Fundamentals of Chemistry

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Description

Supporting Material

RCNET
Regional Center for Nuclear Education & Training

Fundamentals of Chemistry

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Objectives

- Given an incomplete chemical reaction, balance it using the method presented.
- Given sufficient information about a solution, **CALCULATE the pH and pOH of the** solution.

Definition of Chemistry

- Defined as the systematic investigation of the properties, structure, and behavior of matter and the changes matter undergoes.

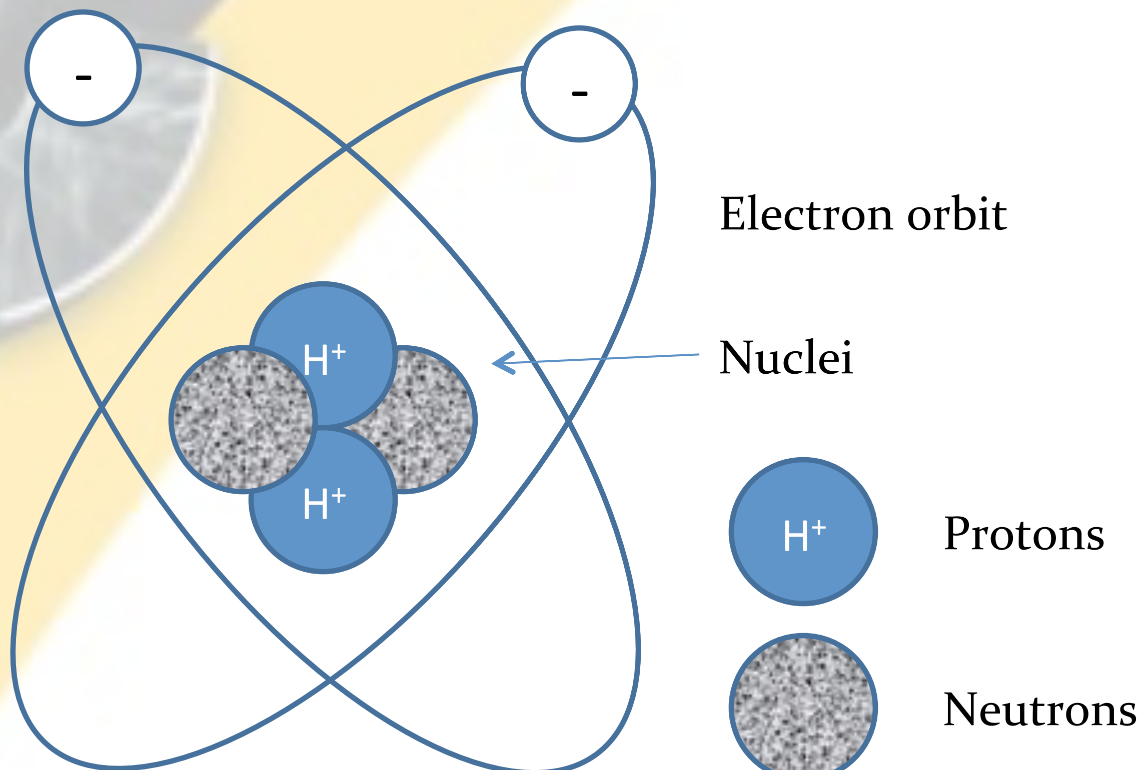
Matter

- States of matter refers to the physical forms that matter exists. It includes solid, liquid and gas.
  - Solids: definite shape and volume.
  - Liquids: specific volume but the shape is determined by the container.
  - Gas: No specific shape or volume.
- Solids experience the strongest interactions between the bonds and gases the weakest.
- All matter can be broken down into fundamental units called atoms.

The Atomic Structure

- All matter is composed of atoms, existing individually or in combination with each other.
- An atom is the smallest unit involved in a chemical reaction.
- Atoms are composed of smaller sub-particles.
- An atom is composed of a positively-charged nucleus orbited by one or more negatively-charged particles called electrons.
  - Neutrons are electrically neutral, protons are electrically positive.

Schematic: Particles in an Atom



The Atomic Structure

- The nuclei represent the mass of the atom.
  - The mass of the proton and neutron are almost identical.
  - The mass of an electron is very small (electron cloud is almost empty space).
  - Neutrons and protons are 1835 times heavier than an electron.
  - A proton has a charge of +1 and the neutron has no charge.
- The atom contains the same number of protons and electrons (no net charge).
  - For the charge to be different than zero a gain or loss of electrons must take place.
- The particles orbiting the nuclei are electrons.

The Atomic Structure

- The mass of one electron is equivalent to 1/1835 of the mass of a proton or neutron.
- The diameter of the atom is determined by the range of the electrons in their travels around the nucleus and is approximately 10^{-8} cm.
- The diameter of the nucleus is roughly 10,000 times smaller, approximately 10^{-13} to 10^{-12} cm.
  - Electrons constitute the largest portion of the volume of the nuclei.

The Atomic Structure

TABLE 1 Properties of the Atom and its Fundamental Particles		
Particle Name	Relative Mass (amu)	Relative Charge (based on charge of proton)
Electron	0.00055 or 1/1835	-1
Proton	1.0	1
Neutron	1.0	0

Amu = atomic mass unit which is a reference unit that assumes that the mass of a proton is approximately one.

Chemical Elements

- An atom is classified chemically by the number of protons in its nucleus.
- Atoms that have the same number of protons in their nuclei have the same chemical behavior.
- Atoms that have the same number of protons are grouped together and constitute a chemical element.
- The number of protons represents the atomic number. The symbol Z is often used for atomic number (or number of protons).
- The name of the elements are represented as using various symbols.
  - Each element has been assigned a specific one or two letter symbol based on the first letter of its chemical name.
  - In some cases the symbol comes from an abbreviation for the old latin name of the element.

The Atomic Mass Number

- The sum of the total number of protons, Z, and the total number of neutrons, N, is called the ***atomic mass number.***
- The symbol used for the atomic mass number is ***A***.
- Not all atoms of the same elements have the same mass.
  - The number of protons is the same but the number of neutrons may vary.

Atomic Weight

- The sum of the total number of protons, Z, and the total number of neutrons, N, is called the *atomic mass number.*
- These units represent a relative scale in which the mass of the isotope carbon-12 is used as the standard and all others are related to it.
  - 1 amu is defined as 1/12 the mass of the carbon-12 atom.
- Since the mass of a proton or a neutron is approximately 1 amu, the mass of a particular atom will be approximately equal to its atomic mass number, A.
- The *atomic weight of an element is defined as the weighted average of the masses of all of* its natural occurring isotopes.

- The elements that have their atomic weights in parentheses are unstable.
- The atomic weight shown for this elements is the weight of the most stable isotope.

TABLE 2
Table of Elements

Name and Symbol		Atomic Number	Atomic Weight (amu)	Name		Atomic Number	Atomic Weight (amu)
Actinium	Ac	89	(227)	Curium	Cm	96	(247)
Aluminum	Al	13	26.981	Dysprosium	Dy	66	162.50
Americium	Am	95	(243)	Einsteinium	Es	99	(252)
Antimony	Sb	51	121.75	Erbium	Er	68	167.26
Argon	Ar	18	39.948	Europium	Eu	63	151.96
Arsenic	As	33	74.921	Fermium	Fm	100	(257)
Astatine	At	85	(210)	Fluorine	F	9	18.998
Barium	Ba	56	137.34	Francium	Fr	87	(223)
Berkelium	Bk	97	(247)	Gadolinium	Gd	64	157.25
Beryllium	Be	4	9.012	Gallium	Ga	31	69.72
Bismuth	Bi	83	208.980	Germanium	Ge	32	72.59
Boron	B	5	10.811	Gold	Au	79	196.967
Bromine	Br	35	79.909	Hafnium	Hf	72	178.49
Cadmium	Cd	48	112.40	Helium	He	2	4.0026
Calcium	Ca	20	40.08	Holmium	Ho	67	164.930
Californium	Cf	98	(251)	Hydrogen	H	1	1.0079
Carbon	C	6	12.011	Indium	In	49	114.82
Cerium	Ce	58	140.12	Iodine	I	53	126.904
Cesium	Cs	55	132.905	Iridium	Ir	77	192.2
Chlorine	Cl	17	35.453	Iron	Fe	26	55.874
Chromium	Cr	24	51.996	Krypton	Kr	36	83.80
Cobalt	Co	27	58.933	Lanthanum	La	57	138.91
Copper	Cu	29	63.546	Lawrencium	Lw	103	(260)

TABLE 2 (Cont.)
Table of Elements

Name and Symbol		Atomic Number	Atomic Weight (amu)	Name		Atomic Number	Atomic Weight (amu)
Lead	Pb	82	207.19	Potassium	K	19	39.102
Lithium	Li	3	6.939	Praseodymium	Pr	59	140.90
Lutetium	Lu	71	174.97	Protactinium	Pa	91	231.03
Magnesium	Mg	12	24.312	Promethium	Pm	61	(145)
Manganese	Mn	25	54.938	Radium	Ra	88	226.02
Mendelevium	Md	101	(258)	Radon	Rn	86	(222)
Mercury	Hg	80	200.59	Rhenium	Re	75	186.2
Molybdenum	Mo	42	95.94	Rhodium	Rh	45	102.90
Neodymium	Nd	60	144.24	Rubidium	Rb	37	85.47
Neon	Ne	10	20.183	Ruthenium	Ru	44	101.07
Neptunium	Np	93	237.05	Samarium	Sm	62	150.35
Nickel	Ni	28	58.71	Scandium	Sc	21	44.956
Niobium	Nb	41	92.906	Selenium	Se	34	78.96
Nitrogen	N	7	14.006	Silicon	Si	34	78.96
Nobelium	No	102	(259)	Silver	Ag	47	107.87
Osmium	Os	76	190.2	Sodium	Na	11	22.989
Oxygen	O	8	15.999	Strontium	Sr	38	87.62
Palladium	Pd	46	106.41	Sulfur	S	16	32.064
Phosphorus	P	15	30.973	Tantalum	Ta	73	180.94
Platinum	Pt	78	195.09	Technetium	Tc	43	(98)
Plutonium	Pu	94	(244)	Tellurium	Te	52	127.60
Polonium	Po	84	(209)	Terbium	Tb	65	158.92

TABLE 2 (Cont.)
Table of Elements

Name and Symbol		Atomic Number	Atomic Weight (amu)	Name		Atomic Number	Atomic Weight (amu)
Thallium	Tl	81	204.37	Vanadium	V	23	50.942
Thorium	Th	90	232.03	Xenon	Xe	54	131.30
Thulium	Tm	69	168.93	Ytterbium	Yb	70	173.04
Tin	Sn	50	118.69	Yttrium	Y	39	88.905
Titanium	Ti	22	47.90	Zinc	Zn	30	65.37
Tungsten	W	74	183.85	Zirconium	Zr	40	91.22
Uranium	U	92	238.03				

Molecules

- Molecules are groups or clusters of atoms held together by chemical bonds.
- Molecules of an element
  - When the two or more atoms of an element are bonded together to form a molecule. Examples: $S_{8,}$ $H_{2,}$ $O_{2.}$
- Molecules involving more than one element are called a compound.
  - Examples include H2O , HCl, CH4.

The Molecular Weight

- The weight of a molecule, the *molecular weight, is the total mass of the individual* atoms. (Sum of the mass of the individual atoms.)

Example 1:

The compound water has a formula of H_2O. This means there is one atom of oxygen and two atoms of hydrogen. Calculate the molecular weight.

Solution:

The molecular weight is calculated as follows:

1 atom	×	16.000	(the atomic weight of oxygen)	=	16.000 amu
2 atoms	×	1.008	(the atomic weight of hydrogen)	=	2.016 amu
			molecular weight of water	=	18.016 amu

Calculating the Molecular Weight

Example 2:

Calculate the molecular weight of H_2SO_4.

Solution:

hydrogen	2	atoms	×	1.008	amu	=	2.016	amu
sulfur	1	atom	×	32.064	amu	=	32.064	amu
oxygen	4	atoms	×	15.999	amu	=	63.996	amu
				molecular weight		=	98.076	amu

Example 3:

Calculate the molecular weight of HCl.

Solution:

hydrogen	1	atom	×	1.008	amu	=	1.008	amu
chlorine	1	atom	×	35.453	amu	=	35.453	amu
				molecular weight		=	36.461	amu

Problems using Molecular Weight

- Calculate the molecular weight of CH_4.
- Calculate the molecular weight for CH_3OH.
- Calculate the molecular weight for NH_4OH.

Avogadro's Number

- The atomic weight of an atom depends on the number of protons and neutrons.
- The atomic weight of sulfur versus oxygen
  - Sulfur – 32 amu
  - Oxygen – 16 amu
- The atomic weight of sulfur is twice the atomic weight of oxygen. One gram of sulfur will contain less atoms than one gram of oxygen.
- Experimentation has shown that, for any element, a sample containing the atomic weight in grams contains 6.022×10^{23} atoms.
  - 15.99 grams of oxygen contain 6.022×10^{23} atoms of oxygen.
- This number (6.02×10^{23}) is know as Avogadro's Number.

Avogadro’s Number and the Mole

- Avogadro’s Number allows to relate the number of atoms per amount of material.
- A *mole* is defined as the quantity of a pure substance that contains 6.022 x 10^{23} units (atoms, ions, molecules, or elephants) of that substance.
- Deals with the macroscopic side of chemistry.
- For any element, the mass of a mole of that element's atoms is the atomic mass expressed in units of grams.
  - 1 mole of C-12 equals 12 grams
  - 1 mole of Na equals 22.981 grams

The Mole

- A *mole is defined as the quantity of a pure substance that contains 6.022 x 10^{23} units (atoms, ions,* molecules, or elephants) of that substance.
- Represents a large system. Chemical analysis is normally done in terms of moles of a substance.
- A mole of a given element is equivalent to the atomic weight in grams.
- For example, to calculate the mass of a mole of copper atoms, simply express the atomic mass of copper in units of grams.
  - The atomic mass of copper is 63.546 amu, then a mole of copper has a mass of 63.546 grams.

Using the Molecular Weight

- The atomic mass of gold is 196.967 amu, which makes the mass of one mole of gold equivalent to 196.967 grams.
- The mass of a mole of atoms is called the *gram atomic weight (GAW).*

Example from DOE Volume 1

Example 1:

A silver bar has a mass of 1870 grams. How many moles of silver are in the bar?

Solution:

Since the atomic mass of silver (Ag) is 107.87 amu, one mole of silver has a mass of 107.87 grams. Therefore, there is one mole of Ag per 107.87 grams of Ag or $\frac{1 \text{ mole Ag}}{107.87 \text{ grams Ag}}$. There are 1870 grams of silver.

$$\frac{1870 \text{ grams Ag}}{1} \times \frac{1 \text{ mole Ag}}{107.87 \text{ grams Ag}} = 17.3 \text{ mole Ag}$$

Example from DOE Volume 1

Example 2:

Mercury (Hg) is the only metal that exists as a liquid at room temperature. It isused in thermometers. A thermometer contains 0.004 moles of mercury. How many grams of mercury are in the thermometer?

Solution:

Since the atomic mass of Hg is 201 amu, one mole of Hg has a mass of 201 grams of Hg or $\frac{201 \text{ grams Hg}}{1 \text{ mole Hg}}$. There are 0.004 moles of Hg.

$$\frac{0.004 \text{ moles Hg}}{1} \times \frac{201 \text{ grams Hg}}{1 \text{ mole Hg}} = 0.8 \text{ grams Hg}$$

Molecular Weight

- The mass of a mole of molecules of a substance is the molecular mass expressed in grams.
- For example, an oxygen molecule (O_2) has a molecular mass equivalent to 32.0 grams because each oxygen atom has a molecular mass of 16.0 grams.
- Calculation is done by adding the atomic weight of each of the elements.
  - Calculate the moles
    - 269.5 g of Al_2O_3
    - 30.9 g of NaCl
    - 56.5 g of Fe_2O_3

The Periodic Table

- Over many years scientists have discovered that if elements are arranged in the order of their atomic numbers, the chemical properties of the elements are repeated somewhat regularly.
  – Physical properties are also repeated.

TABLE 3
Description of the Properties of the First Twenty Elements

Element	Symbol	Atomic Number	Atomic Weight	Description of Properties
Hydrogen	H	1	1.008	Colorless gas, reacts readily with oxygen to form H_2O; forms HCl with chlorine.
Helium	He	2	4.003	Colorless gas, very non-reactive chemically.

Properties of Elements

Lithium	Li	3	6.939	Silvery white, soft metal, very reactive chemically, forms Li_2O and LiCl readily.
Beryllium	Be	4	9.012	Grey metal, much harder than lithium, fairly reactive chemically, forms BeO and $BeCl_2$ easily.
Boron	B	5	10.811	Yellow or brown non-metal, very hard element, not very reactive, but will form B_2O_3, and BCl_3
Carbon	C	6	12.011	Black non-metal, brittle, non-reactive at room temperature. Forms CO_2 and CCl_4.
Nitrogen	N	7	14.007	Colorless gas, not very reactive, will form N_2O_5 and NH_3.
Oxygen	O	8	15.999	Colorless gas, moderately reactive,will combine with most elements, forms CO_2, MgO, etc.
Fluorine	F	9	18.998	Green-yellow gas, extremely reactive, irritating to smell, forms NaF, MgF_2.
Neon	Ne	10	20.183	Colorless gas, very non-reactive chemically.
Sodium	Na	11	22.990	Silvery soft metal, very reactive chemically, forms Na_2O and NaCl.
Magnesium	Mg	12	24.312	Silvery white metal, much harder than sodium. Fairly reactive, forms MgO and $MgCl_2$.
Aluminum	Al	13	26.982	Silvery white metal, like magnesium but not as reactive. Forms Al_2O_3 and $AlCl_3$.
Silicon	Si	14	28.086	Gray, non-metallic, non-reactive at room temperature, forms SiO_2 and $SiCl_4$.
Phosphorus	P	15	30.974	Black, red, violet, or yellow solid, low melting point, quite reactive, forms P_2O_5 and PCl_3.
Sulfur	S	16	32.064	Yellow solid with low melting point. Moderately reactive, combines with most elements, forms SO_2, MgS, etc.
Chlorine	Cl	17	35.453	Greenish-yellow gas, extremely reactive, irritating to smell, forms NaCl, $MgCl_2$.
Argon	Ar	18	39.948	Colorless gas, very non-reactive chemically.
Potassium	K	19	39.102	Silver soft metal, very reactive chemically, forms K_2O and KCl.
Calcium	Ca	20	40.080	Silver-white metal, much harder than potassium, fairly reactive, forms CaO and $CaCl_2$.

Periodic Table

- A table in which elements with similar chemical properties are grouped together is called a periodic table.
- Each horizontal row is called a *period.*
- Elements with similar chemical properties appear in vertical columns called *groups*.
- The number directly above each element is its atomic number, and the number below each element is its atomic weight.
  - If the atomic weight appear within parenthesis, this means that the element is unstable (radioactive).
  - The weight shown corresponds to the most stable isotope.

PERIODIC TABLE OF THE ELEMENTS

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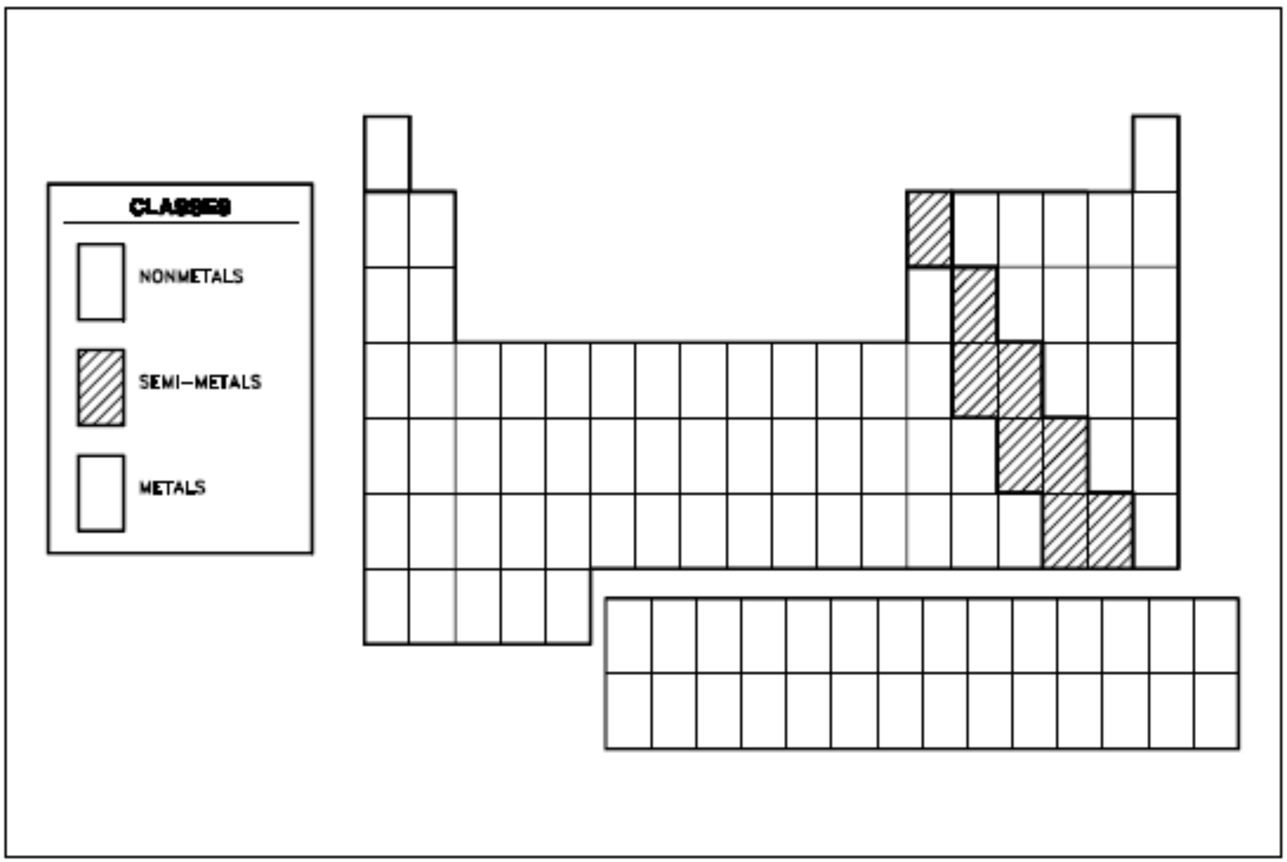
1A 1	2A 2	3B 3	4B 4	5B 5	6B 6	7B 7	8B 8	9	10	1B 11	2B 12	3A 13	4A 14	5A 15	6A 16	7A 17	8A 18
1 **H** Hydrogen 1.00794																	2 **He** Helium 4.00260
3 **Li** Lithium 6.941	4 **Be** Beryllium 9.01218											5 **B** Boron 10.811	6 **C** Carbon 12.011	7 **N** Nitrogen 14.0067	8 **O** Oxygen 15.9994	9 **F** Fluorine 18.998403	10 **Ne** Neon 20.1797
11 **Na** Sodium 22.98977	12 **Mg** Magnesium 24.305											13 **Al** Aluminum 26.98154	14 **Si** Silicon 28.0855	15 **P** Phosphorus 30.97376	16 **S** Sulfur 32.066	17 **Cl** Chlorine 35.4527	18 **Ar** Argon 39.948
19 **K** Potassium 39.0983	20 **Ca** Calcium 40.078	21 **Sc** Scandium 44.9559	22 **Ti** Titanium 47.88	23 **V** Vanadium 50.9415	24 **Cr** Chromium 51.9961	25 **Mn** Manganese 54.9380	26 **Fe** Iron 55.847	27 **Co** Cobalt 58.9332	28 **Ni** Nickel 58.6934	29 **Cu** Copper 63.546	30 **Zn** Zinc 65.39	31 **Ga** Gallium 69.723	32 **Ge** Germanium 72.61	33 **As** Arsenic 74.9216	34 **Se** Selenium 78.96	35 **Br** Bromine 79.904	36 **Kr** Krypton 83.80
37 **Rb** Rubidium 85.4678	38 **Sr** Strontium 87.62	39 **Y** Yttrium 88.9059	40 **Zr** Zirconium 91.224	41 **Nb** Niobium 92.9064	42 **Mo** Molybdenum 95.94	43 **Tc** Technetium (98)	44 **Ru** Ruthenium 101.07	45 **Rh** Rhodium 102.9055	46 **Pd** Palladium 106.42	47 **Ag** Silver 107.8682	48 **Cd** Cadmium 112.411	49 **In** Indium 114.82	50 **Sn** Tin 118.710	51 **Sb** Antimony 121.757	52 **Te** Tellurium 127.60	53 **I** Iodine 126.9045	54 **Xe** Xenon 131.29
55 **Cs** Cesium 132.9054	56 **Ba** Barium 137.327	57 ***La** Lanthanum 138.9055	72 **Hf** Hafnium 178.49	73 **Ta** Tantalum 180.9479	74 **W** Tungsten 183.85	75 **Re** Rhenium 186.207	76 **Os** Osmium 190.2	77 **Ir** Iridium 192.22	78 **Pt** Platinum 195.08	79 **Au** Gold 196.9665	80 **Hg** Mercury 200.59	81 **Tl** Thallium 204.3833	82 **Pb** Lead 207.2	83 **Bi** Bismuth 208.9804	84 **Po** Polonium (209)	85 **At** Astatine (210)	86 **Rn** Radon (222)
87 **Fr** Francium (223)	88 **Ra** Radium 226.0254	89 †**Ac** Actinium 227.0278	104 **Rf** Rutherfordium (261)	105 **Db** Dubnium (262)	106 **Sg** Seaborgium (263)	107 **Bh** Bohrium (262)	108 **Hs** Hassium (265)	109 **Mt** Meitnerium (268)	110 (269)	111 (272)	112 (277)						

*Lanthanide Series	58 **Ce** Cerium 140.115	59 **Pr** Praseodymium 140.9077	60 **Nd** Neodymium 144.24	61 **Pm** Promethium (145)	62 **Sm** Samarium 150.36	63 **Eu** Europium 151.965	64 **Gd** Gadolinium 157.25	65 **Tb** Terbium 158.9254	66 **Dy** Dysprosium 162.50	67 **Ho** Holmium 164.9303	68 **Er** Erbium 167.26	69 **Tm** Thulium 168.9342	70 **Yb** Ytterbium 173.04	71 **Lu** Lutetium 174.967
† Actinide Series	90 **Th** Thorium 232.0381	91 **Pa** Protactinium 231.0359	92 **U** Uranium 238.0289	93 **Np** Neptunium 237.048	94 **Pu** Plutonium (244)	95 **Am** Americium (243)	96 **Cm** Curium (247)	97 **Bk** Berkelium (247)	98 **Cf** Californium (251)	99 **Es** Einsteinium (252)	100 **Fm** Fermium (257)	101 **Md** Mendelevium (258)	102 **No** Nobelium (259)	103 **Lr** Lawrencium (260)

- Types of elements: Metals, Non-metals and Metalloids



Periodic Table

- Metals
  - Includes elements towards the left and center of the periodic table including the lanthanide and actinide series.
  - Metals normally donate electrons when involved in chemical reactions.
  - As a result, ions formed from metals have a positive charge (more protons than electrons).
  - They are usually hard and strong, capable of being shaped mechanically and good conductors of heat and electricity.
    - Lustrous surface when cleaned.

Periodic Table

- Metals can be involved in a wide variety of reactions.
  - Electron donors.
- Properties depend on the location in the periodic table.
- Metals are divided into the following two categories
  - The light metals, which are soft, have a low density, are very reactive chemically, and are unsatisfactory as structural materials. (Group I and IIA)
    - Includes sodium, potassium, magnesium and calcium.
  - The transition metals, which are hard, have a high density, do not react readily, and are useful structural materials. (Transition metals in the middle of the periodic table).
    - Includes elements like gold, platinum, nickel, copper and cobalt.
    - Transition metals are commonly used as catalysts.

Nonmetals

- The nonmetals occupy the part of the periodic table to the right of the heavy, step-like line.
- Nonmetals are often gases at room temperature.
- The nonmetals that are solids are not lustrous, are not malleable or ductile, and are poor conductors of heat and electricity.
  - Carbon conducts electricity but not as well as metals.
  - Some non-metals are very reactive.
  - Tend to gain electrons to form negative ions rather than to lose electrons to form positive ions.
  - The last group to the right is very un-reactive (known as the noble gases). Xenon is the only noble gas that is known to form chemical compounds. Under normal operating conditions the noble gases are considered inert.

Semi-Metals (Metalloids)

- Properties of elements change as we move from left to right and the change is not sharply defined.
- The elements located by the bold line are known as semi-metals.
- Exhibit properties of both metals and non-metals.
- Include boron (B), silicon (Si), germanium (Ge), arsenic (As), and tellurium (Te).
- They are usually classified as semi-conductors of electricity and are widely used in electrical components.

Group Characteristics

- On the left side of the table are Group IA (Alkali metals) elements (except hydrogen), which are soft metals that undergo similar chemical reactions. They loose one electron.
- The elements in Group IIA form similar compounds and are much harder than their neighbors in Group IA. Reactions with this group involves the transfer of two electrons. Known as the Alkaline Earth Family.
- All elements within a particular group have similar physical and chemical properties. (A gradual change is observed in each group). (Not necessarily the case for transition metals and the groups close to the bold line).

Group Characteristics

- Groups with a B designation (IB through VIIB) and Group VIII are called transition groups.
- Within any group in this region, all the elements are metals, but their chemical properties may differ.
- In some cases, an element may be more similar to neighbors within its period than it is to elements in its group.
  - For example, iron (Fe) is more similar to cobalt (Co) and nickel (Ni) than it is to ruthenium (Ru) and osmium (Os).
  - Most of these elements have several charges, and their ions in solution are colored (ions of all other elements are colorless).

Group Characteristics

- In Group IVA, for example, carbon (C) is a nonmetal; silicon (Si) and germanium (Ge) are semi-metals; and tin (Sn) and lead (Pb) are metals.
  - In this group they are not many similarities.
- Chemical Reactivity in the periodic table
  - Most active metals are the Alkali metals (Group IA). Lithium, potassium, cesium and francium are examples.
    - Activity of metals decrease as we move to the right of the periodic table.
  - Most active non-metals: The Halogen group (Group VIIA). Fluorine and chlorine are the most active nonmetals.
    - Activity of non-metals decreases as we move towards the left of the periodic table.

Atomic Structure of Electrons

- Chemical reactions involve the electrons in an element
  - Only elements in the outer most shell are involved.
- Electrons are in constant motion around the nucleus.
- They have both kinetic and potential energy, and their total energy is the sum of the two.
- The total energy is quantized; that is, there are definite, discrete values of total energy that atomic electrons can possess.
- These energy states can be visualized as spherical shells around the nucleus separated by forbidden areas where electrons cannot exist in a stable state.
- Each shell is referred with a number with the shell No. 1 referring to the innermost shell.
- The closer the electrons are to the nuclei the lower its energy state.

Atomic Structure of Electrons

Atomic Structure of Electrons

- The electron shells represent major energy states of electrons.
- Each shell contains one or more subshells called orbitals, each with a slightly different energy.
  - In order of increasing energy, the orbitals are designated by the small letters s, p, d, f, g, h.
  - the first shell contains an s orbital,
  - the second shell contains s and p orbitals,
  - the third shell contains s, p, and d orbitals,
  - the fourth shell contains s, p, d, and f orbitals, and so on.

Atomic Structure of Electrons

- Chemical reactions involve primarily the electrons in the outermost shell of an atom.
- The term outermost shell refers to the shell farthest from the nucleus that has some or all of its allotted number of electrons.
- All of the partially-filled shells have some effect on chemical behavior, but the outermost one has the greatest effect.
- The outermost shell is called the valence shell, and the electrons in that shell are called valence electrons.
- The term *valence* (of an atom) is defined as the number of electrons an element gains or loses, or the number of pairs of electrons it shares when it interacts with other elements.
- For the A group electrons the number of valence electrons equals the group number.

Atomic Structure of Electrons

- The arrangement in which the outermost shell is either completely filled (as with He and Ne) or contains eight electrons (as with Ne, Ar, Kr, Xe, Rn) is called the inert gas configuration.
- Hydrogen has features that are different from any other element, it can gain or loose an electron. If it looses an electron what is left is a bare nucleus.
- The hydrogen ion is very small in comparison with a positive ion of any other element, which must still have some electrons surrounding the nucleus.
- The number of electrons in the outer, or valence, shell determines the relative activity of the element.
- In general, the fewer electrons an element must lose, gain, or share to reach a stable shell structure, the more chemically active the element is.
- The likelihood of elements forming compounds is strongly influenced by this valence shell and on the stability of the resulting molecule.

Chemical Bonding

- The number of electrons in the valence shell determines the relative activity of an element.
- The arrangement of electrons in the outer shell explains why some elements are chemically very active, some are not very active, and others are inert.
- Group I has 1 valence electron, which makes it easy to loose that electron. Group VIIA has seven valence electrons and it only needs to gain one electron to become stable.
- The more stable the resulting molecules are, the more likely these molecules are to form.
- For example, an atom that "needs" two electrons to completely fill the valence shell would rather react with another atom which must give up two electrons to satisfy its valence.
- There are several types of chemical bonds that hold atoms together; three will be discussed, ionic, covalent, and metallic.

Ionic Bonds

- An *ionic bond* is formed when one or more electrons is wholly transferred from one element to another, and the elements are held together by the force of attraction due to the opposing charges.

Ionic Compounds

- Sodium donates an electron to the chlorine atom.
- As a result the sodium ion has a (+1) charge and chloride ion (-1 charge).
- In the ionic compound atoms arrange in a crystalline structure.
- Accordingly, the ionic bond is a force holding many atoms or ions together rather than a bond between two individual atoms or ions.

- Group IA looses one electron. (+1 charge)
- Group IIA looses 2 electrons (+2 charge)
- Group IIIA looses three electrons. (+3 charge)
- Group VIIA gains one electron. (-1 charge)
- Group VIA gains two electron. (-2 charge)
- Group VA gains three electrons. (-3 charge)

Naming Ionic Compounds

- For the cations, the name of the element is followed by the word ion. For example, the potassium ion or magnesium ion.
- For the monoatomic anions, the –ine termination is replaced by an –ide termination.
  - Examples: Oxide ion, sulfide ion, nitride ion, iodide ion, fluoride ion.
- NaF = sodium fluoride
- Na2S = sodium sulfide

Common Polyatomic Ions

- Sulfate ion = $SO4^{2-}$
- Nitrate ion = $NO3^{-}$
- Hydroxide ion = OH^{-}
- Ammonium Ion = $NH4^{+}$

Problems

- Write the ionic compound formed with the given ions.
  - Aluminum and Oxygen- Al2O3
  - Magnesium and Chlorine - MgCl2
  - Calcium and Hydroxide ion Ca(OH)2
  - Sodium and Oxygen Na2O
  - Lithium and the $SO4^{2-}$ Li2SO4

Covalent Bonds

- A *covalent bond* is formed when one or more electrons from an atom pair off with one or more electrons from another atom and form overlapping electron shells in which both atoms share the paired electrons.
- Covalent bonds hold the atoms together.
- Covalent bonding can be single covalent, double covalent, or triple covalent depending on the number of pairs of electrons shared.

Single Covalent Bond

Double Covalent Bond

Four pairs of electrons are shared by the carbon atom, two from each of the two oxygen atoms.

Four pairs of electrons are shared by the carbon atom, two from each of the two oxygen atoms

Coordinate Covalent Bond

Coordinate Covalent Bonds

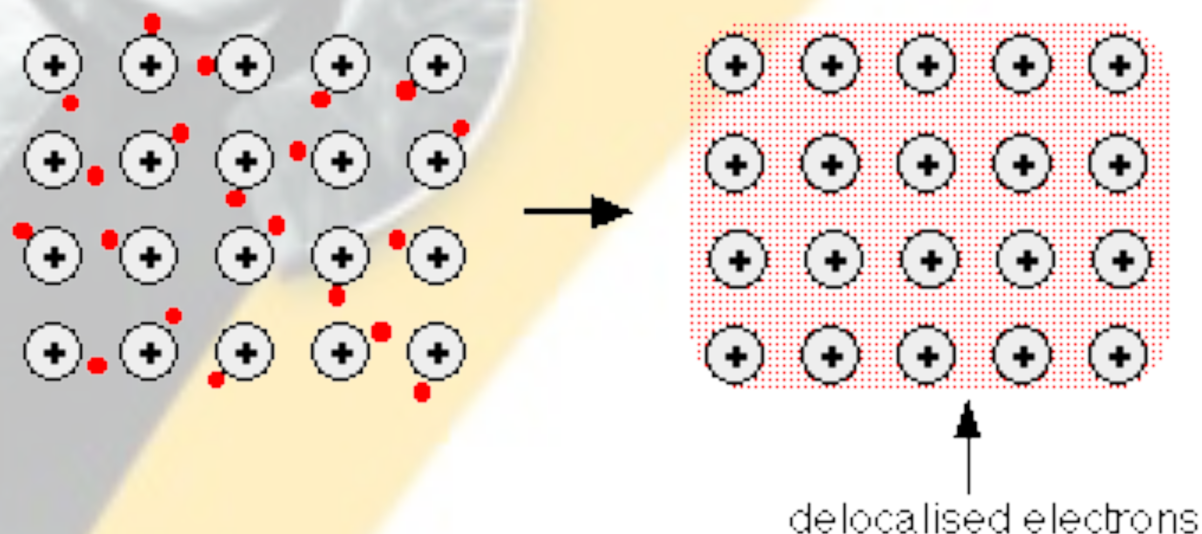
- In a coordinate covalent bond, both electrons come from the same atom. In the chlorate ion both electrons come from the Chlorine atom.
- Figure 9 illustrates the bonds of the negatively-charged chlorate ion.
- The ion consists of one chlorine atom and three oxygen atoms with a net charge of -1, and is formed with two coordinate covalent bonds and one covalent bond.
- The chlorine atom has effectively gained an electron through the covalent bond, which causes the overall negative charge.

Polarity of Molecules

- Covalent bonds can be either polar or non-polar.
- When the shared pair of electrons is not shared equally, one end of the bond is positive, and the other end is negative.
- An uneven share of electron is called a polar covalent bond. This are polar molecules.
- When two atoms of the same element are bonded they share the same attraction. As a result the bond is a non-polar covalent bond.
- If all the bonds in the molecule experience the same behavior the molecule is a non-polar molecules.

Metallic Bonds

- In the metallic bond, an atom achieves a more stable configuration by sharing the electrons in its outer shell with many other atoms.
- Prevail in elements in which the valence electrons are not tightly bound with the nucleus (by metals).
- In this type of bond, each atom in a metal crystal contributes all the electrons in its valence shell to all other atoms in the crystal.
  - Sodium has the electronic structure $1s^2 2s^2 2p^6 3s^1$. When sodium atoms come together, the electron in the 3s atomic orbital of one sodium atom shares space with the corresponding electron on a neighboring atom to form a molecular orbital - in much the same sort of way that a covalent bond is formed.
  - Each sodium atom is being touched by eight other sodium atoms - and the sharing occurs between the central atom and the 3s orbitals on all of the eight other atoms.


delocalised electrons

Metallic Bond

- The electrons can move freely within these molecular orbitals, and so each electron becomes detached from its parent atom. The electrons are said to be ***delocalised.*** The metal is held together by the strong forces of attraction between the positive nuclei and the delocalised electrons.
  - This is what makes metals conductors.
- The more electrons are involved in the metallic bond the stronger the interaction. In transition metals, the 3d electrons are also involved as well as the 4s electrons. As a result the metallic bond is stronger.

Van der Waals Forces

- Occurs within non polar molecules like H2 and Cl2.
- Believed to be formed due to a temporary dipole (unequal charge distribution) as electrons move about in the molecule.
- Van der Waals forces are the only forces experienced between non polar covalent bonds.
- It is important to note that van der Waals forces exist between all kinds of molecules.
  - Some molecules may have these forces, as well as dipole or other intermolecular forces.
- The strength of the van der Waals forces between substances increases with increasing gram molecular mass.

Van der Waals Forces

- In van der Walls forces, the temporary dipole induces a dipole in a neighboring molecule. The electrostatic force of attraction formed during this short dipole is the van der Waals forces.

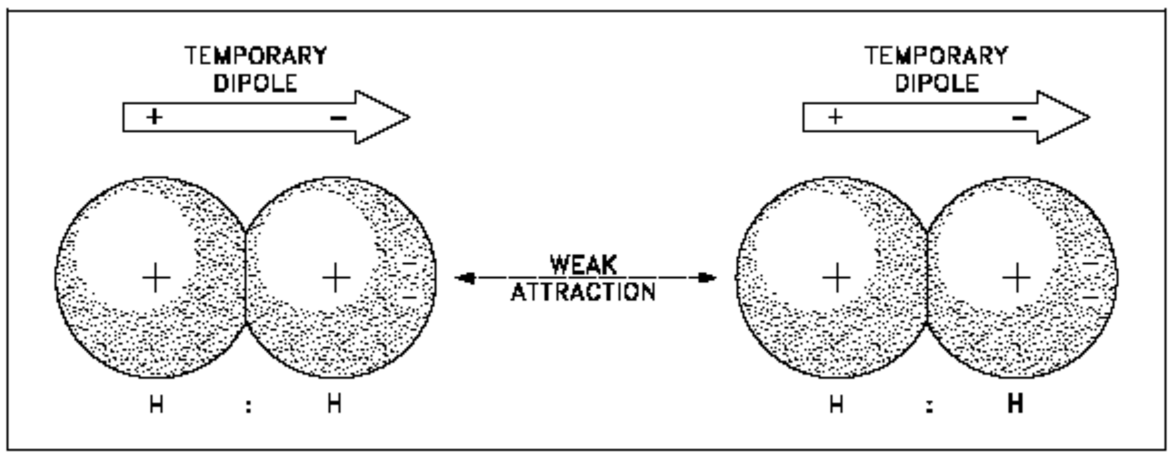


Figure 10 Van der Waals Forces

Organic Chemistry

- Organic chemistry deals with the chemistry of carbon containing compounds.
- This is a very broad area in chemistry that is normally subdivided into smaller subcategories.
- Depending on the groups attached to the carbon and the types of bonding is the subcategory that the compound falls on. This are known as families or classes.
- The carbon atoms can combine to form straight chains, rings, or branched chains.
- The bonds between carbon atoms can be single, double, triple or a combination of these.

Organic Chemistry

- If the chain only contains hydrogen and carbon they are called hydrocarbons.
  - Hydrocarbons can be divided as fatty or aromatic.
  - Aliphatic hydrocarbons are subdivided as saturated and unsaturated.
    - Alkanes, alkenes and alkynes.
  - Alkanes are completely saturated hydrocarbons (no double bonds). They contain the maximum number of hydrogens.
    - General form: CnH2n+2
- The alkanes are colorless, practically odorless, insoluble in water, and readily soluble in nonpolar solvents such as benzene or ether.

Alkanes

- Not as reactive.
- Reactions include halogenation, cracking (thermal decomposition) and combustion.
- In halogenation, a hydrogen atom is replaced by a halogen atom.
  - Example: $CH_4 + Br_2 \rightarrow CH_3Br + HBr$
- Thermal decomposition or cracking involves the decomposition by applying heat.
  - Example: $C_3H_8 \rightarrow CH_4 + C_2H_4$
- Combustion is the burning of a hydrocarbon in the presence of oxygen. It gives off CO2.

Alkenes

- Contain two fewer hydrogen atoms than alkanes.
- General formula: C_nH_{2n}
- Molecule have a double bond.
- Main source of alkenes is the cracking of alkanes.

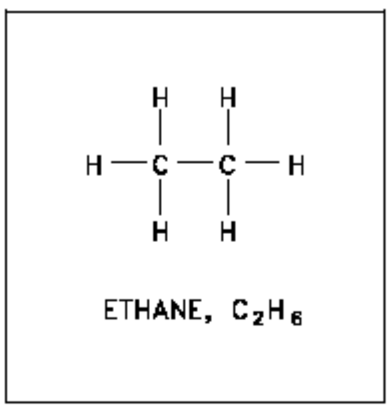


Figure 11 Alkane

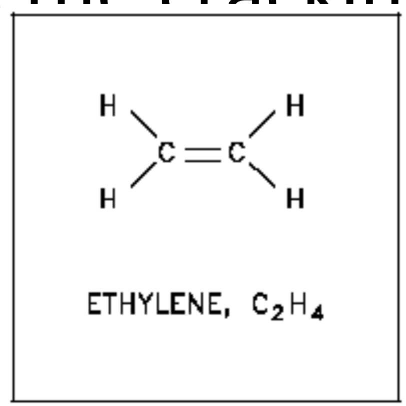


Figure 12 Alkene

Alkynes

- Contain two fewer hydrogens than the corresponding alkene.
- Formula: C_nH_{2n-2}
- Contain one triple bond.
- It is an unsaturated compound.

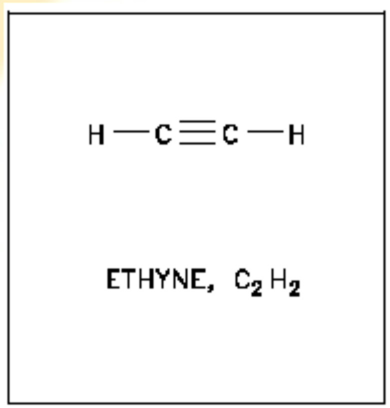


Figure 13 Alkyne

Aromatics

- They are not arranged in a straight chain. They are arranged in cyclical chains.
- They have pleasant odor that may be toxic.
- They have alternating single and double bonds like benzene.
- They undergo s reaction instead of additions.

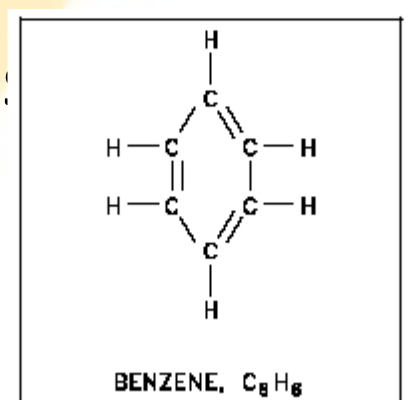


Figure 14 Aromatic

Alcohols

- Composed of aliphatic hydrocarbon with a hydrogen substituted by a hydroxyl group.
- The hydroxyl group does not behave like an ionic compound.
- Alcohols can be used for the synthesis of aliphatic hydrocarbons.

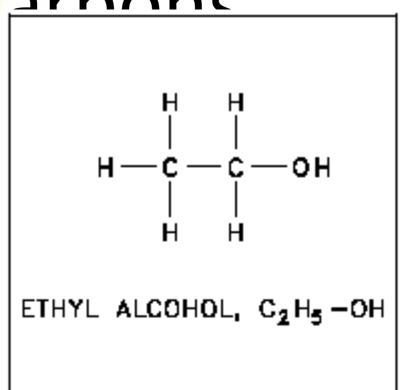


Figure 15 Alcohol

Aldehydes

- A product of the oxidation of alcoh
- Contain a carbonyl group (R-C=O).
- The term aldehyde comes from alc dehydrogenation.

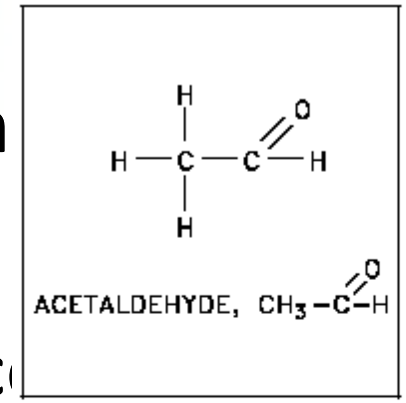


Figure 16 Aldehyde

- The alcohol gives off two hydrogen atoms to form the aldehyde.
- The carbonyl group is always at the end of the chain.

Basic Chemical Laws

- Molecules are groups or clusters of atoms held together firmly by means of chemical bonding.
  - Molecule of an element - Two single atoms of the same element, in certain cases, can become fastened to one another by a chemical bond to form a molecule.
  - Molecules of a compound - A compound contains at least two different kinds of atoms.
- Elements are substances that cannot be decomposed by ordinary types of chemical change nor made by chemical union.
- Compounds are substances containing more than one constituent element and having properties different from those of the individual elements.
- Mixtures consist of two or more substances intermingled with no constant percentage composition. Each component retains its original properties.

Forming Chemical Compounds

- Laws dealing with chemical reactions
  - The Law of Conservation of Mass
    - This law states that in a chemical reaction the total mass of the products equals the total mass of the reactants.
  - The Law of Definite Proportions
    - This law states that no matter how a given chemical compound is prepared, it always contains the same elements in the same proportions by mass.
  - The Law of Multiple Proportions
    - This law states that if two elements combine to form more than one compound, the masses of one of the elements combining with a fixed mass of the other are in a simple ratio to one another.
    - For example, H_2O and H_2O_2. For water the proportion is 2:1 and for hydrogen peroxide is 1:1. The molecular weight of water is 18 amu and for hydrogen peroxide is 34 amu. For water 1 grams of hydrogen are combined with 8 grams of oxygen. For hydrogen peroxide, 1 grams of hydrogen per 16 grams of oxygen.
      - Proportions are constant for each compound.

Combining Elements

- The Laws of Definite Proportions and Multiple Proportions and the related portions of atomic theory form the bases for most quantitative calculations involving chemical reactions.
- Regardless of the type of bond (ionic, covalent, coordinate covalent, or metallic), specific amounts of one element will react with specific amounts of the element(s) with which it is combined.
- If two substances are combined the result is a mixture.
  - Heterogeneous or homogeneous.

Combining Elements

- Homogeneous mixture is uniform throughout. For example: water and sugar.
- The solvent is the medium used to dissolve the sample.
  - Sugar water mixture – the solvent is water.
- The substance dissolved in the solution is known as the solute.
  - Air is a mixture of gases, metal alloys is a mixture of solids. Beer is an example of a liquid solution.

- *Solubility* is defined as the maximum amount of a substance that can dissolve in a given amount of solvent at a specific temperature. At this point, the solution is said to be saturated.

- *Equilibrium* is the point at which the rates of the forward and reverse reactions are exactly equal for a chemical reaction if the conditions of reaction are constant.

- Kinetics is the study of the factors which affect the rates of chemical reactions. There are five principle factors to consider: concentration, temperature, pressure, the nature of the reactants, and the catalyst.

Chemical Equations

- Le Chateliers Principle states that if a dynamic equilibrium is disturbed by changing the conditions, the position of equilibrium moves to counteract the change.
- If conditions like pressure, temperature and concentration are changed the equilibrium of the chemical reaction will change to counter the effect.
  - For an endothermic reaction an increase in temperature, will shift the equilibrium to the right (solubility will increase).

- For an exothermic reaction the reverse is true. Added heat will result in a decrease in solubility.
    - Solid ↔ solution + heat
- The amount of solute dissolved in a solution is very important. High accuracy is important.
- Properties like density, molarity, normality and parts per million.
- Density is a characteristic of a substance.
    - It is defined as mass divided by the volume.
    - It is not a function of the size and shape.

Weight Percent

- Concentration of solution

$$\text{Weight percent} = \frac{\text{weight of solute}}{\text{weight of solution}} \times 100\%$$

- Calculating

$$12\ \%\ \text{NaCl solution} = \frac{12\ \text{g NaCl}}{100\ \text{g solution}}$$

$$\frac{12\ \text{g NaCl}}{(12\ \text{g NaCl} + 88\ \text{g water})} = 12\%\ \text{NaCl solution}$$

Weight Percent Example 1

Question:
What is the weight percent of glucose in a solution made by dissolving 4.6 g of glucose in 145.2 g of water?

Analysis:
To get weight percent we need the weight of the solute and the total weight of the solution.

Determine total weight of solution:

4.6g glucose
+145.2 g water
149.8 g solution

Calculate percent

Weight % glucose =

$$\frac{4.6\text{ g glucose}}{149.8\text{ g solution}} \times 100 = 3.1\%\text{ glucose}$$

Weight Percent Example 2

- **Question:**
  How would you prepare 400. g of a 2.50% solution of sodium chloride?
- **Analysis:**
  We need to find out how much salt is needed and how much water is needed.
- **Determine weight of salt:**

  400. g x 2.50% salt = 10.0 g salt

  $$400.\ \text{g solution} \times \frac{2.50\ \text{g salt}}{100\ \text{g solution}} = 10.0\ \text{g salt}$$

- **Determine weight of water:**

  400. g total
  -10. g salt
  390. g water

- **Answer**

  Dissolve 10.0 g salt in 390. g water.

Weight Percent Questions

- What is the weight percent of ethanol in a solution made by dissolving 5.3 g of ethanol in 85.0 g of water?
- How would you make 250. g of a 7.5% solution of glucose in water?
- A sample of a solution weighing 850.0 g is known to contain .223 moles of potassium chloride. What is the weight percent of potassium chloride in the solution?

- Molarity
  - Defined as moles of solute divided by the liters of solution.
  - If a solution is said to be one molar, it contains one mole of solute per one liter of solution.

Example 1:

Prepare one molar solution of NaCl.

Solution:

a) Calculate the molecular weight of the salt

1 atom of Na	= 22.989 amu
1 atom of Cl	= 35.453 amu
1 molecule of NaCl	= 58.442 amu

One mole is equal to the gram molecular weight, so one mole = 58.442 grams.

b) 58.442 grams of NaCl is weighed out and sufficient water is added to bring the solution to one liter.

Problem

- Prepare 2.5L of 0.5M NaOH.
  - Steps: to solve the problem, calculate the moles by multiplying the molarity and the liters of solutions. Then multiply by the molecular weight (GAW) to obtain the grams needed to prepare the solution.
  - 2.5L*0.5M=1.25 moles
  - 1.25 moles * (23+16+1)=50 grams of NaOH
  - To prepare the solution, measure 50 grams of NaOH and dissolved until using volumetric glassware.

Normality

- Defined as equivalents of solute per liter of solution.

- Equivalents are based on the stoichiometry of the reaction.
- One equivalent of acid is the amount of acid necessary to give up one mole of hydrogen ions in a chemical reaction.
- One equivalent of base is the amount of base that reacts with one mole of hydrogen ions.

Normality

- When dealing with acids normality refers to the concentration of protons in solution.
- For bases, normality refers to the concentration of hydroxide ions.
- The *equivalent weight is defined as the molecular weight of the acid or base divided by the number* of replaceable hydrogen or hydroxyl ions.
- Sulfuric acid has 2 equivalents per mole (two hydrogen ions are transferred).
- For a one to one reaction: one equivalent per mole.

- Determine the equivalent weight for each of the reagents.

- How much sodium carbonate is needed to prepare 500.0mL of 0.50N Na_2CO_3?

Parts per Million

- The term *ppm is defined as the concentration of a solution in units of one part of solute* to one million parts solvent.
- 1 ppm is equivalent as 1 mg of solute per liter of solution.
- Term used for very dilute solutions.
- PPB represents parts of solute per billion of solution. Equivalent to micrograms per L of solution.

- Calculate the concentration of $CaCO_3$ in parts per million if 25mL of a 0.01266M EDTA solution was needed to test 100mL of water .

Chemical Equations

- A chemical equation is the representation of a chemical reaction using shorthand.
- Substances to the left are the reactants and to the right of the arrow the products.
  - $2H_2 + O_2 \rightarrow 2H_2O$
- The single arrow means that the reaction will only go in one direction.
- When writing an equation, always place the reactant on the left and the products on the right even in the case of a reversible reaction.
- A chemical equation represents not only the reaction, but also shows the number of atoms or molecules entering into and produced by the reaction.
- Formulas must be balanced (mass must be conserved).

Balancing Chemical Equations

- The number of atoms or molecules of each substance is shown by the coefficients in the equation.
- The number of atoms of each element must be the same at the left and right hand side of the reaction.

Balancing Chemical Reactions

Guidelines:

a. Once the correct chemical formula for a compound is written in an equation, do not modify it.

b. Select the compound with the greatest number of atoms. Then begin by balancing the element in that compound with the most atoms. There must be the same number of atoms of an element on each side of the equation. As a rule of thumb, this first element should not be hydrogen, oxygen, or a polyatomic ion.

c. Balance the atoms of each element in the compound by placing the appropriate coefficient in front of the chemical symbol or formula.

d. Next, balance the polyatomic ions. In some cases, the coefficient assigned in guideline 2 may have to be changed to balance the polyatomic ion.

e. Balance the hydrogen atoms next, then the oxygen atoms. If these elements appear in the polyatomic ion it should not be necessary to balance them again.

f. All coefficients will be whole numbers. The coefficients should be reduced to the lowest possible ratios.

g. As simple as it sounds, check off each element as it is accounted for since this will prevent double inclusion or a missed atom.

Example 1:

$$FeS_2 + O_2 \rightarrow Fe_2O_3 + SO_2$$

Chemical Reactions

- The mechanism for molecules to form involve more than a mere collision, it involves complex reaction mechanisms.
- Chemical equations do not show whether the reaction proceeds to completion or, if incomplete, the extent of reaction. In most cases, the substances that react never completely disappear; however, their concentration may be exceedingly small.
- Reactions that do not go to completion are usually represented in chemical equations by using double horizontal arrows .
- Chemical equations are very effective in representing chemical reactions on a macroscopic scale.
- The equation weight in grams of a compound or element is defined as the gram molecular weight times the number of molecules of the compound, as shown by the coefficients of the chemical equation for the reaction.
- The sums of the equation weights on each side of a chemical equation must be equal.

Solution:

The equation weight of iron equals the gram atomic weight of iron times the number of atoms shown reacting in the equation, which is two. Using Table 2:

$$\begin{aligned}\textbf{Equation Weight Fe} &= 2 \times 55.8 \text{ grams} \\ &= 111.6 \text{ grams}\end{aligned}$$

Because 27.9 g of iron react, the fraction of the equation weight that reacts is:

$$\frac{27.9 \text{ grams}}{111.6 \text{ grams}} = \frac{1}{4}.$$

Thus, 1/4 of the equation weight of ferric oxide will be formed.

The equation weight of ferric oxide equals the gram molecular weight of ferric oxide times the number of molecules shown formed in the equation, which is one.
Using Table 2:

$$\begin{aligned}\textbf{Equation Weight } Fe_2O_3 &= 2(55.8 \text{ g}) + 3(16.0 \text{ g}) \\ &= 111.6 \text{ g} + 48.0 \text{ g} \\ &= 159.6 \text{ g}\end{aligned}$$

Thus, the amount of ferric oxide formed is:

$$\frac{1}{4}(159.6 \text{ g}) = 39.9 \text{ g}.$$

- Solution using basic chemistry relating the coefficients of products and reactants:

Practice Problems

- An unknown sample is tested for the content of SO_4. It was found via precipitation using $BaCl_2$ that 4.35 g of $BaSO_4$ were precipitated out of solution. Calculate the amount of sulfate contained in the unknown.
  - Reaction: $BaCl_2 + SO_4 \longrightarrow BaSO_4$ + Chloride containing salt
- Calculate the amount of $CaCO_3$ produced if 24.2 g of Na_2CO_3 was consumed in the reaction.

Practice Problems

- Practice writing ionic compounds
  - Calcium hydroxide
  - Sodium carbonate
  - Magnesium nitrate
  - Calcium chloride
  - Aluminum chloride
- What is the charge of each of the ions?
  - $Fe(OH)_2$
  - $AgNO_3$
  - CuO
  - $Mg_3(PO_4)_2$

Acids, Bases, Salts and pH

EO 3.1 DEFINE the following terms:

a. Acid **b. Salt** **c.Base** **d. pH**

e. pOH **f. Dissociation constant of water**

e. Alkalies

EO 3.2 STATE the formula for pH.

EO 3.3 STATE the formula for pOH.

EO 3.4 CALCULATE the pH of a specified solution.

- Electrolytes are substances that when dissolved in water dissociate into ions.
- Electrolytes are divided into: salts, acids and bases.
- Acids are electrolytes that when dissolved in water dissociate to form hydrogen ions, H^+.
- Examples of acids include H2SO4, HCl, H3PO4.
- The letters in the chemical formula in the subscript inside parenthesis represent the state of the reactants.
- H_2SO_4 (aq)$\rightarrow$ $2H^+ + SO_4^{2-}$

Characteristics of Acids in Water

- Acid solutions taste sour.
- Acids react with many metals to produce hydrogen gas.
- $2HCl_{(aq)} + Zn_{(s)} \rightarrow ZnCl_{2\,(aq)} + H_2\,(g)$
- Turn litmus paper red.
- The solution conducts electricity.
- Acids react with bases to form a salt and water.
- $HCl_{(aq)} + NaOH_{(aq)} \rightarrow NaCl_{(aq)} + H2O_{(l)}$
- Acids react with carbonates to form CO2.
  - $CaCO_3 + 2HCl \rightarrow CaCl_2 + H_2O_{(l)} + CO_2\uparrow$

Characteristics of Bases

- Bases are ionic compounds that when dissolved in water produce hydroxide ions.
  - NaOH and KOH are common bases.
  - $NaOH (aq) \rightarrow Na+ + OH-$
- Common bases include ammonia and household soap. Characteristics of bases include:

  1. Basic solutions taste bitter and feel slippery to the touch.
  2. Bases turn litmus paper blue.
  3. Basic solutions conduct electricity.
  4. Bases neutralize acids.

- When an acid is neutralized by a base the product includes a salt.
- A *salt is an ionic* compound composed of positive ions and negative ions.
- The ionic bond is what keeps salts in their molecular form.
- The acid base reaction is known as ***neutralization***.
- NaCl is the product from the neutralization of NaOH and HCl.

- Salts may taste salty, sour, bitter, astringent, sweet, or tasteless.
- Solutions of salts may be acidic, basic, or neutral to acid-base indicators.
- Fused salts and aqueous solutions of salts conduct an electric current

pH and pOH

- Compounds that produce hydrogen ions directly when dissolved in water are called acids, and those that produce hydroxyl ions directly when dissolved in water are called bases.
- pH and pOH are calculated using the concentration of hydrogen or hydroxide ions in solutions, respectively.
- $pH = -\log [H^+]$
- $[H+] = 10-^{pH}$
- $pOH =-\log [OH-]$
- The negative in front of the logarithm is to ensure that the value of the pH is positive.

Example 1: The hydrogen concentration, $[H^+]$, of a solution is 4.8×10^{-8} moles/liter. What is the pH of the solution?

Solution 1:

$$pH = -\log [H^+]$$

$$= -\log (4.8 \times 10^{-8})$$

$$= 7.32$$

Example 2: The pH of a solution is 3.83. What is the hydrogen concentration of the solution?

Solution 2:

$$[H^+] = 10^{-pH}$$

$$= 10^{-3.83} \text{ moles/liter}$$

$$= 1.48 \times 10^{-4} \text{ moles/liter}$$

- The *pOH of a solution is defined as the negative logarithm of the hydroxyl* concentration, represented as [OH] in moles/ liter.
- $[H+][OH-]=1x10-^{14}$
- pH+pOH=14
- pH = 7 indicates a neutral solution.
- For a pH less than 7, the solution is acidic and for a pH greater than 7 the solution is basic.

Example: What is the hydrogen ion concentration $[H^+]$ and the hydroxyl concentration $[OH^-]$ in a solution with a pH of 5.5?

Solution:

$$[H^+] = 10^{-pH}$$
$$= 10^{-5.5}$$
$$= (10^{-6})(10^{.5})$$
$$= 3.16 \times 10^{-6} \text{ moles/liter}$$

$$\text{since} \quad pOH + pH = 14$$
$$pOH = 14 - 5.5$$
$$pOH = 8.5$$

$$\text{and} \quad [OH^-] = 10^{-pOH}$$
$$= 10^{-8.5}$$
$$= (10^{-9})(10^{.5})$$
$$= 3.16 \times 10^{-9} \text{ moles/liter}$$