The Islamic University of Gaza- Environmental Engineering Department

Environmental Chemistry I (EENV 2301)

CHAPTER 1: INTRODUCTION TO ENVIRONMENTAL CHEMISTRY

Chemistry and Environmental Chemistry

The Building Blocks of Matter

Prepared by

Husam Al-Najar

Objectives of the course:

The course address the environmental chemistry of the hydrosphere, discusses the fundamental properties of water, properties of bodies of water, and basic aquatic chemistry, including acid base behavior, phase interactions, oxidation-reduction, chelation. Moreover it introduces the atmosphere and atmospheric chemistry, including the key concept of photochemistry. Inorganic air pollutants, including nitrogen and sulfur oxides, carbon monoxide, and carbon dioxide. The course deals with soil and discuss anthrospheric aspects of environmental chemistry.

Text Book:

Stanley E. Manahan, Fundamentals of Environmental Chemistry, 3rd ed., Taylor & Francis/CRC Press, 2009

References:

Jorge G. Ibanez, Margarita Hemandez-Esparza, Carmen Doria-Serrano, Arturo Fregoso-Infante and Mono Mohan Singh. 2007. Environmental Chemistry Fundementals. Springer.

Ronald A. Hites 2007. Elements of Environmental Chemistry. WILEY-INTERSCIENCE A JOHN WILEY & SONS, INC., PUBLICATION.

Roy M. Harrison. Understanding Our Environment An Introduction to Environmental Chemistry and Pollution. ISBN 0-85404-584-8. The Royal Society of Chemistry 1999

John Wright . 2005. Environmental Chemistry. Taylor & Francis e-Library, 2005.

Eric Lichtfouse, Jan Schwarzbauer and Didier Robert (Editors). Green Chemistry and Pollutants in Ecosystems. ISBN 3-540-22860-8. Springer Berlin Heidelberg New York.

Course Outline

CHAPTER 1: INTRODUCTION TO ENVIRONMENTAL CHEMISTRY
1.1   Chemistry and Environmental Chemistry
1.2   The Building Blocks of Matter
1.3 Chemical Bonds, Compound Formation and Octet Rule

CHAPTER 2: ACIDS, BASES, AND SALTS
2.1 The Importance and nature of Acids, Bases, and Salts
2.2 Dissociation of Acids and Bases in Water
2.3 pH and the Relationship Between Hydrogen Ion and Hydroxide Ion Concentrations
2.4 Preparation of Acids, Bases and Salts

CHAPTER 3: SOLUTIONS
3.1 The Solution Process, Solubility and Concentration
3.2 Standard Solutions and Titrations
3.3 Solution Equilibria
3.4 Colloidal Suspensions

CHAPTER 4: ENVIRONMENTAL CHEMISTRY OF WATER
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4.2 Metal Ions and Calcium in Water
4.3 Oxidation-Reduction - Complexation and Chelation and Water Interactions with Other Phases

Midterm Exam

CHAPTER 5: WATER POLLUTION

5.1 Nature and Types of Water Pollutants
5.2 Elemental Pollutants, Heavy Metal, Metalloid
5.3 Organically Bound Metals and Metalloids
5.4 Inorganic Species
5.5 Oxygen, Oxidants, and Reductants
5.6 Organic Pollutants, Pesticides in Water and Polychlorinated Biphenyls

CHAPTER 6: THE ATMOSPHERE AND ATMOSPHERIC CHEMISTRY

6.1 The Atmosphere and Atmospheric Chemistry
6.2 Physical Characteristics of the Atmosphere
6.3 Energy Transfer in the Atmosphere
6.4 Atmospheric Mass Transfer, Meteorology, and Weather
6.5 Inversions and Air Pollution
6.6 Chemical and Photochemical Reactions in the Atmosphere
6.7 Acid–Base Reactions in the Atmosphere
6.8 Reactions of Atmospheric Oxygen and Nitrogen

CHAPTER 7: SOIL ENVIRONMENTAL CHEMISTRY

7.1 Nature and Composition of Soil
7.2 Acid-Base and Ion Exchange Reactions in Soils
7.3 Macronutrients and Micronutrients in Soil
7.4 Fertilizers
7.5 Wastes and Pollutants in Soil
7.6 Soil Loss and Degradation

Final Exam

Grades: Midterm 30% Homework and Quizzes 20% Final Exam 50%

Chemistry and Environmental Chemistry

Definitions:

Environmental Science: "Environmental Science is the study of the environment, its living and nonliving components, and the interactions of these components." Daniel D. Chiras, *Environmental Science*, 3rd Ed.

Ecology: "Ecology is the scientific study of the relationships between organisms and their environments." S. J. McNaughton and L. L. Wolf, *General Ecology*

Environmental Studies: "Environmental studies is the discipline dealing with the social, political, philosophical and ethical issues concerning man's interactions with the environment." Roderick Nash, University of California, Santa Barbara

Environmentalist: "A person working to solve environmental problems, such as air and water pollution, the exhaustion of natural resources, uncontrolled population growth, etc." Webster's New World Dictionary, Second College Edition.

Ecological Chemistry: "Ecological chemistry is the study of the interactions between organisms and their environment that are mediated by naturally occurring chemicals." International Society of Chemical Ecology

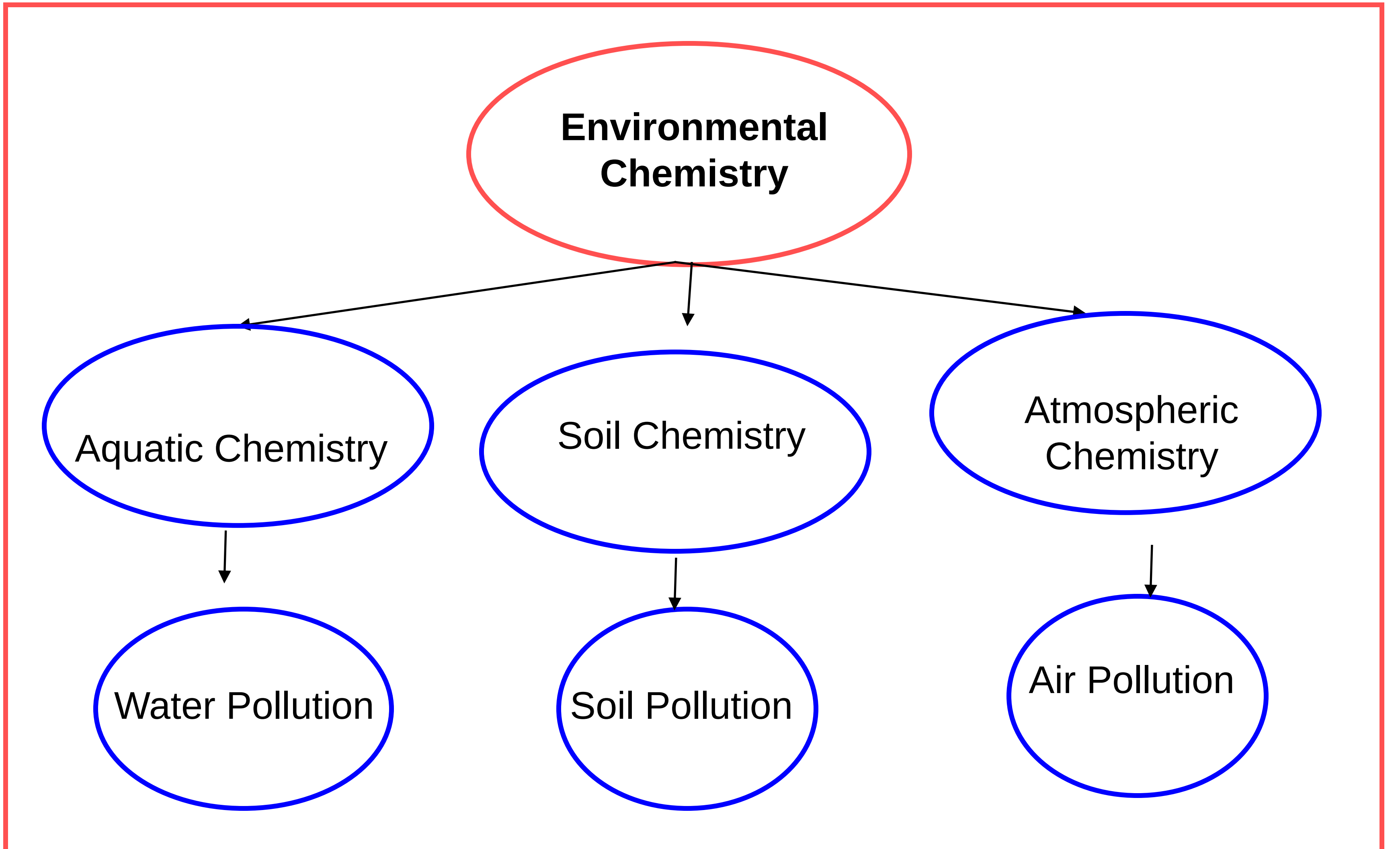
Environmental Biochemistry: "Environmental Biochemistry is the discipline that deals specifically with the effects of environmental chemical species on life

Toxicological Chemistry: "Toxicological chemistry is the chemistry of toxic substances with emphasis upon their interactions with biologic tissue and living organisms." Stanley Manahan*,* *Toxicological Chemistry*, 2nd Ed.

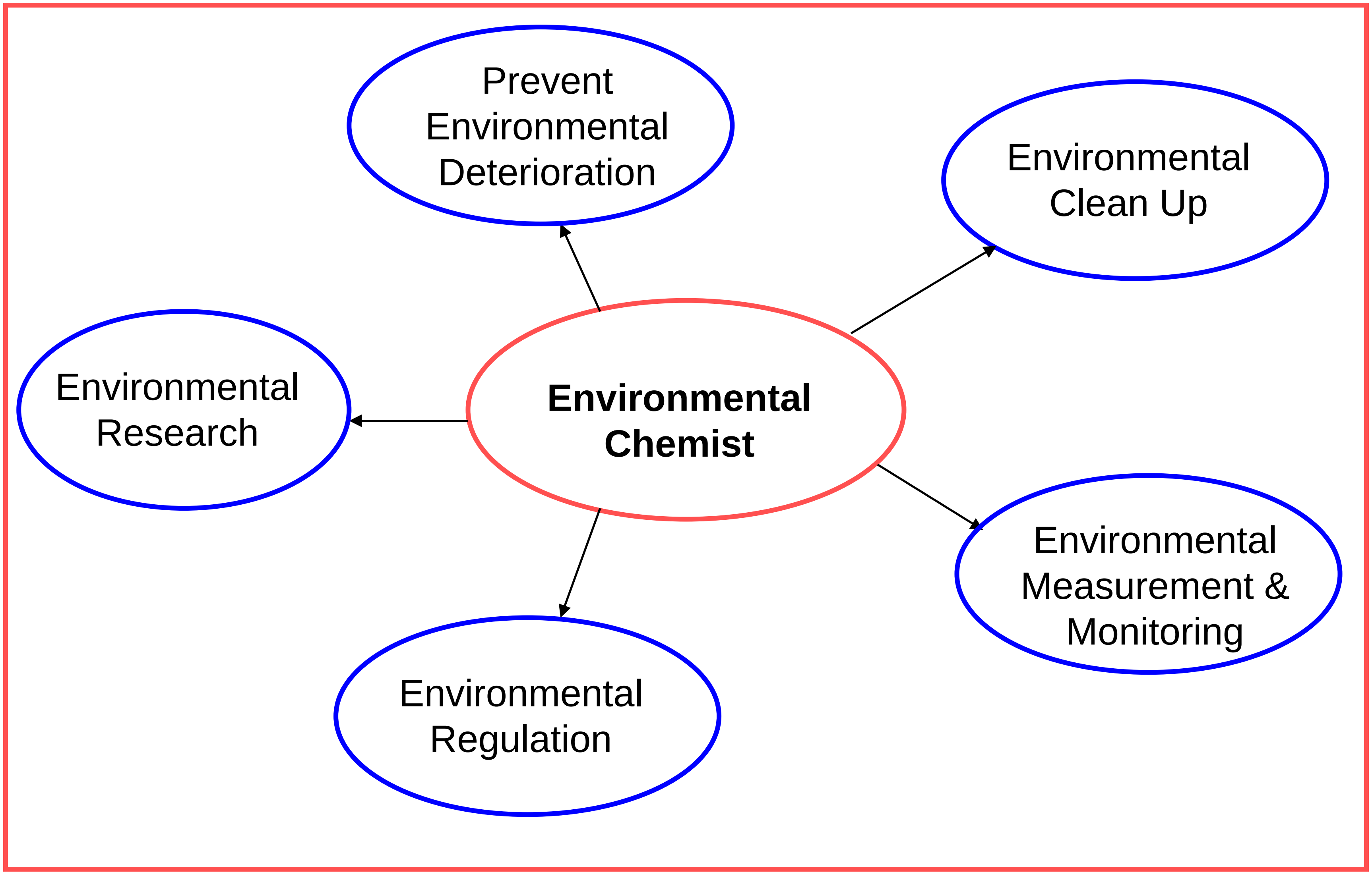
Environmental Analytical Chemistry: Environmental Analytical Chemistry is the application of analytical chemical techniques to the analysis of environmental samples--in a regulatory setting. Richard Foust, *Environmental Analytical Chemistry*, 1st Ed.

Environmental Chemistry: "Environmental chemistry is the study of the sources, reactions, transport, effects, and fates of chemical species in water, soil, and air environments, and the effects of technology thereon." Stanley Manahan, *Environmental Chemistry*, 8h Ed.

Illustration of Environmental Chemistry definition



The Mission of Environmental Chemist



The problems that environmental chemists study are often health related, and the driving force in much of this research is to provide a healthier environment.

Discoveries from basic research are used to identify and define issues, which over time, result in legislation.

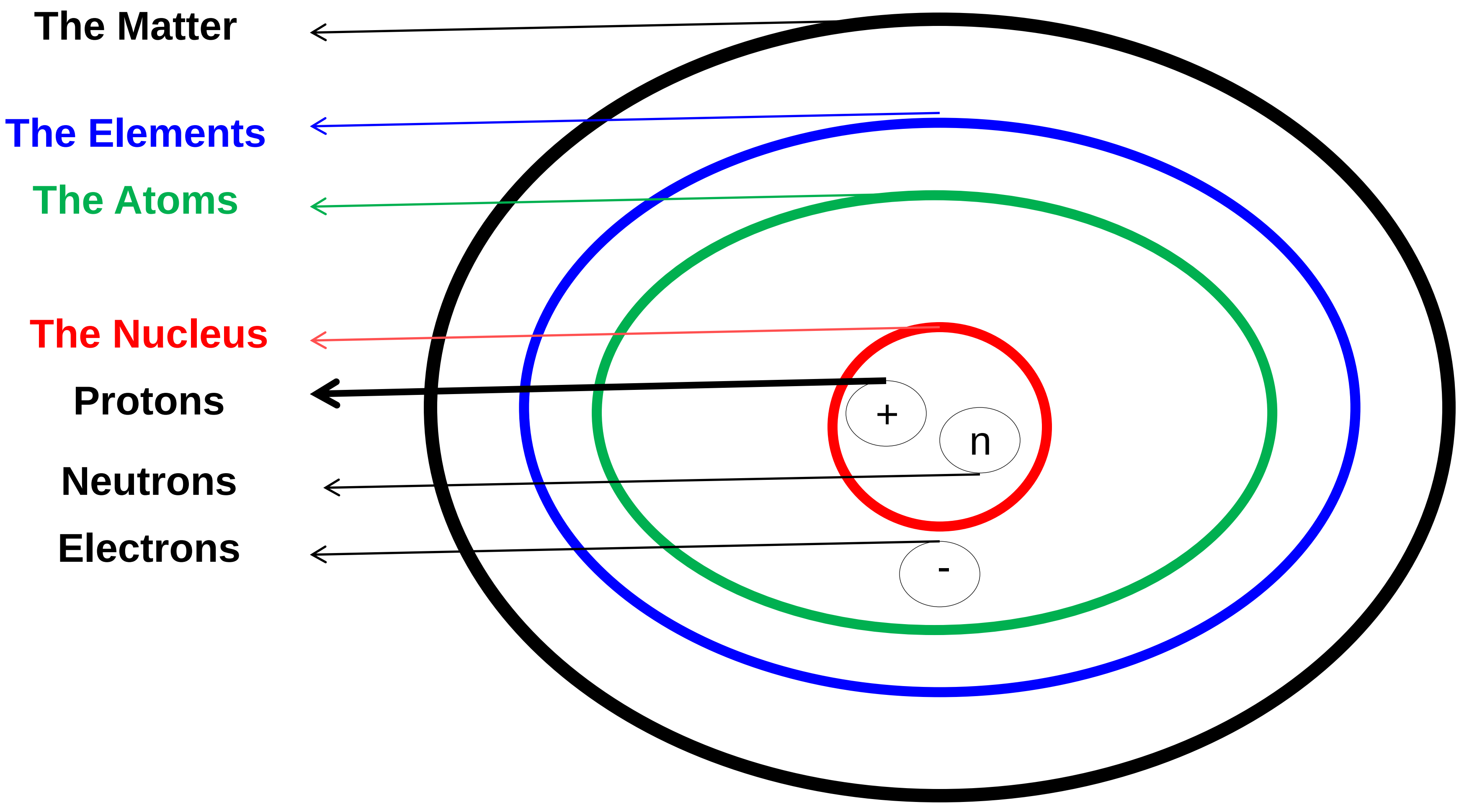
The purpose of environmental legislation is to modify human behavior so as to reduce or eliminate the environmental threats identified through basic research.

Environmental chemists participate in all aspects of this process from collecting the data for basic research, to monitoring environmental quality to developing chemical processes for remediation and environmental cleanup.

Environmental chemists work at the interface of chemistry with biologists, geologists, atmospheric scientists, engineers, lawyers and legislators.

The Building Blocks of Matter

Structure of the Matter



All **matter** is composed of only about a hundred fundamental kinds of matter called **elements.**

Each **element** is made up of very small entities called **atoms**

All atoms of the same element behave identically chemically.

Example: An atom of deuterium, a form of the element hydrogen.
such an atom is made up of even smaller subatomic particles:
Positively charged protons
Negatively charged electrons
Uncharged (neutral) neutrons.

Protons and neutrons have relatively high masses compared with electrons and are contained in the positively charged nucleus of the atom.

The nucleus has essentially all the mass, but occupies virtually none of the volume, of the atom.

An uncharged atom has the same number of electrons as protons.

The electrons in an atom are contained in a cloud of negative charge around the nucleus that occupies most of the volume of the atom.

The atomic number of the element is the number of protons in the nucleus of each atom of an element.

Atomic mass (atomic weight) is the average mass of all atoms of the element, including the various isotopes of which it consists.

The atomic mass unit, u (also called the dalton), is used to express masses of individual atoms and molecules (aggregates of atoms).

Mass Number is the average atomic mass rounded to the nearest whole number, therefore it is the total number of protons and neutrons in an atom's nucleus.

Example: Each C atom has 6 protons (+) in its nucleus,
the atomic number of C is 6.
The atomic mass of C is 12.

Ions: atoms that have lost or gained electrons are called ions

Positive Ions: when an atom loses electrons

$_{11}Na^{23}$ Protons =11 and Electrons = 11

loss 1 Electron $\longrightarrow$ Protons =11 and Electrons = 10 $\longrightarrow$ $^{23}Na_{11}{}^{+}$

Negative Ions: when an atom gains electrons

$_{9}F^{19}$ Protons = 9 and Electrons = 9

gain 1 Electron $\longrightarrow$ Protons = 9 and Electrons = 10 $\longrightarrow$ $^{19}F_{9}^{-}$

Isotopes: atoms with the same number of protons (atomic number is the same) but different numbers of neutrons (mass number is different).

Usually isotopes are referred to by their name (of symbol) and their mass number. Every element has at least 2 isotopes and some elements have as many as 25 isotopes.

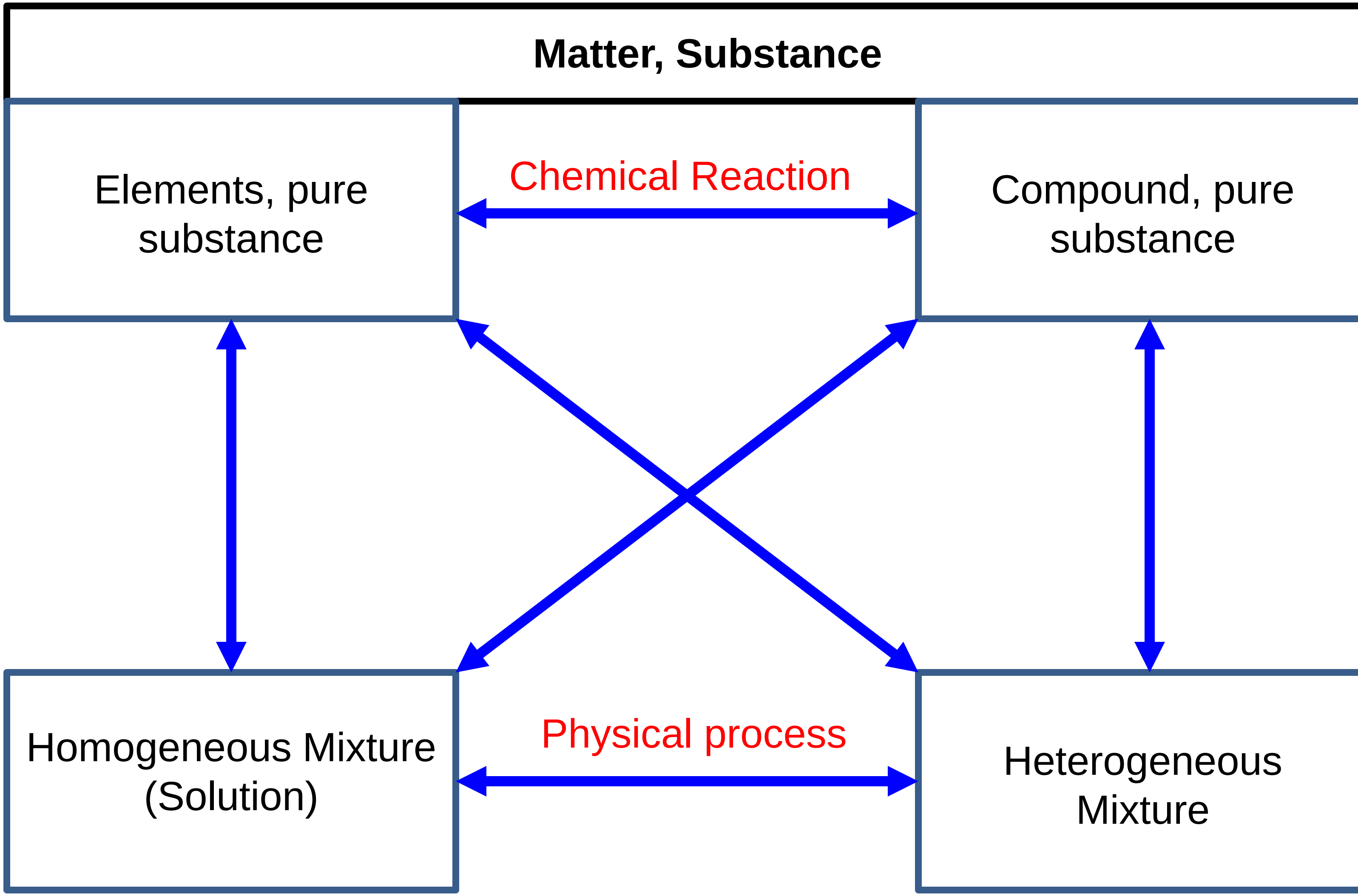
Example: The isotopes of hydrogen have separate names rather than being called hydrogen-1, hydrogen-2, etc. Their names are protium (H-1), deuterium (H-2), and tritium (H-3).

Atoms of most elements consist of two or more **isotopes** that have different numbers of neutrons in their nuclei.

Some isotopes are **radioactive** isotopes or **radionuclides**, which have unstable nuclei that give off charged particles and gamma rays in the form of radioactivity.

This process of radioactive decay changes atoms of a particular element to atoms of another element.

Classification of the Matter



Some General Types of Matter

Elements are divided between metals and nonmetals

several elements with properties of both metals and nonmetals are called **metalloids**.

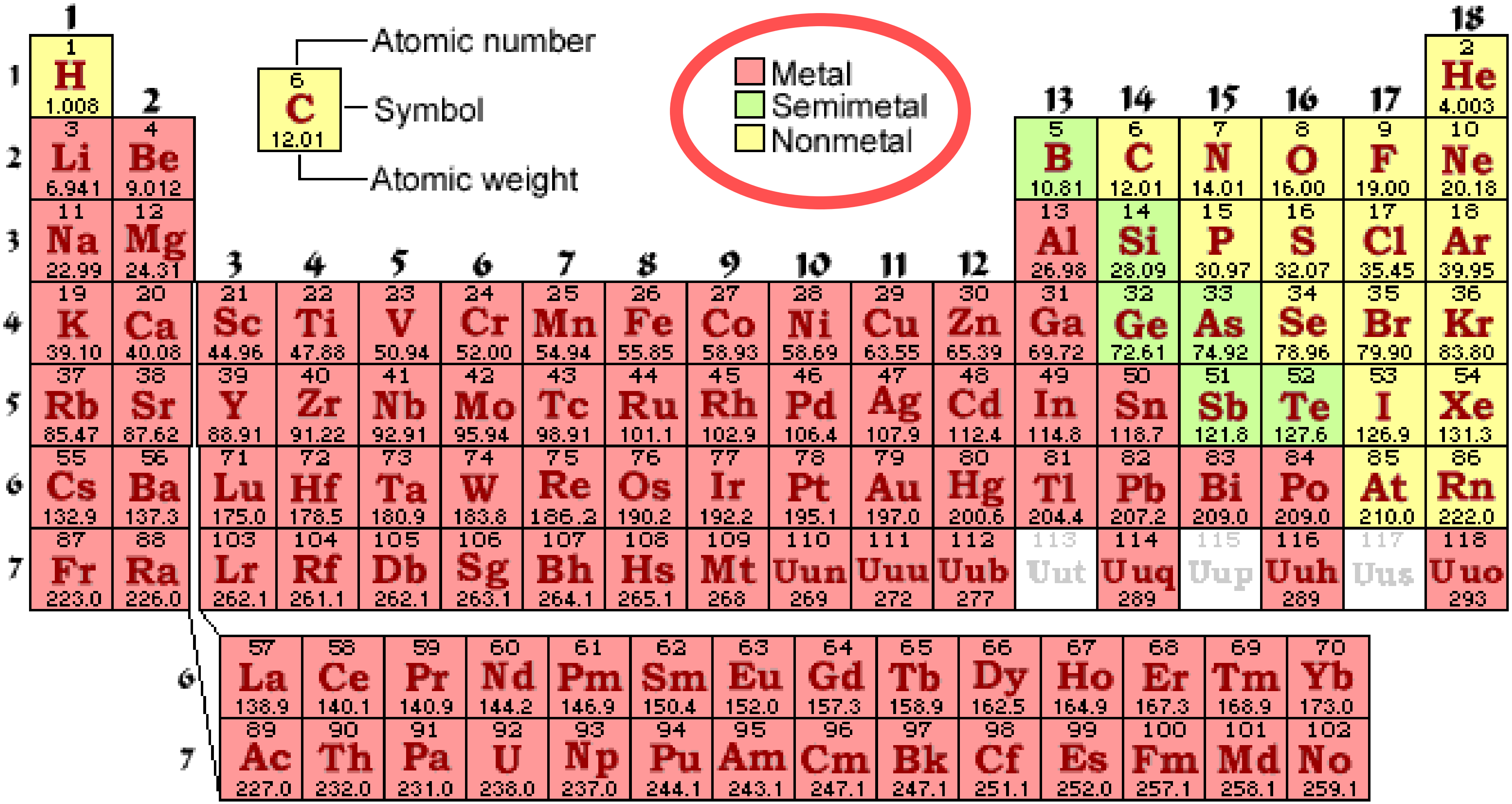
Metals are elements that are generally solid, shiny in appearance, electrically conducting, and malleable (ليونة) —that is, they can be pounded into flat sheets without disintegrating.

Nonmetals often have a dull appearance and are not at all malleable.

nonmetals frequently occur as gases or liquids. Colorless oxygen gas, green chlorine gas and brown bromine liquid are common nonmetals.

Organic substances consist of virtually all compounds that contain carbon, including substances made by life processes (wood, flesh, cotton, wool), petroleum, natural gas (methane), solvents (dry cleaning fluids), synthetic fibers, and plastics.

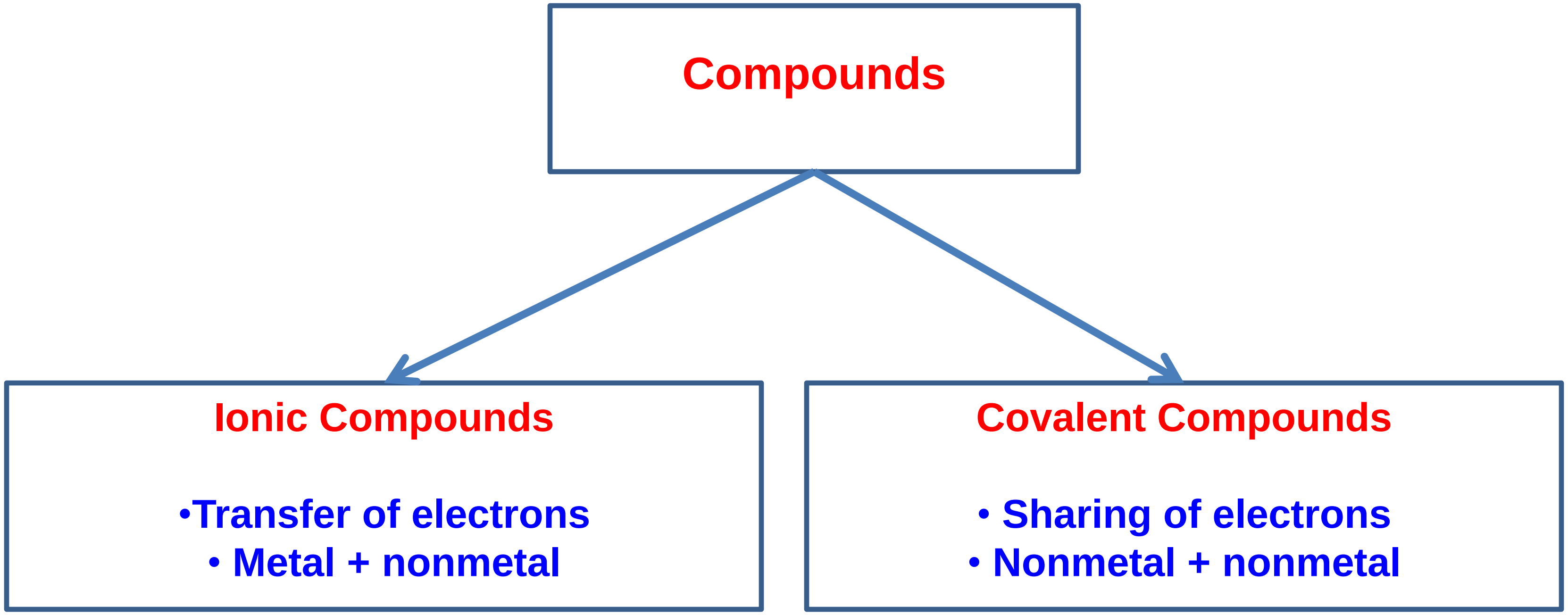
All of the rest of the chemical kingdom is composed of **inorganic substances** made up of virtually all substances that do not contain carbon.

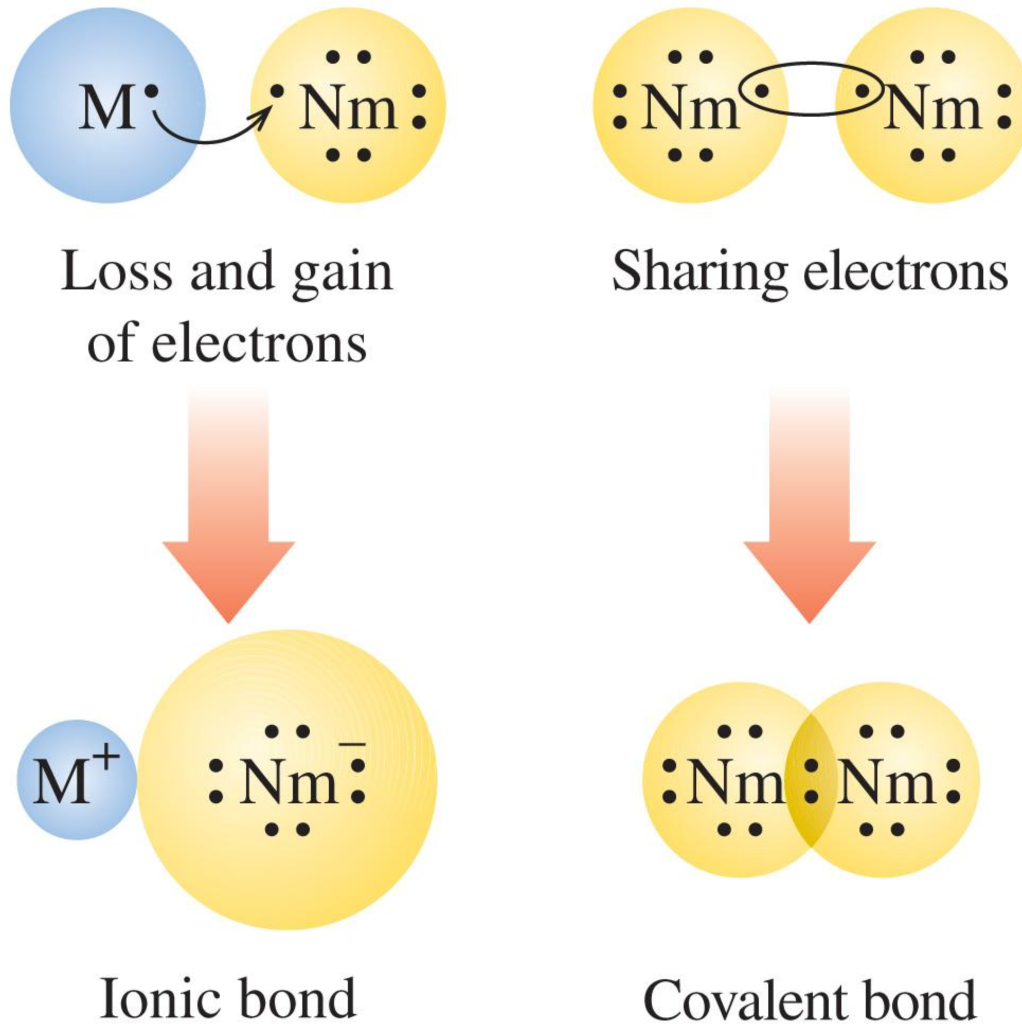

Atomic number
Symbol
Atomic weight
6 C 12.01
Metal
Semimetal
Nonmetal
1 H 1.008
2 He 4.003
3 Li 6.941
4 Be 9.012
5 B 10.81
6 C 12.01
7 N 14.01
8 O 16.00
9 F 19.00
10 Ne 20.18
11 Na 22.99
12 Mg 24.31
13 Al 26.98
14 Si 28.09
15 P 30.97
16 S 32.07
17 Cl 35.45
18 Ar 39.95
19 K 39.10
20 Ca 40.08
21 Sc 44.96
22 Ti 47.88
23 V 50.94
24 Cr 52.00
25 Mn 54.94
26 Fe 55.85
27 Co 58.93
28 Ni 58.69
29 Cu 63.55
30 Zn 65.39
31 Ga 69.72
32 Ge 72.61
33 As 74.92
34 Se 78.96
35 Br 79.90
36 Kr 83.80
37 Rb 85.47
38 Sr 87.62
39 Y 88.91
40 Zr 91.22
41 Nb 92.91
42 Mo 95.94
43 Tc 98.91
44 Ru 101.1
45 Rh 102.9
46 Pd 106.4
47 Ag 107.9
48 Cd 112.4
49 In 114.8
50 Sn 118.7
51 Sb 121.8
52 Te 127.6
53 I 126.9
54 Xe 131.3
55 Cs 132.9
56 Ba 137.3
71 Lu 175.0
72 Hf 178.5
73 Ta 180.9
74 W 183.8
75 Re 186.2
76 Os 190.2
77 Ir 192.2
78 Pt 195.1
79 Au 197.0
80 Hg 200.6
81 Tl 204.4
82 Pb 207.2
83 Bi 209.0
84 Po 209.0
85 At 210.0
86 Rn 222.0
87 Fr 223.0
88 Ra 226.0
103 Lr 262.1
104 Rf 261.1
105 Db 262.1
106 Sg 263.1
107 Bh 264.1
108 Hs 265.1
109 Mt 268
110 Uun 269
111 Uuu 272
112 Uub 277
113 Uut
114 Uuq 289
115 Uup
116 Uuh 289
117 Uus
118 Uuo 293
57 La 138.9
58 Ce 140.1
59 Pr 140.9
60 Nd 144.2
61 Pm 146.9
62 Sm 150.4
63 Eu 152.0
64 Gd 157.3
65 Tb 158.9
66 Dy 162.5
67 Ho 164.9
68 Er 167.3
69 Tm 168.9
70 Yb 173.0
89 Ac 227.0
90 Th 232.0
91 Pa 231.0
92 U 238.0
93 Np 237.0
94 Pu 244.1
95 Am 243.1
96 Cm 247.1
97 Bk 247.1
98 Cf 251.1
99 Es 252.0
100 Fm 257.1
101 Md 258.1
102 No 259.1

Chemical Bonds, Compound Formation and Octet Rule

Atoms with eight valence electrons are particularly stable, an observation called the octet rule.

Atoms form bonds with other atoms to achieve a valence octet.

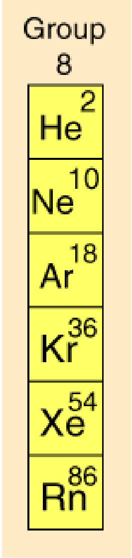




M is a metal

Nm is a nonmetal

CHEMICAL BONDS AND COMPOUNDS



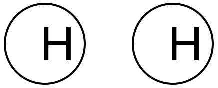
Only a few elements, particularly the noble gases, exist as individual atoms

Most atoms are joined by chemical bonds to other atoms.

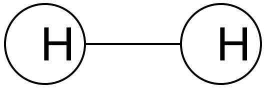
Elemental hydrogen exists as **molecules**, each consisting of 2 H atoms linked by a chemical bond.

Because hydrogen molecules contain 2 H atoms, they are said to be **diatomic** and are denoted by the chemical formula H_2.

The H atoms in the H_2 molecule are held together by a covalent bond made up of 2 electrons, each contributed by one of the H atoms, and shared between the atoms.



The H atoms in elemental hydrogen



are held together by chemical bonds in molecules

H_2

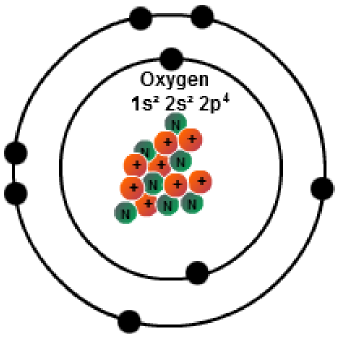
that have the chemical formula H_2.

Chemical Compounds

Most **substances** consist of two or more **elements** joined by chemical bonds.

As an example, consider the chemical combination of the elements hydrogen and oxygen

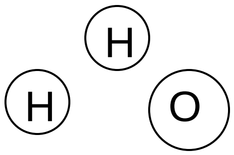
Oxygen, has an atomic number of 8 exists in the elemental form as diatomic molecules of O_2.

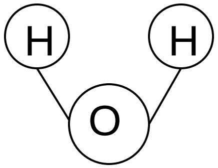


A substance such as H_2O that consists of a chemically bonded combination of two or more elements is called a chemical compound.

A chemical compound is a substance that consists of atoms of two or more different elements bonded together.

Each of the chemical bonds holding a hydrogen atom to the oxygen atom in the water molecule is composed of two electrons shared between the hydrogen and oxygen atoms.





H_2O

A molecule of water, H_2O, formed from 2 H atoms and 1 O atom held together by chemical bonds.

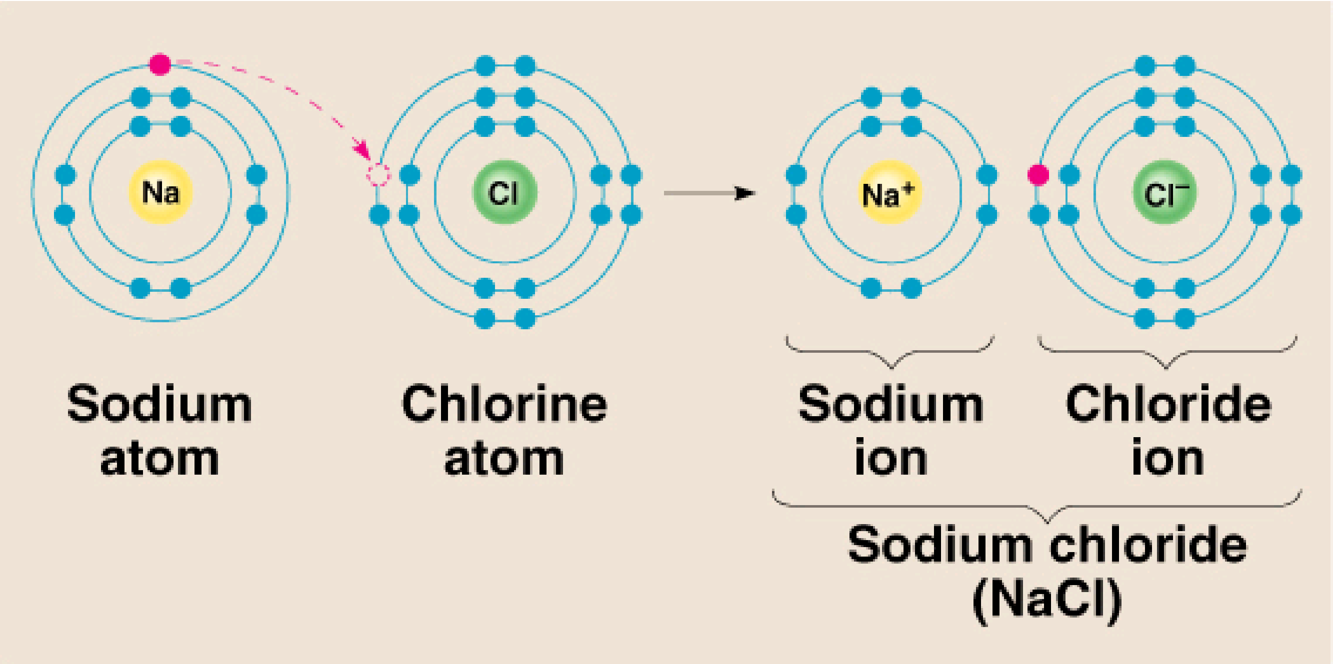
Ionic Bonds:

The **transfer of electrons** from one atom to another produces charged species called **ions**.

Positively charged ions are called **cations** and **negatively charged ions** are called **anions**.

The attracting forces of the oppositely charged ions in the crystalline lattice constitute the ionic bonds in the compound.

Example:



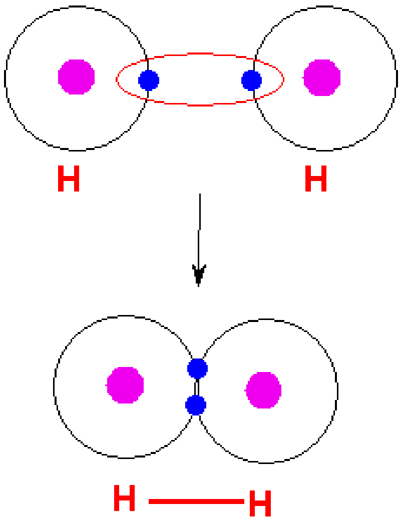
Sodium atom		Sodium ion

Sodium atom		Sodium ion
$Na^{\bullet}$ $- e^{-}$	$\longrightarrow$	Na^{+}
2-8-1		2-8 (= Ne)
$11\ p^{+}$		$11\ p^{+}$
$\underline{11\ e^{-}}$		$\underline{10\ e^{-}}$
0		1^{+}

Magnesium atom		Magnesium ion
$\dot{Mg}^{\bullet}$ $- 2e^{-}$	$\longrightarrow$	Mg^{2+}
2-8-2		2-8 (=Ne)
$12\ p^{+}$		$12\ p^{+}$
$\underline{12\ e-}$		$\underline{10\ e^{-}}$
0		2^{+}

Covalent Bonds: are formed when two atoms share one or more electron pairs.

When two atoms share one pair of electrons, the result is a single bond.

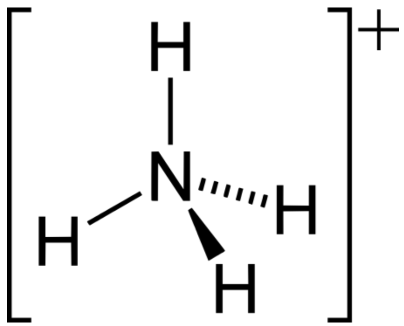


Two shared pairs of electrons is a double bond; three is a triple bond.

Rather than individual atoms that have lost or gained electrons, many ions are groups of atoms bonded together **covalently** and having **a net charge**.

Example ammonium ion, NH_4^+ consisting of 4 hydrogen atoms covalently bonded to a single nitrogen (N) atom and having a net electrical charge of +1 for the whole cation.

The number of valence electrons is the number of electrons in the outer shell, that the atom uses for bonding. Nitrogen has 5 electrons in its n=2 (outer) shell.



Ammonium ion NH_4^+

Mixtures and Pure Substances

A pure substance consisting of only one compound or of only one form of an element

The compound water with nothing dissolved in it is a pure substance.

Air is a mixture of elemental gases and compounds, predominantly nitrogen, oxygen, argon, carbon dioxide, and water vapor.

Drinking water is a mixture containing calcium ion ($Ca^{2+,}$ hydrogen carbonate ion, (bicarbonate, HCO_3^-), nitrogen gas, carbon dioxide gas, and other substances dissolved in the water.

Mixtures can be separated into their constituent pure substances by physical processes. Liquefied air is distilled to isolate pure oxygen.

A heterogeneous mixture is one that is not uniform throughout and possesses readily distinguishable constituents.

Homogeneous mixtures are uniform throughout; to observe different constituents would require going down to molecular levels.

Homogeneous mixtures are also called solutions, a term that is usually applied to mixtures composed of gases, solids, and other liquids dissolved in a liquid.

QUANTITY OF MATTER: THE MOLE

One of the most fundamental characteristics of a specific body of matter is the quantity of it.

In discussing the quantitative chemical characteristics of matter it is essential to have a way of expressing quantity in a way that is proportional to the number of individual entities of the substance- that is, atoms, molecules, or ions.

A mole is defined as the quantity of substance that contains the same number of specified entities as there are atoms of C in exactly 0.012 kg (12 g) of carbon-12.

Example:

A mole of argon, which always exists as individual Ar atoms: The atomic mass of Ar is 40.0. Therefore exactly one mole of Ar is 40.0 grams of argon.

A mole of molecular elemental hydrogen, H_2: The atomic mass of H is 1.0, the molecular mass of H_2 is, therefore, 2.0, and a mole of H_2 is 2.0 g of H_2.

A mole of methane, CH_4: The atomic mass of H is 1.0 and that of C is 12.0, so the molecular mass of CH_4 is 16.0. Therefore a mole of methane has a mass of 16.0 g.

The Mole and Avogadro's Number

Avogadro's number (6.02×10^{23}) the number of specified entities in a mole of substance. The "specified entities" may consist of atoms or molecules or they may be groups of ions making up the smallest possible unit of an ionic compound.

Example:

Substance	Formula unit	Formula mass	Mass of 1 mole	Numbers and kinds of individual entities in 1 mole
Helium	He atom	4.003[a]	4.003 g	6.02×10^{23} (Avogadro's number) of He atoms
Fluorine gas	F_2 molecule	38.00[b]	38.00 g	6.02×10^{23} F_2 molecules $2 \times 6.02 \times 10^{23}$ F atoms
Methane	CH_4 molecules	16.04[b]	16.04 g	6.02×10^{23} CH_4 molecules 6.02×10^{23} C atoms $4 \times 6.02 \times 10^{23}$ H atoms
Sodium oxide	Na_2O	62.00[c]	62.00 g	6.02×10^{23} Na_2O formula units $2 \times 6.02 \times 10^{23}$ Na^+ ions 6.02×10^{23} O^{2-} atoms

[a] Specifically, atomic mass
[b] Specifically, molecular mass

PHYSICAL PROPERTIES OF MATTER

Physical properties of matter are those that can be measured without altering the chemical composition of the matter.

Physical properties are important in describing and identifying particular kinds matter

Physical properties are very useful in assessing the hazards and predicting the fates of environmental pollutants.

An organic substance that readily forms a vapor (volatile organic compound) will tend to enter the atmosphere or to pose an inhalation hazard.

A brightly colored water soluble pollutant may cause deterioration of water quality by adding water “color.” Much of the health hazard of asbestos is its tendency to form extremely small-diameter fibers that are readily carried far into the lungs and that can puncture individual cells.

Several important physical properties of matter are density, color, solubility. And Thermal Properties.

Density (d) is defined as mass per unit volume and is expressed by the formula

d = mass /volume

Density may be expressed in any units of mass or volume. The densities of liquids and solids are normally given in units of grams per cubic centimeter (g/cm^3, the same as grams per milliliter, g/ml

The volume of a given mass of substance varies with temperature, so the density is a function of temperature.

This variation is relatively small for solids, greater for liquids, and very high for gases.

The temperature-dependent variation of density is an important property of water and results in stratification of bodies of water, which greatly affects the environmental chemistry that occurs in lakes and reservoirs.

The combined temperature/pressure relationship for the density of air causes air to become stratified into layers, particularly the troposphere near the surface and the stratosphere from about 13 to 50 kilometers altitude.

Specific Gravity

Often densities are expressed by means of specific gravity defined as the ratio of the density of a substance to that of a standard substance.

For solids and liquids, the standard substance is usually water; for gases it is usually air.

For example, the density of ethanol (ethyl alcohol) at 20˚C is 0.7895 g/ml. The specific gravity of ethanol at 20˚C referred to water at 4˚C is given by

$$\text{Specific gravity} = \frac{\text{density of ethanol}}{\text{density of water}} = \frac{0.7895\ \text{g/mL}}{1.0000\ \text{g/mL}} = 0.7895$$

For an exact value of specific gravity, the temperatures of the substances should be specified.

In this case the notation of “specific gravity of ethanol at 20˚/4˚ C” shows that the specific gravity is the ratio of the density of ethanol at 20˚C to that of water at 4˚C.

Color is one of the more useful properties for identifying substances without doing any chemical or physical tests.

A violet vapor, for example, is characteristic of iodine.

A red/brown gas could well be bromine or nitrogen dioxide (NO_2)

A characteristic yellow/brown color in water may be indicative of organically bound iron.

The human eye responds to colors of electromagnetic radiation ranging in wavelength from about 400 nanometers (nm) to somewhat over 700 nm.

Within this wavelength range humans see light; immediately below 400 nm is ultraviolet radiation, and somewhat above 700 nm is infrared radiation.

Light with a mixture of wavelengths throughout the visible region, such as sunlight, appears white to the eye.

Light over narrower wavelength regions has the following colors: 400–450 nm, blue; 490–550 nm, green; 550–580 nm, yellow; 580–650 nm, orange, 650 nm–upper limit of visible region, red. Solutions are colored because of the light they absorb. Red, orange, and yellow solutions absorb violet and blue light; purple solutions absorb green and yellow light; and blue and green solutions absorb orange and red light. Solutions that do not absorb light are colorless (clear); solids that do not absorb light are white.

STATES OF MATTER

Solids have a definite shape and volume.

Liquids have an indefinite shape and take on the shape of the container in which they are contained.

Solids and liquids are not significantly compressible, which means that a specific quantity of a substance has a definite volume and cannot be squeezed into a significantly smaller volume.

Gases take on both the shape and volume of their containers.

A quantity of gas can be compressed to a very small volume and will expand to occupy the volume of any container into which it is introduced.

Changes in matter from one phase to another are very important in the environment.

Example, water vapor changing from the gas phase to liquid results in cloud formation or precipitation.

Water is desalinated by producing water vapor from sea water, leaving the solid salt behind, and recondensing the pure water vapor as a salt-free liquid.

State of Matter

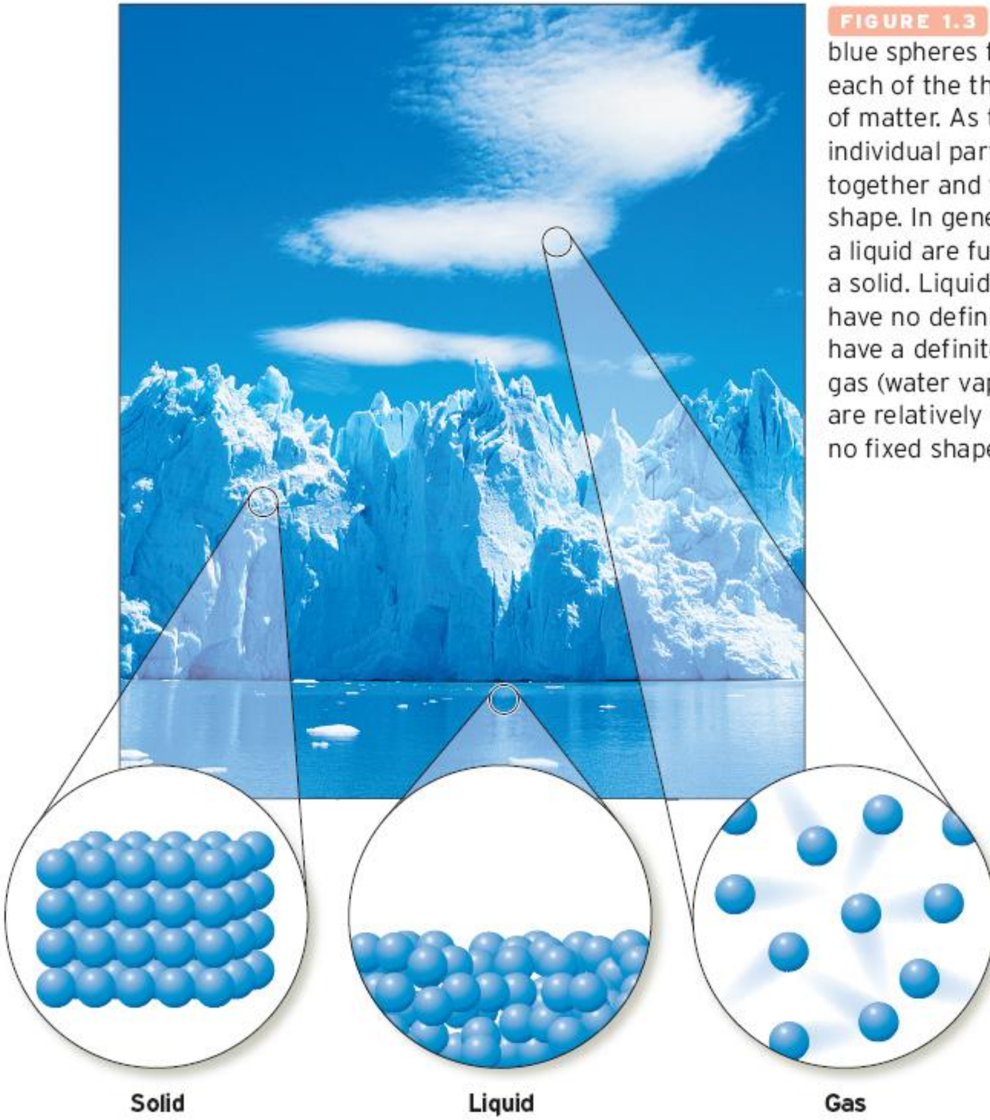


FIGURE 1.3 Water (shown as blue spheres for convenience) in each of the three principal states of matter. As the solid (ice), individual particles are close together and form a definite shape. In general, the particles of a liquid are further apart than as a solid. Liquids such as water have no definite shape but do have a definite volume. As the gas (water vapor), the particles are relatively far apart and have no fixed shape or volume.

GASES

Atmosphere is composed of a mixture of gases, the most abundant of which are nitrogen, oxygen, argon, carbon dioxide, and water vapor.

Gases are crucial to the well-being of life on Earth. Deprived of oxygen, an animal loses consciousness and dies within a short time.

Nitrogen extracted from the air is converted to chemically bound forms that are crucial to plant growth. Plants require carbon dioxide and water vapor condenses to produce rain.

Physically, gases are the “loosest” form of matter. A quantity of gas has neither a definite shape nor a definite volume so that it takes on the shape and volume of the container in which it is held.

The reason for this behavior is that gas molecules move independently and at random, bouncing off each other as they do so.

They move very rapidly; at 0˚C the average molecule of hydrogen gas moves at 3600 miles per hour.

Gas molecules colliding with container walls exert pressure.

The Gas Laws

Boyle's Law

a fixed quantity of gas is inversely proportional to the pressure of the gas. Boyle's law may be stated mathematically as $V = \text{(a constant)}\ 1/P$

Charles' Law

The volume of a fixed quantity of gas is directly proportional to the absolute temperature (˚C + 273) at constant pressure . $V = \text{(a constant)} \times T$ (where T is the temperature in Kelvin).

Example: calculate the volume of a gas with an initial volume of 10.0 L temperature changes from -11.0˚C to 95.0˚C at constant pressure.

$T_1 = 273 + (-11) = 262\ K$

$T_2 = 273 + 95 = 368\ K$

$V_2 = V_1 \times (T_2 / T_1)$

$V_2 = 10.0 \times (368 / 262) = 14\ L$

Avogadro's Law

At constant temperature and pressure the volume of a gas is directly proportional to the number of molecules of gas, commonly expressed as moles .

$V = \text{(a constant)} \times n$ where n is the number of moles of gas.

The General Gas Law

The three gas laws just defined may be combined into a general gas law stating that the volume of a quantity of ideal gas is proportional to the number of moles of gas and its absolute temperature and inversely proportional to its pressure .

$V = (\text{a constant}) \times nT/P$

Designating the proportionality constant as the ideal gas constant, R, yields the ideal gas equation: **$V = RnT/P$ or $PV = RnT$**

The units of R depend upon the way in which the ideal gas equation is used. For calculations involving volume in liters and pressures in atmospheres, the value of **R is 0.0821** L-atm/deg-mol.

The ideal gas equation shows that at a chosen temperature and pressure a mole of any gas should occupy the same volume.

A temperature of 0˚C (273.15 K) and 1 atm pressure have been chosen as standard temperature and pressure (STP).

At STP the volume of 1 mole of ideal gas is 22.4 L . This volume is called the molar volume of a gas.

Example: Calculate the volume of 0.333 moles of gas at 300 K under a pressure of 0.950 atm:

$$V = \frac{nRT}{P} = \frac{0.333 \text{ mol} \times 0.0821 \text{ L atm/K mol} \times 300 \text{ K}}{0.950 \text{ atm}} = 8.63 \text{ L}$$

Example: calculate the temperature of 2.50 mol of gas that occupies a volume of 52.6 L under a pressure of 1.15 atm:

$$T = \frac{PV}{nR} = \frac{1.15 \text{ atm} \times 52.6 \text{ L}}{2.50 \text{ mol} \times 0.0821 \text{ L atm/K mol}} = 295 \text{ K}$$

Where subscripts 1 and 2 denote parameters before and after any change, respectively, the following relationship can be derived from the fact that R is a constant in the ideal gas law equation:

$$\boxed{\frac{P_2V_2}{n_2T_2} = \frac{P_1V_1}{n_1T_1}}$$

LIQUIDS AND SOLUTIONS

Molecules of liquids are close enough together that they can be regarded as "touching" and are strongly attracted to each other.

Like those of gases, however, the molecules of liquids move freely relative to each other. These characteristics give rise to several significant properties of liquids.

Liquids do not expand to fill their containers as do gases. Essentially, therefore, a given quantity of liquid occupies a fixed volume.

Because of the free movement of molecules relative to each other in a liquid, it takes on the shape of that portion of the container that it occupies.

Evaporation and Condensation of Liquids

The molecules of liquid water are in constant motion. They have a distribution of energies such that there are more higher energy molecules at higher temperatures.

They escape from the mass of liquid and enter the gas phase. This phenomenon is called evaporation.

Molecules of liquid in the gas phase, such as water vapor molecules, may come together or strike the surface of the liquid and reenter the liquid phase. This process is called condensation.

Vapor Pressure

In a confined area above a liquid, equilibrium is established between the evaporation and condensation of molecules from the liquid.

For a given temperature, this results in a steady-state level of the vapor, which can be described as a pressure.

Such a pressure is called the vapor pressure .

Vapor pressure is very important in determining the fates and effects of substances, including hazardous substances, in the environment.

Loss of a liquid by evaporation to the atmosphere increases with increasing vapor pressure and is an important mechanism by which volatile pollutants enter the atmosphere.

High vapor pressure of a flammable liquid can result in the formation of explosive mixtures of vapor above a liquid, such as in a storage tank.

Solutions

Solutions are homogeneous mixtures. Solutions are regarded as liquids in which quantities of gases, solids, or other liquids are dispersed as individual ions or molecules.

Rainwater, for example, is a solution containing molecules of N_2, O_2 and CO_2 from air dispersed among the water molecules.

Rainwater from polluted air may contain harmful substances such as sulfuric acid (H_2SO_4) in small quantities among the water molecules.

The predominant liquid constituent of a solution is called the solvent.

A substance dispersed in it is said to be dissolved in the solvent and is called the solute.

Solids

The solid state is the most organized form of matter in that the atoms, molecules, and ions in it are in essentially fixed relative positions and are highly attracted to each other. Therefore, solids have a definite shape, maintain a constant volume, are virtually non-compressible under pressure, and expand and contract only slightly with changes in temperature.

Because of the strong attraction of the atoms, molecules, and ions of solids for each other, solids do not enter the vapor phase very readily at all; the phenomenon by which this happens to a limited extent is called sublimation

Some solids (quartz, sodium chloride) have very well defined geometric shapes and form characteristic crystals.

Other solids have indefinite shapes because their constituents are arranged at random; glass is such a solid. These are amorphous solids.

THERMAL PROPERTIES

The behavior of a substance when heated or cooled defines several important physical properties of it, including the temperature at which it melts or vaporizes.

Melting Point: The melting point of a pure substance is the temperature at which the substance changes from a solid to a liquid.

The melting point of a pure substance is indicative of the identity of the substance.

The melting behavior—whether melting occurs at a single temperature or over a temperature range—is a measure of substance purity.

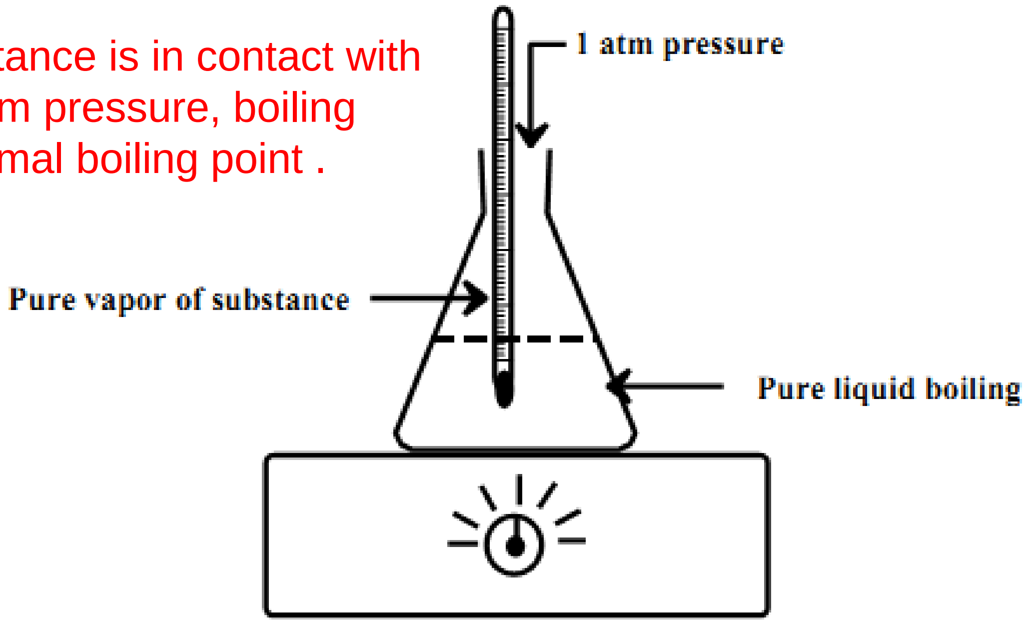
Boiling Point

Boiling occurs when a liquid is heated to a temperature such that bubbles of vapor of the substance are evolved.

The boiling temperature depends upon pressure, and reduced pressures can be used to cause liquids to boil at lower temperatures.

The boiling point is useful for identifying liquids. The extent to which the boiling temperature is constant as the liquid is converted to vapor is a measure of the purity of the liquid.

When the surface of a pure liquid substance is in contact with the pure vapor of the substance at 1 atm pressure, boiling occurs at a temperature called the normal boiling point .



Specific Heat: The amount of heat energy required to raise the temperature of a unit mass of a solid or liquid substance by a degree of temperature varies with the substance.

The very high amount of energy required to increase the temperature of water has some important environmental implications because it stabilizes the temperatures of bodies of water and of geographic areas close to bodies of water.

In contrast to water, hydrocarbon liquids such as those in gasoline require relatively little heat for warming because the molecules interact much less with each other than do those of water.

$$\boxed{\text{Specific heat} = \frac{\text{heat energy absorbed, J}}{(\text{mass, g})(\text{increase in temperature, }^\circ\text{C})}}$$

where the heat energy is in units of joule, J. The most important value of specific heat is that of water, 4.18 J/g˚C.

The amount of heat, q, required to raise the mass, m, of a particular substance over a temperature range of ΔT is $q = (\text{specific heat}) \times m \times \Delta T$

Example: consider the amount of heat required to raise the temperature of 11.6 g of liquid water from 14.3˚C to 21.8˚C:

$$q = 4.18 \frac{\text{J}}{\text{g}^\circ\text{C}} \times 11.6\ \text{g} \times (21.8^\circ\text{C} - 14.3^\circ\text{C}) = 364\ \text{J}$$

Heat of vaporization: is the quantity of heat taken up in converting a unit mass of liquid entirely to vapor at a constant temperature. The heat of vaporization of water is 2,260 J/g (2.26 kJ/g) for water boiling at 100˚C at 1 atm pressure.

When water vapor condenses, similar enormous amounts of heat energy, called heat of condensation, are released.

This occurs when water vapor forms precipitation in storm clouds and is the driving force behind the tremendous releases of heat that occur in thunderstorms and hurricanes.

The heat of vaporization of water can be used to calculate the heat required to evaporate a quantity of water.

Example: The heat, q, required to evaporate 2.50 g of liquid water is

q = (heat of vaporization)(mass of water) = 2.26 kJ/g x 2.50 g = 5.65 kJ

and that released when 5.00 g of water vapor condenses is

q = -(heat of vaporization)(mass of water)

= -2.26 kJ/g x 5.00 g = -11.3 kJ. The latter value is negative to express the fact that heat is released.

Heat of fusion is the quantity of heat taken up in converting a unit mass of solid entirely to liquid at a constant temperature. The heat of fusion of water is 330 J/g for ice melting at 0˚C.

Example: the heat, q, required to melt 2.50 g of ice is

$$q = \text{(heat of fusion)(mass of water)} = 330 \text{ J/g} \times 2.50 \text{ g} = 825 \text{ J.}$$

and that released when 5.00 g of liquid water freezes is

$$q = - \text{(heat of fusion)(mass of water)} = -330 \text{ J/g} \times 5.00 \text{ g} = -1{,}650 \text{ J}$$

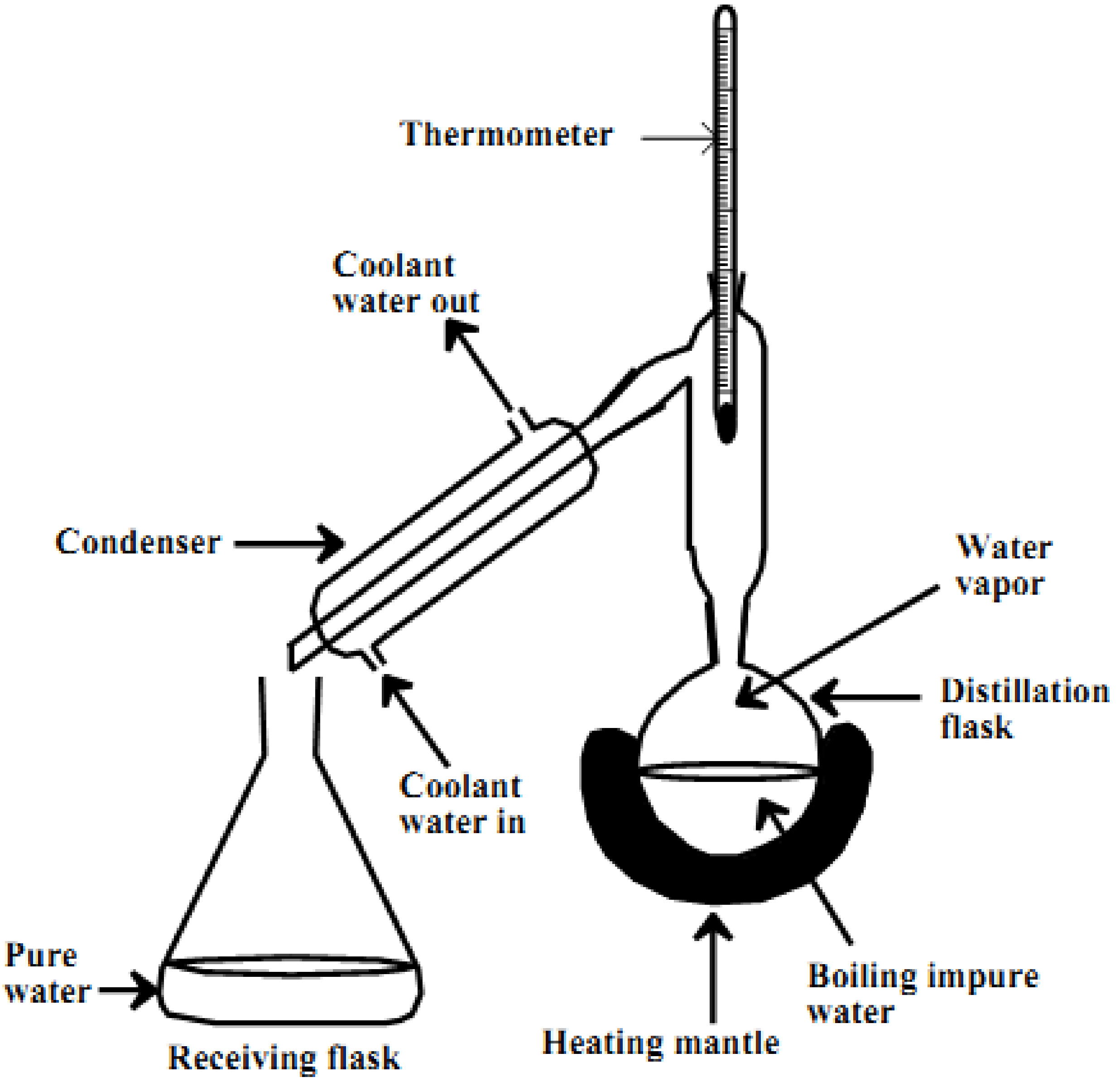
SEPARATION AND CHARACTERIZATION OF MATTER

A very important aspect of the understanding of matter is the separation of mixtures of matter into their constituent pure substances.

Distillation consists of evaporating a liquid by heating and cooling the vapor so that it condenses back to the liquid in a different container.

Less-volatile constituents such as solids are left behind in the distillation flask; more-volatile impurities such as volatile organic compound pollutants in water will distill off first and can be placed in different containers before the major liquid constituent is collected.

Separation by Distillation



The seawater to be purified is placed in a round-bottom distillation flask heated with an electrically powered heating mantle.

The seawater is boiled and pure water vapor in the gaseous state flows into the condenser where it is cooled and condensed back to a purified salt-free product in the receiving flask.

When most of the seawater has been evaporated, the distillation flask and its contents are cooled and part of the sodium chloride separates out as crystals that can be removed.

Sophisticated versions of this distillation process are used in some arid regions of the world to produce potable (drinking) water from seawater.

Distillation is widely used in the petroleum and chemical industries.

A sophisticated distillation apparatus is used to separate the numerous organic components of crude oil, including petroleum ether, gasoline, kerosene, and jet fuel fractions.

This requires fractional distillation in which part of the vapor recondenses in a fractionating column that would be mounted vertically on top of the distillation apparatus.

Molecular Separation: often based upon membrane processes in which dissolved contaminants or solvent pass through a size-selective membrane under pressure.

Reverse osmosis is the most widely used of the membrane techniques. It operates by virtue of a membrane that is selectively permeable to water and excludes ionic solutes.

Reverse osmosis uses high pressures to force permeate through the membrane, producing a concentrate containing high levels of dissolved salts.

Homework

1. Calculate the molecular masses of
(a) C_2H_2,
(b) N_2H_4,
(c) Na_2O,
(d) O_3 (ozone),
(e) PH_3,
(f) CO_2,
(g) C_3H_9O

2. Calculate the density of each of the following substances, for which the mass of a specified volume is given:
(a) 93.6 g occupying 15 mL
(b) 0.992 g occupying 1,005 mL
(c) 13.7 g occupying 11.4 mL
(d) 16.8 g occupying 3.19 mL

3. Calculate the molar concentration of solute in each of the following solutions:
(a) Exactly 7.59 g of NH_3 dissolved in 1.500 L of solution.
(b) Exactly 0.291 g of H_2SO_4 dissolved in 8.65 L of solution.
(c) Exactly 85.2 g of NaOH dissolved in 5.00 L of solution.
(d) Exactly 31.25 g of NaCl dissolved in 2.00 L of solution.
(e) Exactly 6.00 mg of atmospheric N_2 dissolved in water in a total solution volume of 2.00 L.

3. Using the appropriate gas laws, calculate the quantities denoted by the blanks in the table below:

Conditions before				Conditions after			
n_1, mol	V_1, L	P_1, atm	T_1, K	n_2, mol	V_2, L	P_2, atm	T_2, K
1.25	13.2	1.27	298	1.25	(a) _____	0.987	298
1.25	13.2	1.27	298	1.25	(b) _____	1.27	407
1.25	13.2	1.27	298	4.00	(c) _____	1.27	298
1.25	13.2	1.27	298	1.36	(d) _____	1.06	372
1.25	13.2	1.27	298	1.25	30.5	(e) _____	298

4. Calculate the amount of heat required to raise the temperature of 7.25 g of liquid water from 22.7°C to 29.2°C.

5. Calculate the heat (q)
(a) required to evaporate 5.20 g of liquid water
(b) released when 6.50 g of water vapor condenses.