✧.* Stoichiometry

FUNDAMENTALS

Relative molecular mass (covalent substances) (M_r)

- Weighted av of the masses of 1 covalent compound formula: $\frac{1}{12}$ of the mass of 1 ^{12}C atom

Ex: Calculate the M_r in sodium sulfate, Na_2SO_4 → 2(23)+32.1+4(16) = 142.1

✦ All relative isotopic, atomic, formula and molecular masses don't have units ✦

Mass – amt of material in an object (SI unit is kg)

- Isotope carbon-12, has been selected as the standard
- An atom of ^{12}C has 12 amu
- **Atomic mass unit (amu)** is the small unit of mass used to express atomic and molecular masses (1 amu = 1.66054×10^{-27}kg)

Ex: mass of a carbon atom = $(12.011)(1.66054)(10^{-27}) = 1.993 \times 10^{-26} kg$

1. **Relative isotopic mass** – mass of 1 atom of the isotope : $\frac{1}{12}$ of the mass of 1 ^{12}C atom

- Ex: ^{24}Mg is × 2 as heavy as ^{12}C, so it has a relative isotopic mass of
- $\frac{24\ \text{amu}}{\frac{1}{2}(\text{mass of a carbon}-12\text{ atom})} = \frac{24\ \text{amu}}{\frac{1}{2}(12\text{amu})} = 24$

1. **Relative atomic mass** (A_r)

- Element's average mass of all isotopes : $\frac{1}{12}$ of the mass of 1 ^{12}C atom (Atomic mass on ptable)
- Ex: Determine the relative atomic mass of magnesium

Isotope	Isotopic Abundance	Mass × Percent	Result
Magnesium – 24	78.99%	$(24)(\frac{78.99}{100})$	18.9576
Magnesium – 25	10%	$(25)(\frac{10}{100})$	2.5
Magnesium – 26	11.01%	$(26)(\frac{11.01}{100})$	2.8626
			24.3

(6.02 x 10^{23}) Avogadro's constant

mol = $\frac{\text{atom}}{6.02 \times 10^{23}}$

amt

1 mole of any substance (amt / no.) has 6.02 x 10^{23} particles (atoms, ions, molecules)

Ex:

2 mol of O_2 molecules
= (2)(6.02)(10^{23}) O_2 molecules
= (2)(2)(6.02)(10^{23}) O atoms

Ex: 1 mol of NaCl has
6.02 x 10^{23} Na^+ ions
6.02 x 10^{23} Cl^- ions
Total amt of ions in 1 mol of NaCl: 2 mol

MOLE

↳ amt

% by mass

Relative formula mass (ionic compounds)

- Weighted average of the masses of 1 ionic compound formula: $\frac{1}{12}$ of the mass of 1 ^{12}C atom
  - add all the (A_r) of all the atoms in that compound's molecular formula

Calculate the % mass of an element in a componud $\frac{(A_r\ of\ atom)(\#\ of\ atoms)}{M_r\ of\ compound}(100)$

Ex: Calculate the percentage by mass of iron and oxygen, respectively, in iron(III) oxide. Molar mass of Fe = 55.8 g mol-1; Molar mass of O = 16.0 g mol-1

Molar mass of Fe2O3 = 2(55.8) + 3(16.0) = 159.6 $g\ mol^{-1}$

Percentage by mass of iron = $2\left(\frac{55.8}{159.6}\right)(100\%) = 69.9\%$

Percentage by mass of oxygen = $3\left(\frac{16}{159.6}\right)(100\%) = 30.1\%$

Mass of an element in a compound:
$\frac{molar\ mass\ of\ element \times \#\ of\ atoms\ in\ the\ formula}{molar\ mass\ of\ the\ compound}(mass\ of\ compound)$

Ex: Calculate the mass of copper in 103 g of copper(II) oxide.
Molar mass of Cu = 63.5 g mol-1 ; Molar mass of O = 16.0 g mol-1

Molar mass of CuO = 63.5 + 16.0 = 79.5$g\ mol^{-1}$
Mass of copper in 103 g of copper(II) oxide = $\frac{63.5}{79.5}(103\ g)$ = 82.3 g

(mass) SOLID

mass (g)
amt (mol) | molar mass (g mol^{-1})

(vol) LIQUID

Concentration (gdm^{-3} or moldm^{-3})

amt (mol)
conc (mol dm^{-3}) | vol (dm^3)

→ Mr

mass (g)
conc (g dm^{-3}) | vol (dm^3)

When the solute dissolves in the solvent, the mixture is termed a solution.
Diluted solution: a solution with little amount of solute/unit solvent
Concentrated solution: a solution with large amt of solute/unit solvent
Saturated solution: a solution containing the max amt of solute that can dissolve in the solvent at a particular temperature and pressure

Dilutions - adding more of a solvent to a solution
- Results in a drop of concentration (# of moles of solute is dissolved in a greater volume of solvent)
- cV (# of moles) before dilution = cV (# of moles) after dilution

$C_1V_1 = C_2V_2$

A 50.0 cm^3 aqueous solution has a concentration of 1 $mol\ dm^{-3}$. 15 cm^3 of water is added into the solution. What is the new concentration of the solution?

$\frac{n}{0.05(dm^3)} = 1\ mol\ dm^{-3}$

$\frac{n}{0.065dm^3} = 0.769\ mol\ dm^{-3}$

(vol) GAS

vol (dm^3)
amt (mol) | molar vol ($dm^3 mol^{-1}$)
↳ stp : 22.7
rtp : 24

Avogadro's law

at the same temperature and pressure, equal volumes of gases contain the same number of particles (atoms/molecules)

- The fixed volume of any gas always contains the same # of particles
- For gaseous reactants & products, the mole ratio = the vol ratio

∆ dif gases w = masses have dif amt of molecules

Under conditions of fixed temperature and pressure:
$\frac{volume\ of\ gas\ 1}{volume\ of\ gas\ 2} = \frac{amount\ of\ gas\ 1}{amount\ of\ gas\ 2}$

Composition

1. **Empirical formula**
   - Shows the types of elements present in it
   - Shows the simplest ratio of elements present
   - Usually used for ionic compounds/covalent substances with a lattice structure
2. **Molecular formula**
   - Shows all the atoms (exact # of each element) in one molecule
   - It a multiple of its empirical formula
   - Mainly used for organic compounds, some may have the same empirical formula but may have a different molecular formula

$$n = \frac{molar\ mass\ of\ compound}{molar\ mass\ of\ empirical\ formula}$$

Ex: A hydrocarbon contains 85.7 % carbon and 14.3 % hydrogen. Find its empirical formula.

	C	H
Mass of each element in 100g sample/g	85.7	14.3
Molar mass of each element/ $g\ mol^{-1}$	12	1
Amount of each element/ mol	$\frac{85.7}{12} = 7.14$	$\frac{14.3}{1} = 14.3$
Molar ratio (dividing throughout by smallest #)	1	2
Simplest ratio (make whole numbers)	**1**	**2**

Its empirical formula is CH_2

Ex: A sample of a compound has the composition: sodium 9.20 g, sulfur 12.8 g and oxygen 9.60 g. Determine the empirical formula of the compound.

	Na	S	O
Mass of each element/g	9.20	12.8	9.60
Molar mass of each element/ $g\ mol^{-1}$	23	32.1	16
Amount of each element/ mol	0.4	0.4	0.6
Molar ratio (dividing throughout by smallest #)	0.2	0.2	0.3
Simplest ratio	**2**	**2**	**3**

The empirical formula of the compound is $Na_2S_2O_3$

CHEMICAL EQUATION

coefficients of atoms / molecules indicate mol ratio

Limiting reagent

→ less amt so it stops the reaction from continuing

$Mg(s) + 2HCl(aq) \rightarrow MgCl_2(aq) + H_2(g)$
2mol 1mol

① choose an atom
Mg 2HCl
2 mol needs 4 mol
2HCl (aq) is the limiting reagant

② Use the limiting reagant to calc for other stuff

- 0.50 mol of aluminum reacts with 0.72 mol of iodine to form aluminum iodide.
  $2Al + 3I_2 \rightarrow 2AlI_3$
  a) Determine the amount (in moles) of aluminum iodide formed.
  0.5 mol of Al requires $\frac{3}{2}(0.5) = 0.75\ mol\ of\ I_2$ for complete reaction
  0.72 mol of I_2 requires $\frac{2}{3}(0.72) = 0.48\ mol$ of Al for complete reaction
  b) Determine the amount (in moles) of the excess reactant remained.

Al $\frac{0.5}{2} > \frac{0.72}{3}$ I Shortcut?

Iodine is the limiting reactant.
The amount of Al remained: 0.5 - 0.48 = 0.02 mol.

% purity

chemicals that are used in a reaction are assumed to be more, but most of the time they aren't (impure). It is possible to use these reactants in an experiment and it is assumed that the impurities does not react in the reaction.

Formula: $\%\ purity = \frac{mass\ of\ pure\ compound}{mass\ of\ compound\ used\ in\ reaction}(100\%)$

- The reactant used in an actual experiment will always have traces of impurities (its mass is measured using the electronic mass balance)
- The mass of pure compound - data obtained from chemical calculation

% yield

- **Theoretical yield** - the data obtained from chemical calculation, represents the expected mass of product that would be obtained if the reaction goes to completion
- **Experimental yield** - data obtained by performing the actual experiment - represents the actual mass of product obtained from the experiment and is measured using the electronic mass balance
  % yield - indicates the p% of the theoretical yield (of final product) obtained in the experiment. It depicts how successful a reaction is in giving the products.
- The higher the % yield, the more successful the experiment is

Formula: $\%\ yield = \frac{actual\ mass\ of\ product\ (experimental)}{theoretical\ mass\ of\ product}(100\%)$