L-10 / 10.01.2024 /

Chapter-3

Mass Relationship in Chemical Reaction

The Mole:

- 1 mole of atom ⇒ 6.02×10^{23} atoms
- 1 mole of molecules ⇒ 6.02×10^{23} molecules
  → Avogadro's number
- mole ⇒ 6.02×10^{23} of anything

Mole:

- is about mass.
- One mole of substance is its relative formula mass in grams or atomic mass.
- one mole of carbon atom = 12 grams

$$\text{mole, } n = \frac{m}{M}$$

m → mass
M → Molar mass

$$\Rightarrow m = n \times M$$

$$M = \frac{m}{n}$$

Problem - Slide-9

Atomic Mass:

- Mass of an atom, depends on the number of electrons, protons and neutrons.

- amu - atomic mass unit

Problem - Slide-13

Relative Formula Mass (RFM)

- H_2O = $2 \times H + 1 \times O$

$= 2 \times 1 + 1 \times 16$

$= 18$

Formula weight:

$C_{12}H_{22}O_{11} = 12 \times 12 + 22 \times 1 + 11 \times 16$

$= 342$ gm

The Mass Spectrometer:

- most direct and accurate method to measure atomic and molecular mass.

✳ Relative Atomic Mass (RAM):

$$= \frac{\Sigma(\text{isotopic mass} \times \%\ \text{abundance})}{100}$$

Problem - Slide - 29 - 31

✳ Percent Composition of Compounds:

⇒ NH_3

r.f.m. of $NH_3 = 1\times14 + 3\times1$

$= 14 + 3 = 17$

% of $N = \frac{1\times14}{17} \times 100\%$

$= 82\%$

% of $H = \frac{3\times1}{17} \times 100\%$

$= 18\%$

Problem - Slide - 34 | 37, 38

✳ Empirical formula: smallest whole-number ratio of atoms present in a compound.

CH

✳ Molecular formula: Actual number of each type of atom present in a given compound.

C_6H_6

Percent composition:

- The percentage of the mass contributed by each element in a substance.

Formula weight:

- The sum of the atomic weights of all of the atoms in a chemical formula.

Calculating Empinical Formula:

⇒ C : H : N

$\frac{74.03\ g}{12\ g/mol} : \frac{8.70\ g}{1\ g/mol} : \frac{17.27}{14}$

⇒ $\frac{6.16\ mol}{1.238} : \frac{8.70}{1.238} : \frac{1.238}{1.238}$ ↳ divide by smallest

⇒ 5.008 : 7.07 : 1

⇒ 5 : 7 : 1

↳ E.F. ⇒ C_5H_7N

C	:	H	:	Cl
10.06%		0.84%		89.10%

$$\Rightarrow \frac{10.06}{12} : \frac{0.84}{1} : \frac{89.10}{35.5}$$

$$\Rightarrow \frac{0.883}{0.84} : \frac{0.84}{0.84} : \frac{2.5098}{0.84}$$

$$\Rightarrow 1.05 \approx 1 : 1 : 2.976 \approx 3$$

$$\Rightarrow 1:1:3$$

$\hookrightarrow$ E.F. $= CHCl_3$

Molecular Formula (M.F.) = E.F $\times n$

Given,

$MW = 90\ gm = (CH_2O)$

$$\therefore n = \frac{90}{(CH_2O)} = \frac{90}{12+2\times1+16} = 3$$

$\therefore$ M.F. = EF $\times n$

$= (CH_2O) \times 3$

$= C_3H_6O_3$

Problem - Slide - ~~39~~ 40, 41, 44-46

Moles & Chemical Equations:

- Equations can be interpreted at both the microscopic and macroscopic level.

$$C(s) + O_2(g) \Rightarrow CO_2(g)$$

1 C atom + 1 O_2 molecule ⇒ 1 CO_2 molecule

1 mole C + 1 mole O_2 ⇒ 1 mole CO_2

Moles & Equations:

- Equations DO NOT relate to masses of the substances.

$$C + O_2 \Rightarrow CO_2$$

12 g C + 32 g O_2 ⇒ 44 g CO_2

Reactants ⇒ Products

each atom type = # each atom type

Balancing :

- Translate the statement
- Balance the atoms using coefficients
  - formulas cannot be changed
- Adjust coefficients if necessary
- Check that all atoms balance
- Specify states of matter

⇒ Magnesium and oxygen gas react to give magnesium oxide!

⇒ $Mg + O_2 \Rightarrow MgO$

⇒ $2Mg + O_2 \Rightarrow 2MgO$

⇒ $2Mg(s) + O_2(g) \Rightarrow 2MgO(s)$

How to read chemical Equations:

$$2Mg + O_2 \Rightarrow 2MgO$$

- 2 atoms Mg + 1 molecule O_2 makes 2 formula units MgO.

- 2 moles Mg + 1 mole O_2 makes 2 moles MgO.

- 48.6 grams Mg + 32.0 grams O_2 makes 80.6 grams MgO.

✳ Matter's state symbol:

s = solid

l = liquid

g = gas

aq = aqueous (dissolved in water)

✳ 4 type of reaction:

i. Combination reaction

ii. Decomposition reaction

iii. Displacement reaction

iv. Combustion reaction

Combination reaction:

- is a reaction in which two substances combine to form a third substance.

⇒ $2\,Na_{(s)} + Cl_{2(g)} \rightarrow 2\,NaCl_{(s)}$

$2\,Sb + 3\,Cl_2 \longrightarrow 2\,SbCl_3$

$CaO_{(s)} + SO_{2(g)} \longrightarrow CaSO_{3(s)}$

Decomposition reaction:

- is a reaction in which a single compound reacts to give two or more substance.
- Often these reactions occur when the temperature is raised.

⇒ $2\,HgO_{(s)} \xrightarrow{\Delta} 2\,Hg_{(l)} + O_{2(g)}$

$2\,KClO_{3(s)} \xrightarrow[MnO_2]{\Delta} 2\,KCl_{(s)} + 3\,O_{2(g)}$

$CaCO_{3(s)} \xrightarrow{\Delta} CaO_{(s)} + CO_{2(g)}$

Displacement Reaction:

⇒ Single Replacement Reaction:

- Sodium chloride solid reacts with fluorine gas

⇒ $2\,NaCl_{s} + F_{2\,g} \Rightarrow 2\,NaF_{s} + Cl_{2\,g}$

- Aluminum metal reacts with aqueous copper(II) nitrate

⇒ $2\,Al_{s} + 3\,Cu(NO_3)_{2\,aq} \Rightarrow 3\,Cu_{s} + 2\,Al(NO_3)_{3\,aq}$

⇒ Double Replacement Reaction:

- $AgNO_{3\,aq} + NaCl_{s} \longrightarrow AgCl_{s} + NaNO_{3\,aq}$

- $K_2SO_{4\,aq} + Ba(NO_3)_{2\,aq} \longrightarrow 2\,KNO_{3\,aq} + BaSO_{4\,s}$

Combustion Reaction:

- Organic compounds usually burn in oxygen or air to yield carbon dioxide. If the compound contains hydrogen (as most do), water is also a product.

$2\,C_4H_{10\,g} + 13\,O_{2\,g} \longrightarrow 8\,CO_{2\,g} + 10\,H_2O_{g}$

Problem - Slide - 55, 56

$$2C_8H_{18}(l) + 25O_2(g) \longrightarrow 16CO_2(g) + 12H_2O(g)$$

Stoichiometric Calculation:

$$C_3H_8(g) + 5O_2(g) \longrightarrow 3CO_2(g) + 4H_2O(g)$$

⇒ The coefficient in a balanced chemical equation:

- represent the relative number of reactant and product particles
- and the relative number of moles of each

⇒ Since moles are related to mass:

- the equation can be used to calculate masses of reactants and/or products for a given reaction.

The mole ratios from the balanced equation are used as conversion factors.

Stoichiometric Calculations:

- Write and balance the equation for the reaction.
- Convert known masses of reactants to moles.
- Using the numbers of moles of reactants and the appropriate mole ratios, determine which reactant is limiting.
- Using the amount of the limiting reactant and the appropriate mole ratios, compute the number of moles of the desired product.
- Convert from moles of product to grams of product, using the molar mass. (if required)

$$2\,NH_3 + 3\,CuO \longrightarrow N_2 + 3Cu + 3\,H_2O$$

(18.1 g over $2\,NH_3$; 90.4 g over $3\,CuO$)

$M = 17.03\ g$ for NH_3

$M = 79.55\ g$ for CuO

⇒ Given,

$$NH_3 = \frac{18.1}{17.03} = 1.06\ mol$$

$$CuO = \frac{90.4}{79.55} = 1.14\ mol$$

⇒

2 mol NH_3 react. with 3 mol CuO

∴ 1 mol " " " $\frac{3}{2}$ mol CuO

∴ 1.06 mol NH_3 " " $\frac{3 \times 1.06}{2}$ mol CuO

$= 1.59$ mol CuO

⇒

Required 1.59 mol CuO

Given 1.14 mol CuO

Therefore, CuO run out before NH_3 does

Therefore, CuO is the limiting reactant. So we must we the amount of CuO in calculating the amount of N_2 formed.

$\Rightarrow$ from 3 mol CuO, N_2 formed 1 mol

$\therefore$ " 1 mol CuO, N_2 " $\frac{1}{3}$ mol

$\therefore$ " 1.14 mol CuO, N_2 " $\frac{1 \times 1.14}{3}$ mol

$= 0.38$ mol

$\Rightarrow$

$m = n \times M$

$= (0.38 \times 28.02)$ g

$= 10.6$ g N_2

Problem - Slide - 62 - 67

Limiting Reactant:

- The reactant used up first in a reaction is called the limiting reagent, because the maximum amount of product formed depends on how much of this reactant was originally present.

- Excess reagents are the reactants present in quantities greater than necessary to react with

the quantity of the limiting reagent.

$$2NO + O_2 \rightarrow 2NO_2$$

Here,

NO → limiting reagent (Higher Ratio)

O_2 → excess reagent (Lower Ratio)

Problem- Slide-72

Reaction Yield:

⇒ Theoretical Yield ⇒ The amount of product that would result if all the limiting reagents react.

⇒ Actual Yield ⇒ The amount of product actually obtained from a reaction.

⇒ Reaction Yield ⇒ The proportion of the actual yield to the theoretical yield.

$$\Rightarrow \%\ \text{yield} = \frac{\text{actual yield}}{\text{theoretical yield}} \times 100\%$$

Problem- Slide- 74, ~~75~~ -78

⊗ Group-17 ⇒ halogen

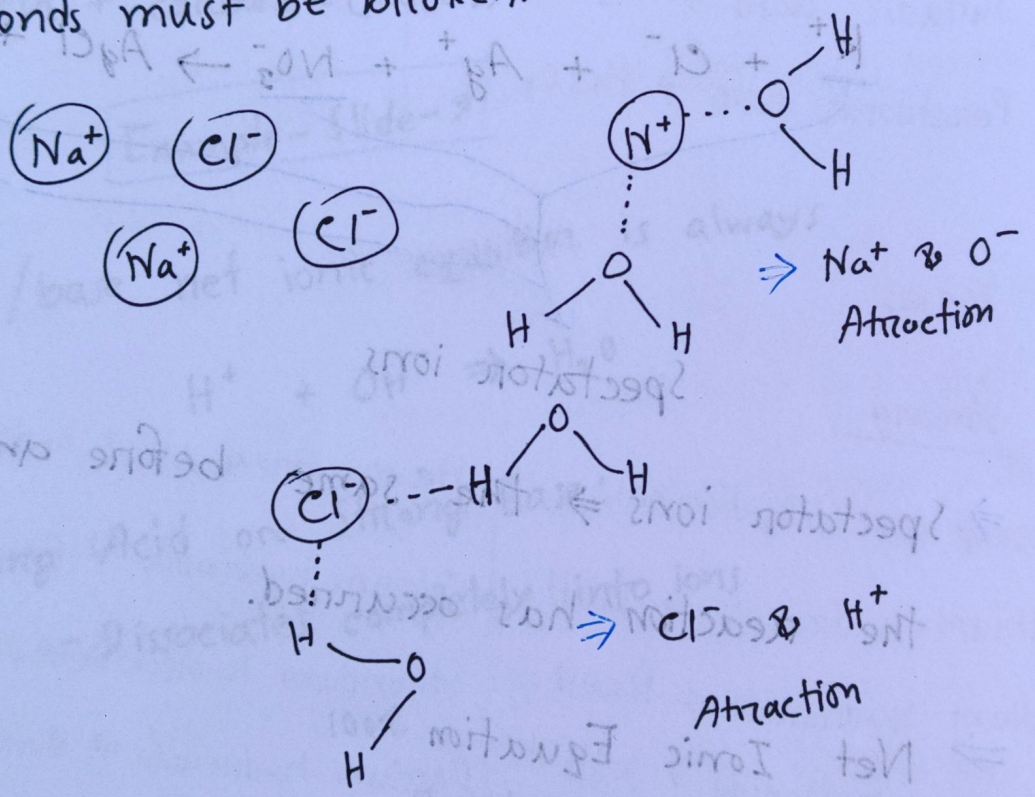
L-11 / 15.01.2024 /

Chapter-4

Reactions in Aqueous Solutions

Dissolution of Ionic Salts

- If ionic compounds are to be dissolved, chemical bonds must be broken.



- Also called dissociation

⇒ Potassium chloride

$$KCl_{(s)} \longrightarrow K^+_{(aq)} + Cl^-_{(aq)}$$

⇒ Silver nitrate

$$AgNO_3(s) \longrightarrow Ag^+(aq) + NO_3^-(aq)$$

$$KCl(s) + AgNO_3(aq) \longrightarrow AgCl(s) + KNO_3(aq)$$

⇒ Total Ionic Equation

$$K^+ + Cl^- + Ag^+ + NO_3^- \rightarrow AgCl + K^+ + NO_3^-$$

Spectator ions

⇒ Spectator ions ⇒ the same before and after the reaction has occurred.

⇒ Net Ionic Equation

$$Ag^+(aq) + Cl^-(aq) \longrightarrow AgCl(s)$$

Problem - Slide - 8,10

Solid, liquid, gas remain in their same form.

Acid - Base Reactions

Neutralization Reactions

- acid + alkali → salt + water
- acid + ~~metal~~ carbonate → salt + water + carbon dioxide
- acid + metal → salt + hydrogen
- acid + alkali (NaOH) → salt + water
- acid + carbonate (Na_2CO_3) → salt + water + CO_2

Example - Slide-9

Acid/base net ionic equation is always

$$H^+ + OH^- \Rightarrow H_2O$$

Strong Acid on Strong Base:

- Dissociates completely into ions

100%.

Weak Acid on Weak Base:

- Does NOT dissociate completely

5 out of 1000, 0.5%.

Acids

Strong

Hydrochloric acid ⇒ HCl

Hydrobromic acid ⇒ HBr

Hydriodic acid ⇒ HI

Nitric acid ⇒ HNO_3

Sulfuric acid ⇒ H_2SO_4

Perchloric acid ⇒ $HClO_4$

Weak

Hydrofluoric acid ⇒ HF

Phosphoric acid ⇒ H_3PO_4

Acetic Acid ⇒ CH_3COOH $(HC_2H_3O_2)$

Bases

Strong

Group 1A(1) hydroxides:

Lithium hydroxide ⇒ $LiOH$

Sodium hydroxide ⇒ $NaOH$

Potassium hydroxide ⇒ KOH

Rubidium hydroxide ⇒ $RbOH$

Cesium hydroxide ⇒ $CsOH$

Heavy Group 2A(2) hydroxides:

Calcium hydroxide ⇒ $Ca(OH)_2$

Strontium hydroxide ⇒ $Sr(OH)_2$

Barium hydroxide ⇒ $Ba(OH)_2$

Weak

Ammonia ⇒ NH_3

Problem - Slide - 12

Acid - Base Titrations:

⇒ In a titration, the concentration of one solution is used to determine the concentration of another.

⇒ An acid-base indicator has different colors in acid and base, and is used to monitor the progress of a titration experiment.

⇒ At the equivalence point, the mol of H^+ from the acid equals the mol of OH^- ion produced by the base.

pH of the solution = 7.

⇒ The end point occurs when there is a slight excess of base and the indicator changes colour permanently.

pH is slightly greater than 7.

⇒ $$\frac{M_A \times V_A}{n_A} = \frac{M_B \times V_B}{n_B}$$

Problem - Slide - 14, 15

Redox Reactions:

Reduction - Oxidation:

⇒ Oxidation is the loss of electrons

⇒ Reduction is the gain of electrons

⇒ Oxidation refers to the increase in oxidation number & reduction refers to the decrease in oxidation number.

Simple mnemonic ⇒ OIL RIG

OIL ⇒ Oxidation Is Loss

RIG ⇒ Reduction Is Gain

A reducing agent ~~gains~~ loses electrons and gets oxidized.

An oxidizing agent gains electrons and get reduced.

A redox reaction involves electron transfer.

Oxidation and reduction occur together.

$$2\,\overset{0}{Mg}_{(s)} + \overset{0}{O_2}_{(g)} \longrightarrow 2\,\overset{+2}{Mg}\overset{-2}{O}_{(s)}$$

Example - Slide - 17

Oxidation Number:

- Oxidation number of an atom in a substance as the actual charge of the atom ~~in a~~ if it exists as a monatomic ion, or a hypothetical charge assigned to the atom in the substance by simple rules.

⇒ Oxidation number:

Elements ⇒ 0

Monatomic ion ⇒ equal to the charge on the ion.

Oxygen ⇒ −2

Hydrogen ⇒ +1

Halogen ⇒ −1

Compound ~~& ions~~ ⇒ 0

Polyatomic ions ⇒ equal to the charge on the ion.

Rules:

Group 1A (1) ⇒ +1

Group 2A (2) ⇒ +2

hydrogen ⇒ +1 with nonmetals

−1 with metal & boron

fluorine ⇒ −1

Oxygen ⇒ −1 in peroxides

−2 (except F)

Group 7A (17) ⇒ −1 with metals & nonmetals (except O) and other halogen lower on the group.

Problem- Slide - 21-28

Practice - Slide - 30-36

PPT

Ions

Al^{3+} ⇒ White ppt ⇒ Soluble in excess – colorless

Pb^{2+} ⇒ White ⇒ s.e. colorless

Zn^{2+} ⇒ White ⇒ s.e. colorless

Cation

Cu^{2+} ⇒ Light Blue ⇒ insoluble in excess

Fe^{2+} ⇒ Dirty-green ⇒ "

Fe^{3+} ⇒ Reddish-brown ⇒ "

END

methane ⇒ CH_4 ⇒

ethene ⇒ C_2H_4 ⇒ $CH_2 = CH_2$

propen ⇒ C_3H_6 ⇒ $CH_2 = CHCH_3$

butene ⇒ C_4H_8 ⇒ $CH_2 = CH(CH_2)CH_3$

pentene ⇒ C_5H_{10} ⇒ $CH_2 = CH(CH_2)_2CH_3$

Alkane = C_nH_{2n+2}

Alkene = C_nH_{2n}