Chapter 2: Atoms, Molecules, and Ions

2.1 **Introduction**

Pure substances contain **compounds** (**combinations of elements**) or **pure elements**

Compound: salt =NaCl =sodium chloride: Elements: Na =sodium, Cl =chlorine

Elements are composed of **atoms**

A **molecule** is a collection of atoms

2.2 Fundamental Chemical Laws

Law of Conservation of Mass: Mass is neither created or destroyed. Input = Output**.**

Law of Definite Proportion: A given compounds (such as CO_2) always contains exactly the same proportion of elements of mass (such as 1 part of C and 2 parts of O)

2.5 The Modern View of Atomic Structure: an Introduction

The simplest view of Atom: It consists of a tiny nucleus (with a diameter of about 10^{-13} cm) and electrons that move about the nucleus at an average distance of about 10^{-8} cm from it.

Protons (positive charged)
Nucleus
Neutrons
Atom
Electrons (negative charged)

The number of protons = The number of electrons in an atom

Table 2.1 The mass and charge of the electron, proton and neutron		
Particle	Mass (kg)	Charge*
Electron	9.11×10^{-31}	1-
Proton	1.67×10^{-27}	1+
Neutron	1.67×10^{-27}	None

- the magnitude of the charge of the electron and the proton is 1.60×10^{-19}C

Expression of an Atom using symbol and number

Mass number

$^{A}_{Z}\mathbf{X}$ Elemental symbol

Atomic number

Atomic number Z = the number of protons

Mass number A = the total number of protons and neutrons

Isotopes = atoms with the same number of protons but different number of neutrons

Example 2.2 (P55) Writing the symbols for atoms

2.6 Molecules and Ions

Molecule = a collection of atoms through chemical bonds

Chemical bond = the forces that hold atoms together in compounds

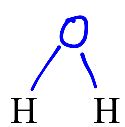
Covalent bond = electrons shared between 2 or more atoms. Hydrogen (H_2), Oxygen (O_2), methane (CH_4), ammonia (NH_3)

Ionic bond = the force of attraction between oppositely charged ions. Sodium chloride
(NaCl)

Chemical formula= The simplest way to represent molecules. The symbols for elements are used to indicate the types of atoms present and subscripts are used to indicate the relative numbers of atoms.

Carbon dioxide: CO_2. methane: CH_4

Structural formula = individual bonds are shown (indicated by the lines). Structural formula may or may not indicate the shape of the molecule. For example, H-O-H or O



Space-filling model= shows the relative sizes of the atoms as well as their relative orientation in the molecule.

Ball-and-stick model = the individual bonds and the relative orientation of atoms in the molecule are shown

IONS= An atom or a group of atoms that has a net positive (lost electrons)or negative charge (receive electrons). Example: Sodium chloride= Na^+, Cl^-

Cation = a positively charged ion (atoms lose electrons). Na^+: Na $Na^+ + e^-$

Anion = a negatively charged ion (atoms receive electrons). Cl^-: Cl $+ e^-$ - Cl^-

2.7. An Introduction to the Periodic Table

Periodic Table = A chart that summarizes all the elements we know. The symbols of atoms are used to represent the elements. The atomic number (proton number) is listed above the symbol. The average atomic mass is listed in a blanket under the symbol.

Mendeleer (1869) –Chemical and physical properties of elements vary in a periodic way-"periodic law".

The periodic table is arranged so that elements in the same vertical columns (called groups or families) have *similar chemical properties*.

Groups 1A -8A, Group 3B-12B (transition metals), Lanthanides group and Actinides group

2.8 Naming Simple Compounds

See the photocopies of the text book (p60-70).

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Chapter 3 Stoichiometry

3.2 Atomic Mass

The modern system of atomic masses, instituted in 1961, is based on ^{12}C (carbon twelve) as the standard. In this system, *^{12}C is assigned a mass of exactly 12 atomic mass unit (amu)*, and the masses of all other atoms are given relative to this standard.

Average atomic mass of natural carbon, which includes isotopes ^{12}C (98.89%) and ^{13}C (1.11%):

98.89% x 12 amu + 1.11% x 13.0034 amu = 12.01 amu.

Example 3.1 (P85) The average mass of an element or average atomic mass

$$69.09 \text{ ATOMS}\left(62.93 \frac{u}{\text{ATOM}}\right) + 30.91 \text{ ATOMS}\left(64.93 \frac{u}{\text{ATOMS}}\right) = 6355u$$

$\therefore$ AVG MASS of A Cu ATOM: $\frac{6335u}{100 \text{ ATOMS}} = 63.55\ u/\text{ATOMS}$

3.3 The Mole

Mole is a mass unit = the number equal to the number of carbon atoms in exactly 12 grams of pure ^{12}C = 6.02214×10^{23} = **Avogadro's Number**

One mole of something consists of 6.02214×10^{23} units of that substance.

Mole is defined such that a sample of a natural element with a mass equal to the element's atomic mass expressed in grams contains 1 mole of atoms.

Based on this definition, 1 mole of ^{12}C = 12 g ^{12}C.

Also, 6.022 x10^{23} atoms x 12 amu/atom = 12 g ---- 6.022 x10^{23} amu = 1 g

Example 3.2 (p87) Determination the mass of a sample of atoms

CALCULATE MASS OF 6 AMERICIUM ATOMS:

$Am = 243u$

$6 \text{ ATOMS} \cdot 243 \frac{u}{\text{ATOM}} = 1.46 \times 10^3 u$

$6.022 \times 10^{23} u = 1g$

$m = (6.022 \times 10^{23} \frac{u}{g})^{-1} (1.46 \times 10^3 u) = \underline{2.42 \times 10^{-21} g}$

Example 3.3 (p88) Determining Moles of Atom

RTF: # MOL IN 10.0g Al.

$1 \text{ MOL } (6.022 \times 10^{23} \text{ ATOMS}) = 26.93g$

$\left(\frac{26.93g}{10.0g}\right)^{-1} = 0.371333 \text{ MOL}$

$(0.371333 \text{ MOL})(6.022 \times 10^{23} \frac{\text{MOLEC.}}{\text{MOL}}) = \underline{2.23 \times 10^{23} \text{ ATOMS}}$

Example 3.4 (p89) Calculating Numbers of Atoms

$m = 5.68 \text{ mg} = 0.00568g \rightarrow$ HOW MANY ATOMS?

REF. PERIODIC TABLE. Si: 28.09 g/MOL

$\frac{0.00568g}{28.09g} = (2.02207 \text{ MOL})(6.022 \times 10^{23} \frac{\text{ATOMS}}{\text{MOL}}) = \boxed{1.22 \times 10^{20} \text{ ATOMS}}$

Example 3.5 (p89-90) Calculating the Number of Moles and Mass

- 5.00×10^{20} ATOMS, RTF # MOL

$\frac{5.00 \times 10^{20}}{6.022 \times 10^{24}} = 8.30 \times 10^{-4} \text{ MOL Co}$

$(8.30 \times 10^{-4} \text{ MOL})(\frac{58.93g}{\text{MOL}}) = \boxed{4.89 \times 10^{-2} g \text{ Co}}$

3.4 Molar Mass

The **molar mass** of a substance is the mass in grams of one mole of that compound.

Molar Mass = The sum of the atomic masses of the component atoms

For example, the molar mass of methane (CH_4) molecule: 1 mole of CH_4 molecule contains 1 mole of carbon and 4 moles of hydrogen. Thus

The molar mass of CH_4 molecule= 1 x 12.01 + 4 x1.008 =16.04 g.

Example 3.6 (p90-91) Calculating Molar Mass I

JUGLONE: $C_{10}H_6O_3$

a) 10(12.01 g/mol) + 6(1.01 g/mol) + 3(16.01 g/mol)

= 174.1 g/mol

b) SAMPLE: 1.56×10^{-2} g

$\frac{1.56 \times 10^{-2} g}{174.1 g/mol} = 8.96 \times 10^{-5}$ mol

Example 3.7 (p91-92) Calculating Molar Mass II

CALCITE: $CaCO_3$

a) 1(40.08 g/mol) + 1(12.01 g/mol) + 3(16.01 g/mol) = 100.09 g

b) 4.86 mol, mass in g?

mass of CO_3^{2-} present?

(4.86 mol)(100.09 g / mol) = 486 g $CaCO_3$,

CO_3 only: 60.01 g/mol

4.86 mol CO_3^{2-} · 60.01 g / mol = 292 g CO_3^{2-}

Example 3.8 (p92) Molar Mass and Number of Molecules

$C_7H_{14}O_2$

1 μg = 1×10^{6} g → 7(12.01) + 14(1.008) + 2(16.01) = 130.18 g/mol → $\frac{1E-6 g}{130.18 g/mol}$

= (8E-9 mol)(6.022E23 atoms/mol)

= 5E15 molecules

4E16 C atoms

3.6 Percent Composition of Compounds

Mass Percent (or weight percent) of the element = mass of the element in 1 mole of that compound/ mass of 1 mole of that compound

For example, ethanol (C_2H_5OH),

Molar mass of C_2H_5OH = 2x 12.01 +5 x1.008 + 1x16 +1x1.008=46.07 g

Mass percent of C in C_2H_5OH = mass of C in 1 mole of C_2H_5OH/ mass of 1 mole of C_2H_5OH
= 2 x12.01/ 46.07 =52.14%

Mass percent of H in C_2H_5OH = 6 x1.008/46.07 =13.13%

Mass percent of O in C_2H_5OH = 1 x16/46.07 = 34.73%

Example 3.9 (p95-96) Calculating Mass Percent

CARVONE → $C_{10}H_{14}O$

TOTAL ATOMIC MASS: 10(12.01 g/mol) + 14(1.008 g/mol) + 16.00 g/mol
= 150.2 g/mol

C: $\frac{120.1\ g/mol}{150.2} = 79.96\%$

H: $\frac{14.112}{150.2} = 9.4\%$

O: $\frac{16.00}{150.2} = 10.65\%$

3.7 Determining the Formula of a Compound (self-study)

3.8 Chemical Equations

Chemical equation is a tool that describes the rearrangement of atoms in chemical reactions with the reactants on the left side and the products on the right side.

For example, combustion of methane: CH_4 (g) + $2O_2$ (g) ---- CO_2(g) + 2 H_2O (g)

Information from Chemical Equation: 1.) the nature of the products and reactants, and 2.) the relative number of each

The physical state (liquid (l), gas (g), solid (s), aqueous solution (aq)) of the reactants and products should be indicated in the chemical equation.

Chemical reaction is the rearrangement of atoms, broken of chemical bonds in the reactants, and formation of new chemical bonds in the products.

Stiochiometery: Quantitative description of the relative amount of substances involved in a reaction.

Balancing Equation: The chemical equation must be atomically and electrically balanced.

3.9 Balancing Chemical Equation

Procedures involved in balancing chemical equation:

1.) Determine what reaction is occurring. What are the reactants, the products, and the physical states involved.
2.) Write the unbalanced equation that summarizes the reaction described in step 1.
3.) Balance the equation by inspection, starting with the most complicated molecules. Determine what coefficients are necessary so that the same number of each type of atom appears on both reactant and product sides. Do not change the identities (formula) of any of the reactants or products.

Example, Burning of ethanol,

Steps 1 and 2. C_2H_5OH(l) + O_2 (g) ----- CO_2 (g) + H_2O (g)

Step 3 a.) balance the number of carbon: $C_2H_5OH(l)$ + O_2 (g) ----- $2CO_2$ (g) + H_2O (g)

3 b.) balance the number of hydrogen: $C_2H_5OH(l)$ + O_2 (g) -- $2CO_2$ (g) + $3H_2O$ (g)

3 c.) balance the number of oxygen: $C_2H_5OH(l)$ + $3O_2$ (g) --- $2CO_2$ (g) + $3H_2O$ (g)

Overall balanced chemical equation:

$$C_2H_5OH(l) + 3O_2(g) \text{ --- } 2CO_2(g) + 3H_2O(g)$$

Example 3.13 (p107) balancing a chemical equation I

$$(NH_4)_2Cr_2O_{7(s)} \rightarrow Cr_2O_{3(s)} + N_{2(g)} + 4H_2O_{(g)}$$

Example 3.14 (p108) balancing a chemical equation II

$$2\ NH_{3(g)} + 3O_{2(g)} \rightarrow 2NO_{(g)} + 3H_2O_{(g)}$$

3.10 Stoichiometric Calculations: Amounts of Reactants and Products

Procedures involved in stoichiometric calculation of the amount of reactants and products:

1.) balance the chemical equation of the reaction;
2.) convert the known mass of the reactant or product to moles of that substance;
3.) use the balanced equation to set up an appropriate mole ratios
4.) Use the appropriate mole ratios to calculate the number of moles of the desired reactant or product;
5.) Convert from moles back to grams if required by the problem.

Example, **what mass of oxygen will react with 96.1 g of propane**?

C_3H_8 (g) + 5 O_2 (g) ----- 3CO_2(g) + 4 H_2O (g)

Mole no. of propane= 96.1 g/ 44.1 g/mole = 2.18 mole C_3H_8

From the above balanced equation, it is clear that 1 mole of C_3H_8 will need 5 moles of O_2 to complete the reaction and produce 3 moles of CO_2 and 4 moles of H_2O.

Thus the mole no. of oxygen required for burning 96.1 g propane = 2.18 x 5/1 = 10.9 mole O_2

10.9 mole of O_2 = 10.9 mole x 32 g/mole = 349 g O_2

Example 3.15 (p112) Chemical Stoichiometry I

$2LiOH_{(s)} + CO_2(g) \rightarrow Li_2CO_3(s) + H_2O_{(l)}$

1.00 Kg LiOH.

WHAT MASS OF CO_2 ABSORBED BY 1.00 Kg LiOH.

$LiOH = 6.941 + 16.00 + 1.008$
$= 23.949$ g/MOL

$\frac{1000 \text{ g}}{23.949 \text{ g/mol}} = 41.7554$ MOLS LiOH ∴ 20.8777 MOL

CO_2 : 44.01 g/MOL

↳ (20.8777)(44.01g)

= 918.828 g = 0.919 Kg

Example 3.16 (p113) Chemical Stoichiometry II

$NaHCO_{3(s)} + HCl_{(aq)} \longrightarrow NaCl_{(aq)} + H_2O_{(l)} + CO_{2(aq)}$

$Mg(OH)_{2(s)} + 2HCl \longrightarrow 2H_2O(l) + MgCl_2$

$NaHCO_3$: 84.008 g/mol

$Mg(OH_2)$: 60.342 g/mol. ← MORE EFFECTIVE

3.11 The Concept of Limiting Reagent

When there are several compounds participating in chemical reactions, they are going to react according to the stoichiometry. For example,

$$O_2 \ + \ 2H_2 \ \text{------} \ \ 2H_2O$$

1 mole O_2 is going to react with 2 moles H_2 to produce 2 moles H_2O. If the reactants contain 2 moles O_2 and 2 moles H_2, then we say O_2 has excess amount (1 mole), and H_2 is the limiting reactant, because 1 mole O_2 is going to react with 2 moles H_2.

Example 3.17 (p119) Stoichiometry: Limiting Reactant

$N_{2(g)}$

$2NH_{3(g)} + 3CuO \longrightarrow N_{2(g)} + 3Cu_{(s)} + 3H_2O_{(g)}$

NH_3	CuO	N_2
18.1 g	90.4 g	? g
17.034 g/mol	79.55 g/mol	
$\frac{1.06258\ mol}{2}$	$\frac{1.136392\ mol}{3}$	(28.02 g/mol)(0.378797 mol)
0.531290 mol	0.378797 mol LIMITING	= 10.6139 g
		= 10.61 g

Example 3.18 (p121-123) Calculating Percent Yield

$$\% \text{ YIELD} = \frac{\text{ACTUAL}}{\text{THEORY}} \times 100\%$$

CH_3OH, METHYL ALCOHOL

$$CO_{(g)} + 2H_{2(g)} \longrightarrow CH_3OH_{(g)}$$

68.5 kg 8.60 kg ? kg → (2132.94 MOL)(32.042 g/MOL)

$$\frac{68500\text{ g}}{28.01\text{ g/MOL}} = 2445.56\text{ MOL}$$

$$\frac{8600\text{ g}}{2.016\text{ g/MOL}} = \frac{4265.87\text{ MOL}}{2} = 2132.94\text{ MOL}$$

LIMITING

= 68343.66 g

= 68.3 kg THEORY

$\%$ YIELD = 52.2 $\%$