I. Chemistry Review

A. ELEMENTS (CHAPTER 2)

Element – a substance that cannot be separated into simpler substances by physical or chemical means.

106 named chemicals (92 naturally occurring) 25 are necessary for life.
Four make up 96% of all living matter: **COHN**
Other 4% {Ca, P, K, S, Na, Cl, Mg, and trace elements} Trace elements – living organisms need these but in smaller amounts: B, Cr, Co, Cu, F, I, Fe, Mn, Mo, Se, Si, Sn, V and Zn.

Atom - smallest particle of an element that has the characteristics of that element.
Two or more atoms can combine chemically and form a molecule.

Compound – any pure substance that contains two or more different atoms.

Atom = H Molecule = H_2 Compound = H_2O

B. FORMS OF MATTER

Solid – has definite shape and definite volume
Liquid – has no definite shape but has definite volume
Gas – has no definite shape and no definite volume

C. SUBATOMIC PARTICLES

Atoms have a dense nucleus (protons & neutrons) with a very low density electron cloud surrounding it.

Neutron (n^o) - neutral subatomic particle in the nucleus of an atom that has a mass of 1 amu or 1 dalton.

amu – Atomic mass unit. A unit used to measure the mass of very small particles such as atoms, protons, and neutrons.

Proton ($p+$) - positively charged subatomic particle in the nucleus that has a mass of 1 amu.

Electron ($e-$) - negatively charged and has a small mass (1/2000 of a proton). This small mass is considered to be negligible (0) in atomic mass calculations.

Strong nuclear forces hold the protons and neutrons together, while the electrons are attracted to the positive charge of the protons. Protons and neutrons can be broken down into small particles called quarks.

Atomic number - the number of protons in an atom; distinguishes atoms of different elements.

Atomic mass – mass of an atom.

Mass number – protons + neutrons in an atom of a given isotope.

Isotope - atoms of the same element having a different number of neutrons and therefore have different atomic masses.

Hydrogen: 1p, 1e- 1 amu	Deuterium: 1p, 1n, 1e- 2 amu	Tritium: 1p, 2n, 1e- 3 amu

Some combinations of protons and neutrons are stable, but other combinations are internally unstable and break down spontaneously. When this happens, the atoms release various subatomic particles and radiation. These isotopes are called radioactive isotopes.

D. ELECTRON ORBITALS

- The first energy level (closest to nucleus) can hold up to 2 electrons (1s)
- The second energy level can hold up to eight electrons (2s, 2p)
- The third energy level can hold up to eight electrons (3s, 3p)

There are more than three energy levels, but biologists are concerned with 18 total electrons. Atoms are most stable when their outer energy level is filled with electrons. Of the three subatomic particles, only the electrons are directly involved in the chemical reactions between atoms.

Not every electron has the same amount of energy (the ability to do work). Potential energy, the amount of energy that matter stores, is due to the position or location of the matter. Electrons have potential energy in relation to the nucleus.

To move to a shell farther out from the nucleus, the electron must absorb energy (ex: Light energy can excite an electron to a higher energy level). To move to a shell closer in, an electron must lose energy, which is usually released to the environment in the form of heat.

E. HOW ATOMS FILL THEIR OUTER SHELL

An atom with its outer shell filled with electrons is a stable atom. Outer electrons are **valence electrons**. Outermost electron shell is the **valence shell**. The valence of H is 1; O, 2; nitrogen, 3; and carbon, 4. An atom with a completed valence shell is unreactive.

The **noble gases**: He, Ne, Ar have full valence shells and are called **inert elements.**

One slightly more complicated case is phosphorus (P). It can have a valence of 3 as we predict. In biologically important molecules, however, it generally has a valence of 5, forming three single bonds and one double bond.

Atoms react with other atoms chemically by filling their outer shells in one of three ways:

1. Ionic Bonds
    a. Gain electrons from another atom
    b. Lose electrons from its outer shell to another atom
2. Covalent Bonds
    a. Share one or more pairs of electrons with another atom

F. BONDS

1. Ionic bonds and Ions

When two atoms are so unequal in their attraction for valence electrons that the more electronegative atom strips an electron completely away from its partner.

Electronegativity is the attraction of electrons to an atom.

Sodium has 11 electrons: $1s^2, 2s^2, 2p^6, 3s^1$
(Na needs to gain 7 more electrons or lose 1 electron.)

Chlorine has 17 electrons: $1s^2, 2s^2, 2p^6, 3s^2, 3p^5$
(Cl has to lose seven electrons or gain one electron.)

Na donates one electron to Cl, these two atoms combine to form a compound, sodium chloride salt. Ionic compounds are called **salts**. An **ion** is any charged atom. Sodium donates an electron, which is negatively charged and becomes a positively charged ion. The chlorine receives an electron and becomes a negatively charged ion. The two ions are Na^+ and Cl^-. When the two atoms give and receive electrons, they form ions and ionic bonds.

Cation – positively charged ion
Anion – negatively charged ion

2. Covalent Bonds

Electrons between atoms are shared. Covalent bonds occur when the electronegativities of the atoms are similar.

A. Nonpolar covalent bonds form when the electrons are <u>shared equally</u>, such as in O_2, the electronegativities are identical and both atoms pull equally on the electrons.
B. Polar covalent bonds form when electrons are <u>shared unequally</u>. The atom with the greater electronegativity will be slightly negative due to the fact that a negative electron spends more time around its nucleus. The other atom has a slightly positive charge. In a molecule of water (H_2O), for example, electrons are shared between the oxygen atom and each hydrogen atom. Oxygen, with a greater electronegativity exerts a stronger pull on the shared electrons than does the hydrogen atom. This unequal distribution of electrons creates a negative pole near the oxygen and positive poles near each hydrogen atom.

- Single covalent, double covalent, and triple covalent bonds form when, two, four, and six electrons are shared, respectively.

C. Hydrogen Bonds are weak bonds between molecules. They form when a positively charged hydrogen atom in one covalently bonded molecule is attracted to a negatively charged area of another covalently bonded molecule. Hydrogen bonds are 20 times weaker than covalent bonds.

nonpolar covalent bonds	polar covalent bonds	ionic bonds
e- shared equally	*e- shared unequally*	*e- transferred*

⟵⟶

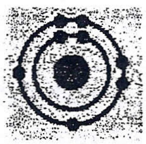
increasing difference of electronegativity between bonding atoms ⟶

Note: Covalent bonds form when electronegativities of the atoms are similar. Ionic bonds occur when the electronegativity of the atoms are very <u>different</u> and one atom has a stronger pull on the electron.

Structural Formula O = O

Molecular formula O_2

Electron configuration for the element O



Chemical reactions make and break chemical bonds. Matter is conserved in a chemical reaction: reactions cannot create or destroy matter but can only rearrange it.

Photosynthesis:

$$6CO_2 + 6H_2O \longrightarrow C_6H_{12}O_6 + 6\,O_2$$

reactants yields products

⇔ opposite arrow heads indicate that the reaction is reversible. Most chemical reactions are reversible.

Chemical equilibrium – equilibrium established when the rate of the forward reaction equals the rate of the reverse reaction.

A molecule's biological function is related to its shape. Shape is determined by the positions of it's atoms' valence orbitals. Shape is often the basis of the recognition of one biological molecule by another.

II. CHEMISTRY OF WATER (CHAPTER 3)

1. Water Properties

A. Powerful Solvent

Water is able to dissolve anything polar due to its polarity. Water separates ionic substances. Covalent compounds dissolve in water are called hydrophilic. Nonpolar substances that do not dissolve in water are called hydrophobic.

B. Capillary Action

1. Adhesion: the attraction between water and other substances
2. Cohesion: the attraction of water molecules to other water molecules

These two properties allow capillary action. The meniscus, in a column of water, is formed because gravity pulls down on the water molecules in the center while water molecules at the sides of the container "climb".

C. High Surface Tension

Water is attracted to itself, and this attraction, due to **hydrogen bonds**, is stronger than the attraction to the air above it. Strong cohesion between water molecules creates a surface that is firm enough to allow many insect to walk upon without sinking.

D. High Specific Heat

The amount of heat that must be absorbed or lost for 1 g of that substance to change its temperature by 1°C. Water changes temperature very slowly, therefore, the temperatures of large bodies of water are very stable in response to the temperature changes of the surrounding air. Water heats up as the hydrogen atoms vibrate (molecular kinetic energy-energy of molecular motion). 1. Water covers ¾ of the earth, keeps temp fluctuation within a range suitable for life 2. Coastal areas have milder climates than inland 3. Marine environments are relatively stable

E. High Boiling Point

A great deal of energy must be present in order to break the hydrogen bonds to change water from a liquid to a gas.

F. Good Evaporative Coolant

When humans sweat, water absorbs heat from the body. When the water turns into water vapor, it takes that energy (heat) with it.

G. High Freezing Point & Lower Density As A Solid Than A Liquid

Water's maximum density is 4°C, while freezing is 0°C. This is why ice floats, this allows for aeration of still ponds in Spring and Fall and the reason that ponds don't freeze from the bottom up. Expands when freezes. Lakes don't freeze because ice floats. As water freezes it releases heat to the water below and insulates it, Water freezes, hydrogen bonds form releasing heat. As ice melts, hydrogen bonds break absorbing heat.

H. Dissociation and pH Scale

Many substances come apart (dissociate) in water. Some dissociate completely, while others dissociate only partly. In a solution, some molecules are intact while others are ionized (gain or lose e-). Water dissociates into H^+ and OH^- equally (hydrogen ion and hydroxide ion).

A mole (mol) is equal in number to the molecular weight of a substance, but upscaled from Daltons to units of grams. Sucrose (C12H22O11) molecular wt = 342g

Molecular weight = sum of the weight of all atoms in a molecule (expressed in Daltons)

C = 12 dal	12 dal x 12 = 144
H = 1 dal	1 dal x 22 = 22
O = 16 dal	16 dal x 11 = 176
	342 g

1. Acids

   Substances that yield H^+ when they dissociate in water are called acids. Acids add H^+ to the solution, increasing the H^+ concentration.

   $$HCl \longrightarrow H^+ \text{ and } Cl^-$$

Molarity = number of moles of solute per liter of solution

2. Bases

   Substances that yield OH^- when they dissociate in water are called bases (e.g. $NaOH \longrightarrow Na^+$ and OH^-). Bases also accept H^+. Bases reduce the amount of H^+ in a solution.

   $$OH^- + H^+ \longrightarrow H_2O \quad \text{or} \quad NH_3 + H^+ \longrightarrow NH_4^+$$

3. Salts

   A salt is a substance in which the H^+ of an acid is replaced by another positively charged ion. $HCl + Na \longrightarrow NaCl$ and H^+

4. pH The acidity or alkalinity (base) is known as pH.

   p18 Schwaum's logs

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

   If pH = 6 then the concentration of H^+ per liter is 10^{-6} in a solution.

   PH Scale

   0 1 2 3 4 5 6 $\boxed{7}$ 8 9 10 11 12 13 14

   acidic ⟵ neutral ⟶ basic

   A pH of 5 is 10 times more acidic than a pH of 6. The more H^+, the ↑ the $[H^+]$ and the more acidic. Acid precipitation is rain, snow, or fog with a pH below 5.6.

Overhead 3.9: The pH of some aqueous solutions

5. Buffers

   Buffers are substances that takes up and releases H^+ or OH^- to prevent swings in pH. An important buffer is H2CO3 (carbonic acid). H2CO3 dissociates to H^+ and $HCO3^-$. The H^+ is a base acceptor, and the $HCO3^-$ is an acid acceptor. A buffer consists of an acid-base pair that combines reversibly with hydrogen ions.

   carbonic acid — bicarbonate — hydrogen ion

   pH falls ⟶

   $$H_2CO_3 \Leftrightarrow HCO_3^- + H^+$$

   pH rises ⟵

   Bicarbonate acts a base to accept excess H^+ ions when the pH starts to fall; the reaction moves to the left

   When pH rises, H^+ ions are donated by carbonic acid, the reaction shifts to the right

Periodic Table of the Elements

Period	I	II										III	IV	V	VI	VII	Noble gases VIII
1	1.00794 **H** 1 1																4.00260 **He** 2
2	6.941 **Li** 3 2 1	9.01218 **Be** 4 2 2										10.81 **B** 5 2 3	12.011 **C** 6 2 4	14.0067 **N** 7 2 5	15.9994 **O** 8 2 6	18.998403 **F** 9 2 7	20.179 **Ne** 10
3	22.98977 **Na** 11 2 8 1	24.305 **Mg** 12 2 8 2										26.98154 **Al** 13 2 8 3	28.0855 **Si** 14 2 8 4	30.97376 **P** 15 2 8 5	32.06 **S** 16 2 8 6	35.453 **Cl** 17 2 8 7	39.948 **Ar** 18
4	39.0983 **K** 19 2 8 8 1	40.08 **Ca** 20 2 8 8 2	44.9559 **Sc** 21 2 8 9 2	47.88 **Ti** 22 2 8 10 2	50.9415 **V** 23 2 8 11 2	51.996 **Cr** 24 2 8 13 1	54.9380 **Mn** 25 2 8 13 2	55.847 **Fe** 26 2 8 14 2	58.9332 **Co** 27 2 8 15 2	58.69 **Ni** 28 2 8 16 2	63.546 **Cu** 29 2 8 18 1	65.38 **Zn** 30 2 8 18 2	69.72 **Ga** 31 2 8 18 3	72.59 **Ge** 32 2 8 18 4	74.9216 **As** 33 2 8 18 5	78.96 **Se** 34 2 8 18 6	79.904 **Br** 35 2 8 18 7	83.80 **Kr** 36
5	85.4678 **Rb** 37 2 8 18 8 1	87.62 **Sr** 38 2 8 18 8 2	88.9059 **Y** 39 2 8 18 9 2	91.22 **Zr** 40 2 8 18 10 2	92.9064 **Nb** 41 2 8 18 12 1	95.94 **Mo** 42 2 8 18 13 1	[98] **Tc** 43 2 8 18 13 2	101.07 **Ru** 44 2 8 18 15 1	102.9055 **Rh** 45 2 8 18 16 1	106.42 **Pd** 46 2 8 18 18 0	107.8682 **Ag** 47 2 8 18 18 1	112.41 **Cd** 48 2 8 18 18 2	114.82 **In** 49 2 8 18 18 3	118.69 **Sn** 50 2 8 18 18 4	121.75 **Sb** 51 2 8 18 18 5	127.60 **Te** 52 2 8 18 18 6	126.9045 **I** 53 2 8 18 18 7	131.29 **Xe** 54
6	132.9054 **Cs** 55 2 8 18 18 8 1	137.33 **Ba** 56 2 8 18 18 8 2	Lanthanide Series 174.967 **Lu** 71 2 8 18 32 9 2	178.49 **Hf** 72 2 8 18 32 10 2	180.9479 **Ta** 73 2 8 18 32 11 2	183.85 **W** 74 2 8 18 32 12 2	186.207 **Re** 75 2 8 18 32 13 2	190.2 **Os** 76 2 8 18 32 14 2	192.22 **Ir** 77 2 8 18 32 15 2	195.08 **Pt** 78 2 8 18 32 17 1	196.9665 **Au** 79 2 8 18 32 18 1	200.59 **Hg** 80 2 8 18 32 18 2	204.383 **Tl** 81 2 8 18 32 18 3	207.2 **Pb** 82 2 8 18 32 18 4	208.9804 **Bi** 83 2 8 18 32 18 5	[209] **Po** 84 2 8 18 32 18 6	[210] **At** 85 2 8 18 32 18 7	[222] **Rn** 86
7	[223] **Fr** 87 2 8 18 32 18 8 1	226.0254 **Ra** 88 2 8 18 32 18 8 2	Actinide Series [260] **Lr** 103 2 8 18 32 32 9 2	[261] **Unq** 104 2 8 18 32 32 10 2	[262] **Unp** 105 2 8 18 32 32 11 2	[263] **Unh** 106 2 8 18 32 32 12 2	[261?] **Uns** 107 2 8 18 32 32 13 2	[265] **Uno** 108 2 8 18 32 32 14 2	[266?] **Une** 109 2 8 18 32 32 15 2									

METALS — TRANSITION ELEMENTS — NONMETALS

RARE EARTH ELEMENTS

Lanthanide Series														
138.9055 **La** 57 2 8 18 18 9 2	140.12 **Ce** 58 2 8 18 20 8 2	140.9077 **Pr** 59 2 8 18 21 8 2	144.24 **Nd** 60 2 8 18 22 8 2	[145] **Pm** 61 2 8 18 23 8 2	150.36 **Sm** 62 2 8 18 24 8 2	151.96 **Eu** 63 2 8 18 25 8 2	157.25 **Gd** 64 2 8 18 25 9 2	158.9254 **Tb** 65 2 8 18 27 8 2	162.50 **Dy** 66 2 8 18 28 8 2	164.9304 **Ho** 67 2 8 18 29 8 2	167.26 **Er** 68 2 8 18 30 8 2	168.9342 **Tm** 69 2 8 18 31 8 2	173.04 **Yb** 70 2 8 18 32 8 2	

Actinide Series														
227.0278 **Ac** 89 2 8 18 32 18 9 2	232.0381 **Th** 90 2 8 18 32 18 10 2	231.0359 **Pa** 91 2 8 18 32 20 9 2	238.0289 **U** 92 2 8 18 32 21 9 2	237.0482 **Np** 93 2 8 18 32 22 9 2	[244] **Pu** 94 2 8 18 32 24 8 2	[243] **Am** 95 2 8 18 32 25 8 2	[247] **Cm** 96 2 8 18 32 25 9 2	[247] **Bk** 97 2 8 18 32 27 8 2	[251] **Cf** 98 2 8 18 32 28 8 2	[252] **Es** 99 2 8 18 32 29 8 2	[257] **Fm** 100 2 8 18 32 30 8 2	[258] **Md** 101 2 8 18 32 31 8 2	[259] **No** 102 2 8 18 32 32 8 2	

A value given in brackets denotes the mass number of the isotope of longest known half-life.