(II) phosphate = sum of the atomic weights of its component atoms = $3 \times 55.85 + 2 \times 30.97 + 8 \times 16.00 = 357.49 \text{ g mol}^{-1}$.
 \therefore moles of iron(II) phosphate in 50.0 g = $50.0 \div 357.49 = 0.140 \text{ mol}$ or $1.40 \times 10^{-1} \text{ mol}$.

(ix) Sulfur is an element and is represented by the symbol for its atom, S. The atomic weight of sulfur is 32.07 or, its molar mass = 32.07 g mol^{-1} .
 \therefore moles of sulfur in 2.80 g = $2.80 \div 32.07 = 0.0873 \text{ mol}$ or $8.73 \times 10^{-2} \text{ mol}$.

(x) The element chlorine consists of diatomic molecules and so it is represented by the molecular formula, Cl_2 and not by the symbol for its atom, Cl. The atomic weight of chlorine is 35.45 so its molecular weight = $2 \times 35.45 = 70.90$ or its molar mass = 70.90 g mol^{-1} .
 \therefore moles of chlorine in 18.3 g = $18.3 \div 70.9 = 0.258 \text{ mol}$ or $2.58 \times 10^{-1} \text{ mol}$.

(xi) The compound ammonium sulfate is ionic, consisting of NH_4^+ and SO_4^{2-} ions. Thus the formula for ammonium sulphate is $(\text{NH}_4)_2\text{SO}_4$. For an ionic compound, its empirical formula is used as the basis for calculating the molar mass. The molar mass of ammonium sulphate = sum of the atomic weights of its component atoms = $2 \times 14.01 + 8 \times 1.008 + 32.07 + 4 \times 16.00 = 132.14 \text{ g mol}^{-1}$.
 \therefore moles of ammonium sulfate in 42.8 g = $42.8 \div 132.14 = 0.324 \text{ mol}$ or $3.24 \times 10^{-1} \text{ mol}$.

(xii) The compound potassium dichromate is ionic, consisting of K^+ and $\text{Cr}_2\text{O}_7^{2-}$ ions. Thus the formula for potassium dichromate is $\text{K}_2\text{Cr}_2\text{O}_7$. For an ionic compound, its empirical formula is used as the basis for calculating the molar mass. The molar mass of potassium dichromate = sum of the atomic weights of its component atoms = $2 \times 39.10 + 2 \times 52.00 + 7 \times 16.00 = 294.20 \text{ g mol}^{-1}$.
 \therefore moles of potassium dichromate in 94.6 g = $94.6 \div 294.20 = 0.322 \text{ mol}$ or $3.22 \times 10^{-1} \text{ mol}$.

(xiii) The compound copper(II) sulfate is ionic, consisting of Cu^{2+} and SO_4^{2-} ions. Thus the formula for copper(II) sulphate is CuSO_4 . For an ionic compound, its empirical formula is used as the basis for calculating the molar mass. The molar mass of copper(II) sulphate = sum of the atomic weights of its component atoms = $63.55 + 32.07 + 4 \times 16.00 = 159.62 \text{ g mol}^{-1}$.
 \therefore moles of copper(II) sulfate in 20.4 g = $20.4 \div 159.62 = 0.128 \text{ mol}$ or $1.28 \times 10^{-1} \text{ mol}$.

(xiv) The compound sulfur dioxide consists of covalent molecules having molecular formula SO_2 . The molecular weight of sulfur dioxide = sum of the atomic weights of its component atoms = $32.07 + 2 \times 16.00 = 64.07$ or its molar mass = 64.07 g mol^{-1} .
 \therefore moles of sulfur dioxide in 15.0 g = $15.0 \div 64.07 = 0.234 \text{ mol}$ or $2.34 \times 10^{-1} \text{ mol}$

3. Explanations and Worked Solutions.

All parts involve calculating the mass of a given number of moles of the specified element or compound. **Again, the starting point is writing the correct formula for that element or compound.** These calculations are all based on the principle that one mole of any element or compound has a mass equal to its molar mass (gram formula weight) which is deduced from the sum of the atomic weights of its constituent atoms.

(i) Silver is an element and its formula is taken as the symbol of its atom, Ag. The atomic weight of silver is 107.87 so its molar mass = $107.87 \text{ g mol}^{-1}$.
 \therefore the mass of 0.150 mol of silver = $0.150 \times 107.87 \text{ g} = 16.2 \text{ g}$.

(ii) Silicon is an element and its formula is taken as the symbol of its atom, Si.
The atomic weight of silicon is 28.09 so its molar mass = 28.09 g mol^{-1} .
 \therefore the mass of 2.00 mol of silver = $2.00 \times 28.09 \text{ g} = 56.2 \text{ g}$.

(iii) Hydrogen is an element which consists of diatomic molecules so its formula is taken as H_2 rather than as the symbol of the single atom, H.
One mole of hydrogen therefore refers to one mole of H_2 molecules, the molecular weight of which is $2 \times 1.008 = 2.016$.
Thus the molar mass of hydrogen = 2.016 g mol^{-1} .
 \therefore mass of 2.00 mol of hydrogen = $2 \times 2.016 \text{ g} = 4.03 \text{ g}$.

(iv) Iodine is another element which exists normally as diatomic molecules so its formula is taken as I_2 rather than as the symbol of the single atom, I. One mole of iodine therefore refers to one mole of I_2 molecules, the molecular weight of which is $2 \times 126.90 = 253.80$.
Thus the molar mass of iodine = 253.8 g mol^{-1} .
 \therefore mass of 0.100 mol of iodine = $0.100 \times 253.8 \text{ g} = 25.4 \text{ g}$.

(v) The compound water consists of molecules having molecular formula H_2O .
The molecular weight of water = sum of the atomic weights of its component atoms
= $2 \times 1.008 + 16.00 = 18.02$ or its molar mass = 18.02 g mol^{-1} .
 \therefore mass of 2.00 mol of water = $2.00 \times 18.02 \text{ g} = 36.0 \text{ g}$

(vi) The compound sodium chloride is ionic, consisting of Na^+ and Cl^- ions. Thus the formula for sodium chloride is NaCl . For an ionic compound, its empirical formula is used as the basis for calculating the molar mass.
The molar mass of sodium chloride = sum of the atomic weights of its component atoms = $22.99 + 35.45 = 58.44 \text{ g mol}^{-1}$.
 \therefore mass of 3.50 mol of sodium chloride = $3.50 \times 58.44 \text{ g} = 205 \text{ g}$

(vii) The compound barium chloride is ionic, consisting of Ba^{2+} and Cl^- ions. Thus the formula for barium chloride is BaCl_2 . For an ionic compound, its empirical formula is used as the basis for calculating the molar mass.
The molar mass of barium chloride = sum of the atomic weights of its component atoms = $137.34 + 2 \times 35.45 = 208.24 \text{ g mol}^{-1}$.
 \therefore mass of 0.370 mol of barium chloride = $0.370 \times 208.24 \text{ g} = 77.0 \text{ g}$

(viii) The compound sulfur trioxide consists of covalent molecules having molecular formula SO_3 .
The molecular weight of sulfur trioxide = sum of the atomic weights of its component atoms = $32.07 + 3 \times 16.00 = 80.07$ or its molar mass = 80.07 g mol^{-1} .
 \therefore mass of 1.87 mol of sulfur trioxide = $1.87 \times 80.07 \text{ g} = 150 \text{ g}$.

(ix) The compound iron(II) phosphate is ionic, consisting of Fe^{2+} and PO_4^{3-} ions. Thus the formula for iron(II) phosphate is $\text{Fe}_3(\text{PO}_4)_2$. For an ionic compound, its empirical formula is used as the basis for calculating the molar mass.
The molar mass of iron(II) phosphate = sum of the atomic weights of its component atoms = $3 \times 55.85 + 2 \times 30.97 + 8 \times 16.00 = 357.49 \text{ g mol}^{-1}$.
 \therefore mass of 0.550 mol of iron(II) phosphate = $0.550 \times 357.49 \text{ g} = 197 \text{ g}$.

(x) The compound potassium dichromate is ionic, consisting of K^+ and $\text{Cr}_2\text{O}_7^{2-}$ ions. Thus the formula for potassium dichromate is $\text{K}_2\text{Cr}_2\text{O}_7$. For an ionic compound, its empirical formula is used as the basis for calculating the molar mass. The molar mass of potassium dichromate = sum of the atomic weights of its component atoms = $2 \times 39.10 + 2 \times 52.00 + 7 \times 16.00 = 294.20 \text{ g mol}^{-1}$.

\therefore 1.70 mol of potassium dichromate has a mass = $1.70 \times 294.20 \text{ g} = 500 \text{ g}$.

(xi) The compound ammonium sulfate is ionic, consisting of NH_4^+ and SO_4^{2-} ions. Thus the formula for ammonium sulfate is $(\text{NH}_4)_2\text{SO}_4$. For an ionic compound, its empirical formula is used as the basis for calculating the molar mass.

The molar mass of ammonium sulfate = sum of the atomic weights of its component atoms = $2 \times (14.01 + 4 \times 1.008) + 32.07 + 4 \times 16.00 = 132.15 \text{ g mol}^{-1}$.

\therefore 1.70 mol of ammonium sulfate has a mass = $1.70 \times 132.15 \text{ g} = 225 \text{ g}$.

(xii) The compound copper(II) nitrate is ionic, consisting of Cu^{2+} and NO_3^- ions. Thus the formula for copper(II) nitrate is $\text{Cu}(\text{NO}_3)_2$. For an ionic compound, its empirical formula is used as the basis for calculating the molar mass.

The molar mass of copper(II) nitrate = sum of the atomic weights of its component atoms = $63.55 + 2 \times (14.01 + 3 \times 16.00) = 187.56 \text{ g mol}^{-1}$.

\therefore 0.480 mol of copper(II) nitrate has a mass = $0.480 \times 187.56 \text{ g} = 90.0 \text{ g}$.

(xiii) The compound manganese(II) hydroxide is ionic, consisting of Mn^{2+} and OH^- ions. Thus the formula for manganese(II) hydroxide is $\text{Mn}(\text{OH})_2$. For an ionic compound, its empirical formula is used as the basis for calculating the molar mass.

The molar mass of manganese(II) hydroxide = sum of the atomic weights of its component atoms = $54.94 + 2 \times (16.00 + 1.008) = 88.96 \text{ g mol}^{-1}$.

\therefore 1.50 mol of Manganese(II) hydroxide has a mass = $1.50 \times 88.96 \text{ g} = 133 \text{ g}$.

(xiv) The compound nickel(II) iodide is ionic, consisting of Ni^{2+} and I^- ions. Thus the formula for nickel(II) iodide is NiI_2 . For an ionic compound, its empirical formula is used as the basis for calculating the molar mass.

The molar mass of nickel(II) iodide = sum of the atomic weights of its component atoms = $58.69 + 2 \times 126.90 = 312.49 \text{ g mol}^{-1}$.

\therefore 2.90 mol of nickel(II) iodide has a mass = $2.90 \times 312.49 \text{ g} = 906 \text{ g}$.

(xv) The compound lithium carbonate is ionic, consisting of Li^+ and CO_3^{2-} ions. Thus the formula for lithium carbonate is Li_2CO_3 . For an ionic compound, its empirical formula is used as the basis for calculating the molar mass.

The molar mass of lithium carbonate = sum of the atomic weights of its component atoms = $2 \times 6.94 + 12.01 + 3 \times 16.00 = 73.89 \text{ g mol}^{-1}$.

\therefore 5.30 mol of lithium carbonate has a mass = $5.30 \times 73.89 \text{ g} = 392 \text{ g}$.

4. Explanations and Worked Solutions.

(i) Sodium chloride has the formula NaCl . Thus one mole of NaCl contains one mole of Na^+ ions and one mole of Cl^- ions.

i.e. moles of Na^+ = moles of Cl^- = moles of NaCl .

Molar mass of NaCl = sum of the atomic weights of the component atoms = $22.99 + 35.45 = 58.44 \text{ g mol}^{-1}$.

Moles of NaCl in 65.0 g = $65.0 \div 58.44 = 1.11 \text{ mol}$

\therefore moles of Na^+ ions = 1.11 mol and moles of Cl^- ions = 1.11 mol.

(ii) Potassium sulfide has the formula K_2S . Thus one mole of K_2S contains two moles of K^+ and one mole of S^{2-} ions.

i.e. moles of $K^+ = 2 \times$ moles of K_2S and
moles of $S^{2-} =$ moles of K_2S .

Molar mass of $K_2S =$ sum of the atomic weights of the component atoms
 $= 2 \times 39.10 + 32.07 = 110.27 \text{ g mol}^{-1}$.

Moles of K_2S in 23.5 g $= 23.5 \div 110.27 = 0.213 \text{ mol}$

\therefore moles of K^+ ions $= 2 \times 0.213 = 0.426 \text{ mol}$ or $4.26 \times 10^{-1} \text{ mol}$
and moles of S^{2-} ions $= 0.213 \text{ mol}$ or $2.13 \times 10^{-1} \text{ mol}$.

(iii) Cobalt(II) chloride has the formula $CoCl_2$. Thus one mole of $CoCl_2$ contains one mole of Co^{2+} and two moles of Cl^- ions.

i.e. moles of $Co^{2+} =$ moles of $CoCl_2$ and
moles of $Cl^- = 2 \times$ moles of $CoCl_2$.

Molar mass of $CoCl_2 =$ sum of the atomic weights of the component atoms
 $= 58.93 + 2 \times 35.45 = 129.83 \text{ g mol}^{-1}$.

Moles of $CoCl_2$ in 20.0 g $= 20.0 \div 129.83 = 0.154 \text{ mol}$

\therefore moles of Co^{2+} ions $= 0.154 \text{ mol}$ or $1.54 \times 10^{-1} \text{ mol}$
and moles of Cl^- ions $= 2 \times 0.154 = 0.308 \text{ mol}$ or $3.08 \times 10^{-1} \text{ mol}$.

(iv) Potassium oxide has the formula K_2O . Thus one mole of K_2O contains two moles of K^+ and one mole of O^{2-} ions.

i.e. moles of $K^+ = 2 \times$ moles of K_2O and
moles of $O^{2-} =$ moles of K_2O .

Molar mass of $K_2O =$ sum of the atomic weights of the component atoms
 $= 2 \times 39.10 + 16.00 = 94.20 \text{ g mol}^{-1}$.

Moles of K_2O in 50.0 g $= 50.0 \div 94.20 = 0.531 \text{ mol}$

\therefore moles of K^+ ions $= 2 \times 0.531 = 1.06 \text{ mol}$
and moles of O^{2-} ions $= 0.531 \text{ mol}$ or $5.31 \times 10^{-1} \text{ mol}$.

(v) Lithium carbonate has the formula Li_2CO_3 . Thus one mole of Li_2CO_3 contains two mole of Li^+ and one mole of CO_3^{2-} ions.

i.e. moles of $Li^+ = 2 \times$ moles of Li_2CO_3 and
moles of $CO_3^{2-} =$ moles of Li_2CO_3 .

Molar mass of $Li_2CO_3 =$ sum of the atomic weights of the component atoms
 $= 2 \times 6.94 + 12.01 + 3 \times 16.00 = 73.89 \text{ g mol}^{-1}$.

Moles of Li_2CO_3 in 32.5 g $= 32.5 \div 73.89 = 0.440 \text{ mol}$

\therefore moles of Li^+ ions $= 2 \times 0.440 = 0.880 \text{ mol}$ or $8.80 \times 10^{-1} \text{ mol}$
and moles of CO_3^{2-} ions $= 0.440 \text{ mol}$ or $4.40 \times 10^{-1} \text{ mol}$.

(vi) Iron(III) phosphate has the formula $FePO_4$. Thus one mole of $FePO_4$ contains one mole of Fe^{3+} and one mole of PO_4^{3-} ions.

i.e. moles of $Fe^{3+} =$ moles of $PO_4^{3-} =$ moles of $FePO_4$.

Molar mass of $=$ sum of the atomic weights of the component atoms
 $= 55.85 + 30.97 + 4 \times 16.00 = 150.82 \text{ g mol}^{-1}$.

Moles of $FePO_4$ in 100 g $= 100 \div 150.82 = 0.663 \text{ mol}$

\therefore moles of Fe^{3+} ions $= 0.663 \text{ mol}$ or $6.63 \times 10^{-1} \text{ mol}$ and
moles of PO_4^{3-} ions $= 0.663 \text{ mol}$ or $6.63 \times 10^{-1} \text{ mol}$

(vii) Aluminium sulfate has the formula $\text{Al}_2(\text{SO}_4)_3$. Thus one mole of $\text{Al}_2(\text{SO}_4)_3$ contains 2 moles of Al^{3+} and 3 moles of SO_4^{2-} ions.

i.e. moles of $\text{Al}^{3+} = 2 \times$ moles of $\text{Al}_2(\text{SO}_4)_3$ and
moles of $\text{SO}_4^{2-} = 3 \times$ moles of $\text{Al}_2(\text{SO}_4)_3$

Molar mass of $\text{Al}_2(\text{SO}_4)_3 =$ sum of the atomic weights of the component atoms
 $= 2 \times 26.98 + 3 \times (32.07 + 4 \times 16.00) = 342.17 \text{ g mol}^{-1}$.

Moles of $\text{Al}_2(\text{SO}_4)_3$ in 50.0 g $= 50.0 \div 342.17 = 0.146 \text{ mol}$

\therefore moles of Al^{3+} ions $= 2 \times 0.146 = 0.292 \text{ mol}$ or $2.92 \times 10^{-1} \text{ mol}$

and moles of SO_4^{2-} ions $= 3 \times 0.146 \text{ mol} = 0.438 \text{ mol}$ or $4.38 \times 10^{-1} \text{ mol}$.

5. Explanations and Worked Solutions.

(a) (i) One mole of any compound contains an Avogadro number of molecules.

i.e. 6.022×10^{23} molecules is one mole.

$\therefore 4.53 \times 10^{24}$ molecules is $(4.53 \times 10^{24}) \div (6.022 \times 10^{23})$ moles of $\text{CO}_2 = 7.52 \text{ mol}$

(ii) One mole of CO_2 has a molar mass = sum of the atomic weights of the constituent atoms $= 12.01 + 2 \times 16.00 = 44.01 \text{ g mol}^{-1}$.

$\therefore 7.52 \text{ mol}$ has a mass $= 7.52 \times 44.01 \text{ g} = 331 \text{ g}$.

(b) (i) One mole of any compound contains an Avogadro number of molecules.

i.e. 6.022×10^{23} molecules of water is one mole.

$\therefore 7.34 \times 10^{23}$ molecules is $(7.34 \times 10^{23}) \div (6.022 \times 10^{23})$ moles of $\text{H}_2\text{O} = 1.22 \text{ mol}$

(ii) One mole of H_2O has a molar mass = sum of the atomic weights of the constituent atoms $= 2 \times 1.008 + 16.00 = 18.016 \text{ g mol}^{-1}$.

$\therefore 1.22 \text{ mol}$ has a mass $= 1.22 \times 18.016 \text{ g} = 22.0 \text{ g}$.

(c) One mole of NaCl contains one mole of Na^+ ions and one mole of Cl^- ions.

i.e. 6.022×10^{23} Na^+ ions and 6.022×10^{23} Cl^- ions.

Molar mass of $\text{NaCl} = 22.99 + 35.45 \text{ g} = 58.44 \text{ g mol}^{-1}$

\therefore moles of NaCl in 5.15 g $= 5.15 \div 58.44 = 8.812 \times 10^{-2} \text{ mol}$

Number of Na^+ ions present = number of Cl^- ions

= number of moles of $\text{NaCl} \times$ Avogadro number

$= (8.812 \times 10^{-2}) \times (6.022 \times 10^{23}) = 5.31 \times 10^{22}$ of each.

(d) Moles of sucrose in a sugar cube containing 6.50 g of sucrose, $\text{C}_{12}\text{H}_{22}\text{O}_{11}$

$= 6.50 \div$ molar mass of sucrose

$= 6.50 \div (12 \times 12.01 + 22 \times 1.008 + 11 \times 16.00) \text{ mol}$

$= 6.50 \div 342.30 \text{ mol} = 1.90 \times 10^{-2} \text{ mol}$.

One mole of sucrose contains an Avogadro number of $\text{C}_{12}\text{H}_{22}\text{O}_{11}$ molecules.

$\therefore 1.90 \times 10^{-2}$ moles contains $(1.90 \times 10^{-2}) \times 6.022 \times 10^{23}$ sucrose molecules

$= 1.14 \times 10^{22}$ molecules