Moles

Definition

Much like the word dozen, a mole represent a certain amount of a substance. Mole is basically 6 * 10^23 of anything. This is often used to convert between amu (atomic mass units) and grams. Additionally, it is a convention and used in experiments. A formal definition is that mass of substance containing the same number of fundamental units as there units as there are atoms in exactly 12.00g of carbon–12

Convert between the amount of a substance (moles) and number of atoms, molecules, ions, electrons and formula units

To convert from atoms and molecules to moles:

Divide by 6.02×10^{23}

e.g. 8.12×10^{23} molecules = 1.35 Mol

To convert from moles to atoms and molecules:

Multiple by 6.02×10^{23}

e.g. 1 Mole of O_2 = 12.044×10^{23} O atoms

Mole Ratio

The mole ratio is the number of moles of each substance involved in a chemical reaction. The mole ratio can only be calculated if the chemical equation is balanced

e.g.

H2 + O2 = H2O **Given Equation**

2H2 + O2 = 2H2O **Balanced Equation**

In this equation, the mole ratio of Hydrogen gas (H2) to Oxygen gas (O2) is 2:1, and the ratio of Hydrogen gas (H2) to Water (H2O) is 2:2 or 1:1 when simplified. This means that when Oxygen and Hydrogen gas react, 1 mole of O2 reacts with 2 moles of H2 to produce 2 moles of H2O.

Moles and Mass

Determining the Mass of One Mole

The mass of one mole of a substance corresponds to its molar mass, which is derived from the sum of the atomic masses of its constituent elements.

Examples:

Hydrogen (H) = 1

Carbon (C) = 12

Oxygen (O) = 16

Understanding Mass Number (A), Relative Atomic Mass (Ar), and Relative Molecular Mass (Mr)

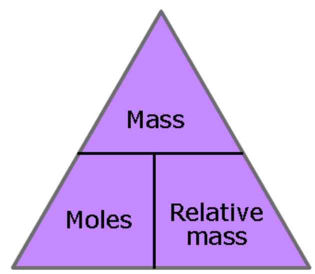
Mass Number (A): The total number of protons and neutrons present in the nucleus of an atom.

Relative Atomic Mass (Ar): The average mass of an atom of an element compared to the mass of a Carbon-12 atom, accounting for the natural abundance of its isotopes.

Relative Molecular Mass (Mr): The sum of the relative atomic masses of all the atoms in a molecule.

Formular

Mass = Formula Mass x Number of Moles



<u>Moles and Concentration</u>

Distinguish between the terms solute, solvent, solution, and concentration

Solute - A pure substance which would be dissolved in a solvent.

Example: Salt and sugar

Solvent - A substance where the solute is dissolved.

Example: Water

Solution - A substance in which a solute is completely dissolved in a solvent.

Example: Salt (solute) completely dissolves in water (solvent) to form salt water (solution)

Concentration - The proportion of the amount of solute to the volume of the solvent, measured in mol/dm3 or g/dm3

Saturated - A saturated solution is when the maximum amount of a solute is dissolved in a solvent at a specific temperature. Factors affecting Saturation include temperature, pressure, and chemical composition.

Solve problems involving concentration, amount of solute, and volume of solution

Concentration (c) = Number of moles (n) / Volume (v)

If the amount of a substance is given in grams, it must first be converted to moles using the formula:

Number of moles (n) = Mass (m)/Molar mass(Mr)

Define STP and SATP

STP - Standard Temperature and Pressure

Standard temperature is 0° Celsius and standard pressure is 1 ATM or 100k pascals.

SATP - Standard Ambient Temperature and Pressure or RTP (Room Temperature and Pressure)

Standard ambient temperature is 25°C and standard pressure is 1 ATM or 100k pascals.

Calculation of reacting volumes of gases using Avogadro's Law

Formulas to calculate for number of moles:

Solid = Mass / Molar Mass

Liquid = Concentration x Volume

Gas = Volume / Molar Volume (STP or SATP)

Avogadro's Law states that equal volumes of all gases, when measured at the same temperature and pressure, contain an equal number of particles.

Molar Volume is the volume occupied by one mole of gas at a given temperature and pressure.

At STP, 1 mole of gas (6.023 x 10^23 particles) occupies 22.7 dm3 at 0°C and 100kPa.

At SATP, 1 mole of gas (6.023 x 10^23 particles) occupies 24 dm3 at 25°C and 100kPa.

<u>Practical Applications</u>

Calculate the Percentage Composition by Mass of a Compound from its Formula

% of the mass of an element in a compound =100×(Number of atoms of the element × Relative formula mass of the compound)

Example: Water (H_2O)
% of mass of oxygen = 100×(1×16/18) = 88.89% of H_2O is oxygen

Use Equations to Calculate the Masses of Reactant and Products

$CaCO_3$ + 2HCL → $CaCl_2$ + H_2O + CO_2
Given: Calcium Carbonate = 20g
Calculate mass of CO_2:

i. First balance the chemical equation
ii. Convert the mass of calcium carbonate to moles
Mass / Molar Mass = 20/(40+12+(16×3)=20/100=0.2mol
iii. Use mole ratio to compare 1 mole of $CaCO_3$ gives out 1 mole of CO_2
Since ratio is 1:1
CO_2 = 0.2 moles
iv. Mass of CO_2 = Moles x Molar Mass
Mass of CO_2 = 0.2(12+32) = 9.8g

Calculating the Limiting Reactants

Limiting reactants - The limiting reactant determines how much product can be formed.
Yield - The yield from a chemical reaction is the mass of product made.
Theoretical Yield - The theoretical yield is the quantity of the product that can be formed from the complete conversion of the limiting reactant.
Actual Yield = The amount of product actually produced is the actual yield.
Percentage yield = (Actual yield/Theoretical yield) × 100

Example:
$CaCO_3 \rightarrow CO_2 + CaO$
If 22 grams of CO_2 is produced with the decomposition of one mole of calcium carbonate, what is the percentage yield of this experiment?
The theoretical yield assumes that 1 mole of CO_2 was produced because the mole ratio between calcium carbonate and carbon dioxide is 1:1.
Mass of one mole of CO_2 = 12+(2×16)=12+32=44g
Percentage Yield = (22/44)×100=50%

Reasons for inefficient reactions and actual yields:

- The reaction may be incomplete.
- Some product is lost during practical preparation.
- There may be unwanted reactions taking place.
- It can be difficult to separate the product from other products.
- The product may be impure.

<u>Skills</u>
Distinguishing Qualitative from Quantitative Chemistry

Qualitative Chemistry: Focuses on non-numerical data, observations, and conclusions.

Quantitative Chemistry: Concerned with numerical data, calculations, and measurements.

Understanding Physical and Chemical Changes

Physical Change: Alteration in state without changing the substance's chemical composition. Reversible through processes like heating, cooling, and evaporating.

Example: Water → Ice

Chemical Change: Results from a chemical reaction, changing the substance's chemical composition. The process can be exothermic or endothermic and is irreversible through simple means.

Example: Egg, Flour, Sugar → Cake

Differentiating Empirical from Molecular Formulas

Empirical Formula: Represents the simplest whole ratio of atoms in a compound.

Example: Glucose's empirical formula is CH_2O.

Molecular Formula: Shows the actual number of atoms of each element in a compound.

Example: Glucose's molecular formula is $C_6H_{12}O_6$.

Calculating the Empirical Formula of a Compound

Empirical formulas can be determined by utilizing percentage composition and dividing the percentage of each element by its respective molar mass.

Example: 10.8% Mg, 31.8% Cl, 57.4% O *Calculate subscripts based on these values.*

$$\frac{10.8}{24}, \frac{31.8}{35.5}, \frac{57.4}{16}$$

$$\frac{0.45}{0.45}, \frac{0.89}{0.45}, \frac{3.6}{0.45}$$

$$= 1,2,8$$

Compound formula: $MgCl_2O_8$ *after dividing by lowest number of moles*

Distinguishing Subscripts and Coefficients in Chemical Equations

Subscripts: Indicate the number of a specific atom in the chemical equation. They remain constant during equation balancing.

Example: O_2

Coefficients: Represent the number of molecules of a substance.

Example: 2HCl

Deriving Chemical Equations from Given Reactants and Products

Reactants and products must be balanced.

Examples:

a. $C + O_2 \rightarrow CO$

b. $SO_2 + O_2 \rightarrow SO_3$

Converting Volume Units

cm^3 to dm^3: 1 cm^3 = 0.001 dm^3 = 1 ml

dm^3 to litres: 1 dm^3 = 1 litre = 1000 ml

Laboratory Techniques in Titrations

Titration: Experimental method to determine the concentration of a solution using a known concentration solution (titrant).

Titrant is added to the other solution until neutralization, identified using a universal indicator. Formula $n = cv$ helps determine the amount and concentration of the unknown solution.

Example: 50 cm^3 of 0.2 mol/dm^3 Sodium Hydroxide neutralizes 25 cm^3 of unknown concentration Sulphuric Acid.

Balanced equation: $2NaOH + H_2SO_4 \rightarrow Na_2SO_4 + 2H_2O$

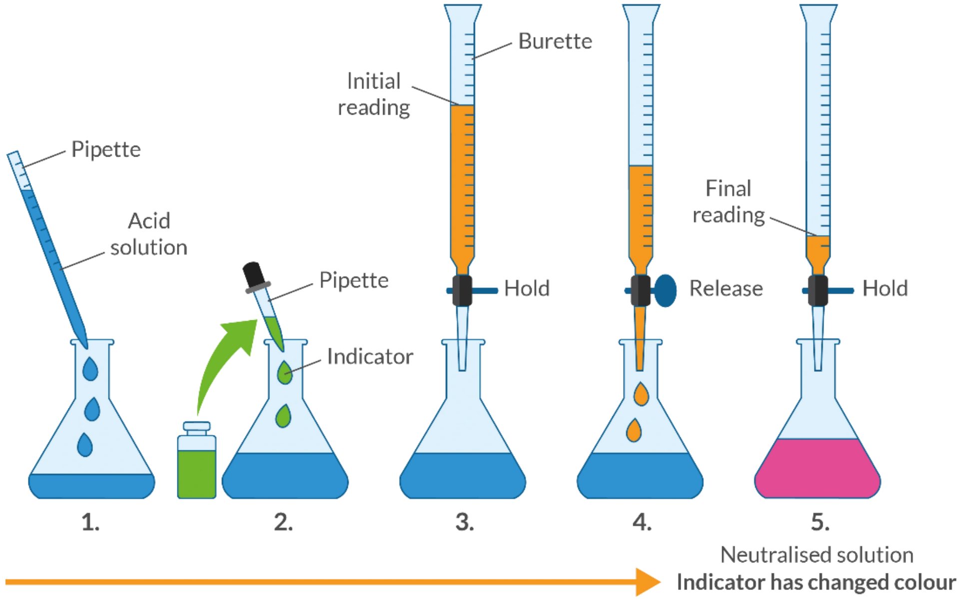
Moles of NaOH: 0.05 moles

Moles of H_2SO_4: 0.005 moles **using mole ratios**

Concentration of H_2SO_4: 0.2 mol/dm^3

See image below

Titration


Pipette
Acid solution
Pipette
Indicator
Initial reading
Burette
Hold
Release
Final reading
Hold
1.
2.
3.
4.
5.
Neutralised solution
Indicator has changed colour