

Chapter 14

1. There are two different elements in a water molecule.
2. Atoms heavier than hydrogen were made by **nuclear fusion**.
3. If we doubled the magnifying power of the most powerful optical microscope in the world, **we would still not be able to see an atom**.
4. **Chemical compounds** are made up of about 100 distinct elements. Atoms are the smallest subdivision of matter that still retains chemical properties of a substance. A **quark** is the smallest particle that exists. **Electrons** form atoms that in turn determine chemical properties of a substance.
5. The number of **protons** makes an element distinct.
6. Brownian motion has to do with **random motions of atoms and molecules**.
7. A **quark** is the smallest particle.
8. A molecule has **structure, mass, and energy**.
9. Solid matter is mostly empty space. **Electrical forces** prevent the solids from falling through one another.
10. **Electrical forces** determine the chemical properties of an atom.
11. The air in this room had **no energy, no weight, and no mass**.
12. Assuming all the atoms exhaled by Julius Caesar in his last dying breath are still in the atmosphere, then we probably breathe one of those atoms with each **single breath**.
13. Nuclei of atoms that make up a newborn baby were made in **ancient stars**.
14. The reason a granite block is mostly empty space is because the atoms in the granite are **mostly empty space themselves**.
15. **Neutrons** are electrically neutral. **Protons** are electrically positive. Electrons are electrically negative.
16. In an electrically neutral atom, the number of protons in the nucleus is balanced by an equal number of **orbital electrons**.
17. **Lead** has the greatest number of protons in its nucleus (compared to: Silver, gold, & mercury.)

18. The weight of matter comes mostly from its **protons and neutrons**.
19. The volume of matter comes mostly from its **electrons**.
20. The chemical properties of matter come mostly from its **electrons**,
21. Compared to the mass of a hydrogen atom, the mass of an oxygen atom is **16 times greater**.
22. The total number of protons and neutrons in a water molecule is 18.
23. When two protons are removed from an oxygen nucleus, the result is **carbon**.
24. To change mercury into gold, pair protons must be added to the gold nucleus.
25. If pair of helium nuclei is fused together, the result is **beryllium**.
26. An electrically neutral iron atom contains 26 protons and **26 electrons**.
27. **A one gram sample of Carbon-12** has more atoms than a one gram sample of carbon-13.
28. The most likely set of mass numbers for the two bromine isotopes are Br-79, **Br-81**.
29. **Uranium** has the most mass (Compared to iron, hydrogen, & lead.)
30. Strontium is dangerous to humans because it tends to accumulate in the calcium-dependent bone marrow tissues. **Sr and Ca are both metals**.
31. The atomic masses listed in the periodic table are not whole numbers because **the atomic masses are average atomic masses**.

Chapter 15

1. A hydrogen atom, with only one electron, can have many spectral lines because **one electron can be boosted to many different energy levels**.
2. Suppose a certain atom possesses only four distinct energy levels. Assuming that all transitions between levels are possible, the atom will exhibit **four** spectral lines.
3. An electron de-excites from the fourth quantum level to the third and then directly to the first. Two frequencies of light are emitted. How do their combined energies compare to the energy of the single frequency that would be emitted by de-excitation from the fourth level directly to the first level? Answer: The combined

energies of the two frequencies emitted by the one electron are **equal to the energy of the single frequency.**

4. **Blue** is a higher frequency than red, and therefore corresponds to higher energy level transition.
5. The wave model of electrons orbiting the nucleus counts for the fact that the electrons can have only discrete energy values because **when an electron wave is confined, it is reinforced only at particular frequencies.**
6. **A broad spectrum of all colors** would be observed if an atom appears if its electrons were not restricted to particular energy levels.
7. An electron in a 7s orbital has more energy than one in a 1s orbital because it is **farther from the nucleus.**
8. The number of elements in a period is **equal to the electron capacity of the shells.**
9. In the outermost shell of strontium (Sr, atomic # 38) there **two electrons.**
10. In the outermost shell of phosphorus (P, atomic # 15) there are five electrons.
11. In the outermost shell of sulfur (S, atomic # 16) there are two **unpaired electrons.**
12. In the outermost shell of aluminum (Al, atomic # 13) there is **one unpaired electron.**
13. The more massive an **atom becomes the greater number of shells it has.**
14. As atoms get more massive they get smaller in size because **more mass means more protons, which act to pull electrons in closer to the nucleus.**
15. The alkali metals (group 1) tend to form 1+ ions while the alkali- earth metals (group 2) tend to form 2+ ions because **the charges of these ions corresponds to the number of valence electrons that may be lost**
16. Two shells are completely filled for the chlorine ion, Cl⁻ (Cl, atomic # 17.)
17. **Cl (atomic # 17) < Br (atomic # 35) < K (atomic # 19) < Rb (atomic # 37)** is the order of increasing atomic size.
18. Osmium, (atomic # 76) has the greatest density of all elements, and, with some exceptions, the closer an element is positioned to osmium, the greater its density. Increasing density: **copper, Cu < silver, Ag < gold, Au < platinum, Pt.**

Chapter 16

1. **Gamma rays** is radiation that has no electric charge associated with it
2. The mass of an atomic nucleon is nearly **two thousand times the mass of an electron.**
3. Once an alpha particle is outside the nucleus it is **electrostatic ally repelled.**
4. When a nucleus emits a beta particle, **its atomic number changes, but its mass number remains constant.**
5. A quark is **NOT** an elementary particle, a building block of leptons or nucleons.
6. An atom with an imbalance of electrons to protons is an **ion.**
7. The atomic number of an element is the same as the number of its **protons.**
8. Deuterium and tritium are both **forms of hydrogen and isotopes of the same element.**
9. Different isotopes of an element have different numbers of **neutrons.**
10. Electric forces within an atomic nucleus tend **to push it apart.**
11. The larger the nucleus is, the greater its **instability.**
12. The half- life of an isotope is one day. At the end of two days the amount that remains is **one-quarter.**
13. The half-life of a radioactive substance is independent of **the age of the substance, the temperature of the substance, the number (if large enough) of atoms in the substance, and whether the substance exists in an elementary state or in a compound.**
14. When an element undergoes nuclear transmutation, the result is a completely different **element.**
15. When an alpha particle is ejected from a nucleus then has less **mass and charge.**
16. The fate of the world's uranium supply is to eventually become **lead.**
17. The origin of cosmic rays is the **cosmos.**
18. **Carbon-14** is a radioactive isotope.
19. There is a greater proportion of Carbon-14 in **new bones,**
20. Carbon-14 is produced in the atmosphere principally by cosmic ray bombardment,
21. Carbon dating requires that the object being tested contains **organic material.**

22. All deposits of natural uranium contain appreciable amounts **of lead**.
23. Most of the radioactivity we personally encounter comes from the **natural environment**.
24. The source of the earth's natural heat is **radioactive decay in the earth's core**.
25. The most harmful radiations are those that **damage living cells**.
26. A sample of relatively active **radioactive material is somewhat warmer than the environment**.
27. When radium ($A=88$) emits an alpha particle, the resulting nucleus has atomic number 86.
28. When a nucleus emits a positron, its atomic number **decreases by 1**.
29. An element will decay to an element with higher atomic number in the periodic table if it emits **a beta particle**.
30. Artificially induced radioactive elements generally have **short half lives**.
31. It's Impossible for a hydrogen atom to emit an alpha particle.
32. In bubble chambers, charged particles move in spirals because **of energy dissipation**.
33. An element emits 1 alpha particle, 1 positron, and 3 beta particles. Its atomic number **stays the same**.
34. **Two separate lumps** will leak more neutrons.
35. A nucleon has more mass when it is **outside the nucleus**.
36. The nucleus with the greatest mass is **plutonium** (compared to hydrogen, Iron, lead, and uranium).
37. The nucleus with the most tightly bound nucleons is **iron** (compared to hydrogen, lead, uranium, and plutonium).
38. Detonation of a fission type atomic bomb is **started by igniting a small thermonuclear bomb**.
39. The most abundant element in the universe is **hydrogen**
40. When U-239 emits a beta particle, the nucleus left behind has **94 protons**.

Homework CH.5

1. To convert from a given mass in grams to the number of moles, you would multiply by **1/molar mass**.
2. To convert from a given number atoms to the number of moles, you would multiply by **1/Avogadro's number**. (6.02×10^{23})
3. The average mass of one atom of iron is 55.85 amu. The mass of Avogadro's number of atoms is 55.85 g.
4. The weight, in grams, of one mole of hydrogen atoms (use atomic weight: H, 1.01 amu) is **1.01g**.
5. How many grams of sulfur make up 3.01 mol of sulfur (use atomic weight: S, 32.06 amu) = **96.5g**.
6. How many moles are there in one ounce (28.4g) Of pure gold (use atomic weight: Au, 197.0 amu) = **0.144 mol**.
7. Howmany atoms of sulfur are present in 155 g of sulfur (use atomic weight: S, 32.06amu) = **2.9×10^{24} atoms**.
8. How many iron atoms are present in 3.01 mol of iron = **1.81×10^{24}** .
9. The formula weight of carbon dioxide (use atomic weight; C,12.01 amu; O,16.00 amu) = **44.01 g**
10. Aspirin is the common name for acetyl salicylic acid, $C_9H_8O_4$. A tablet has 0.325 g of aspirin. There is 1.80×10^{-3} **mol**.
11. There are **3.24g** in 0.0200mol of nicotine, a yellow liquid. (Use formula weight: nicotine, 162.2 amu.
12. There are 1.20×10^{22} **molecules** in 0.0200 mol of nicotine, a yellow liquid.
13. There are **$1.09 \times$ molecules** in 0.325g of aspirin, (use formula weight: aspirin, 180.2 amu)
14. In HCl there are 4 moles that can be formed when 2 mol of hydrogen gas react with **chlorine**.

15. One mole of hydrogen gas is needed to react with oxygen to form **one mole of water.**
16. The mass in grams of oxygen needed to react with 1.000 mol of C₃H₈ to form **carbon dioxide and water.** (use atomic weight: O, 16.00 amu)
17. Iron reacts with oxygen to form iron (III) oxide (Fe₂O₃). There will be 178.8 g of product will be formed from 125.5 g of Fe. (use atomic weights: Fe, 55.85 amu; O, 16.00 amu)
18. **3 A₂ + 2 B₂ → C + 2 D. 3.3 mol of D can be formed from 5.0 mol of A₂ and excess B**
19. To convert a given number of moles into a number of atoms, you would multiply by **6.02 x 10²³ atoms/1 mol.**

Homework CH.6

1. One standard atmosphere of pressure in units of mm Hg is **760 mm Hg.**
2. A sample of oxygen occupies 1.00 L. If the temperature remains constant, and the pressure on the oxygen is triples, the new volume is 0.333 L.
3. The volume occupied by one mole of helium at 0°C and 1 atm pressure is 22.4 L. (1 mol of any gas occupies 22.4 L at STP (standard temp, pressure))
4. A helium-filled weather balloon is launched from the ground where the pressure is 752 mmHg and the temperature is 21°C. Its volume is 75.0 L, when it has climbed to an altitude where the pressure is 89 mmHg and temperature is 0° C, its volume is 588 L.
5. In a gas-filled balloon, which has a volume of 67.0 L at a pressure of 742 mmHg and a temperature of 25°C there are **2.67 mol of gas.**
6. Carbon dioxide acts as a greenhouse gas by **absorbing infrared radiation.**

7. **Absolute temperature** is the quantity that is directly proportional to the kinetic energy of the particles in a gas.
8. H₂ is a gas that will behave most like an ideal gas.
9. At the membrane barrier in lung tissue between the blood and the surrounding atmosphere, the relationship between the partial pressure of atmospheric oxygen to that of the oxygen present in the blood is **higher**.
10. **Surface tension** is decreased with increasing temperature, is affected by temperature, is higher for polar substances than non polar ones, is lowered by surfactants, and is the same as viscosity (resistance to pouring (motion)).
11. **Evaporation** is a liquid changing to a vapor at a temperature less than its boiling point. **Sublimation** is a solid to gas without liquid phase. **Dissociation** is a liquid into ions. **Super cooling**(sweet tea in a refrigerator)
12. **Dimethylether** is a substance whose structure will not display hydrogen bonding.
13. The density of gas is proportional to its **molecular weight**.
14. Approximately 99% of the total pressure of dry air is due to molecules of N₂ and O₂. (1% is argon, 21% oxygen, and 78% is nitrogen)
15. The boiling point of a liquid is a dependent on the **atmospheric pressure**.
16. **Polar** compounds generally have higher boiling points than **non polar** compounds of similar molecular weight.
17. **Ionic compounds** tend to have higher melting points than molecular compounds.
18. **The barometer** is a device used to measure atmospheric pressure.
19. The **Pascal (Pa)** is a unit for expressing pressure.
20. **Pressure** is the experimental quantity measures force per unit area.

21. **Boyle's Law** states that the volume of a gas is inversely proportional to the pressure, if the # of moles (or mass) and the temperature of the gas are kept constant.
22. **Avogadro's Law** states that the volume of a gas is directly proportional to the absolute temperature, if the number of moles (or mass) and the pressure of the gas are kept constant.
23. The density of oxygen gas (O₂) at STP, in g/L (use molar mass: O₂, 32.0 g/mol) = 1.43×10^{-3} L
24. The volume (L) occupied by a mole of an ideal gas, if the pressure is 626 mmHg and the temperature is 25 °C is = 29.7 L.
25. A gas sample is prepared in which the components have the following partial pressures: nitrogen, 555 mmHg; oxygen, 149 mmHg; water vapor, 13 mmHg; argon, 7 mmHg, **724 mmHg** is the total pressure of the mixture.
26. **Viscosity** is the resistance of a liquid to flow.
27. **Surface tension** is a measure of the attractive forces between molecules at the surface of a liquid.
28. **Condensation** is the process that is responsible for the formation of dew on the grass early in the morning.
29. Pure water can be made to boil at a temperature above 100°C by raising the pressure **to more than one atmosphere.**
30. For a compound to display hydrogen bonding in a general structural, it must have hydrogen atoms bonded to **small, electronegative atoms such as N, O or F.**
31. Hydrogen bonding is **more extensive in water** than in hydrogen fluoride, because each water molecule has two δ⁺ sites and two δ⁻ sites, all of which can be used for hydrogen bonding. **The hydrogen fluoride** has three δ⁻ sites but only one δ⁺ site,

and the shortage of the latter limits the number of hydrogen bonds which can form per molecule to half the number in water.

32. Four types of crystalline solid are: **ionic solid, NaCl; covalent solid, diamond; molecular solid, ice; metallic solid, iron.**
33. The two types of matter that are least compatible are **solids and liquids.**

Homework CH.7

1. **Carbon dioxide and water** are the two products formed when octane (C_7H_{18}) burns completely in excess oxygen gas.
2. One method to distinguish between a beaker containing a true solution and one containing a colloidal suspension is **when a direct narrow beam of light horizontally through the two beakers. The colloidal suspension will scatter light (the Tyndall effect) making the beam visible as it passes through; the true solution will show no scattering.**
3. A solution contains 1.65g of NaOH in a total volume of 150.0 mL. The concentration expressed as % (W/V).
4. **Molarity** is the concentration of a solution expressed as the number of moles of solute per liter of solutions.
5. There are 0.2 mol of KNO_3 that are contained in one liter of 0.2 M KNO_3 solution.
6. The molarity of 50.0 mL of a 0.660 M NaOH solution after it has been diluted to 450.0 mL is 0.0733 M.
7. **A colligative property** is the concentration of a solution expressed as the number of moles of solute per liter of solution.

8. An aqueous solution is warmed from 20°C to 30°C. **The volume of the solution is likely to increase, causing a decrease in the molarity. Since mass is not affected by temperature, the molarity will stay the same.**
9. **Semipermeable membranes** allow solvent molecules to pass through but do not allow solute molecules to pass through.
10. When red blood cells are placed in a hypertonic solution the cells lose water, by osmosis, to the hypertonic solution, and they collapse. The process is known as **crenation**.
11. The osmolarity of a 2.0×10^{-3} M Na_3PO_4 solution (Na_3PO_4 is an ionic compound and produces an electrolytic solution.) is 8.0×10^{-3} **Osm**.
12. The osmotic pressure of a 6.0×10^{-2} M solution of NaCl at 20°C (293K) is 2.9 **atm**.
13. Water is referred to as the "**Universal Solvent**."
14. **Dialysis and osmosis** are similar because both involve the selective movement of small molecules through a membrane, from a solution of high concentration of those molecules to one of lower concentration. They differ in that osmosis involves only movement of solvent (water) molecules, whereas in dialysis, solute molecules can also pass through the membrane.
15. In a true solution **filtration** is the process cannot separate solute from solvent.
16. The solubility of gases in liquids is **highest** at low temperature and high pressure.
17. When 15.0g of sodium chloride is dissolved in enough water to make 300.0 mL of solution the concentration is 5.00 % (W/V).
18. In a 0.1250 M KCl solution containing 2.330g of KCl (use formula weight: KCl, 74.55 amu) is 250.0mL.
19. Density is not a colligative property of a solution.

20. The concentration of Mg^{2+} in solution is 3.0×10^{-3} ; **the concentration is expressed at 6.0 meq/L.**
21. **Nitrates and acetates** are generally soluble.
22. The usual products of an acid-base reaction are a **salt and water.**
23. In normal room lighting, the eye **cannot** distinguish a true solution from a collidal one.
24. Colligative properties depend **only** on the concentration of solute particles, not on their identity.

Homework CH.8

1. **Thermodynamics** deals with energy changes in chemical reactions.
2. The three thermodynamic quantities represented by H, S and G are: **enthalpy, entropy, and free energy.**
3. The system in thermodynamics is the reaction or process being studied; in surroundings in thermodynamics are **the remainders of the universe.**
4. The first law of thermodynamic is also known as the **law of conservation of energy.**
5. The name and symbol of the thermodynamic quantity (function) which is equal to the heat absorbed or liberated in a chemical reaction at constant pressure is **enthalpy change, ΔH° .**
6. 1000 cal is equivalent to one nutritional Calorie.
7. **The combustion of carbohydrates** is the process for the main source of energy in the human body.

8. A substance which speeds up a chemical reaction without being consumed in the process is a **catalyst**.
9. A biological catalyst is also known as **enzyme**.
10. **LeChatelier's principle** is if a stress is applied to system at equilibrium, the equilibrium will shift in such a way as to minimize that stress.
11. When a cold-pack is activated, a chemical reaction occurs and the temperature of the pack contents drops sharply. **The reaction is endothermic; $\Delta H^\circ > 0$.**
12. The measure of the randomness or disorder of a chemical system is **entropy**.
13. Solid has the **lowest** entropy.
14. Oxygen gas, under normal conditions, has the **highest** entropy.
15. A granolabar contains 185 nutritional Calories. This **774kJ (kilojoules)**.
16. Ethylene glycol has specific heat of 0.578 cal/ (g. °C). If 23,2g of ethylene glycol absorbs 75.6 cal of heat energy, the temperature will be **5.64°C**.
17. **Adding** a catalyst can alter the activation energy of a reaction.
18. One effect of a catalyst being **added** to a reaction mixture is to provide anew pathway for the reaction.
19. In kinetics, the order of a reaction depends on the **rate constant**.
20. When a chemical reaction reaches **equilibrium** the rates of the forward and reverse reactions are equal.
21. At equilibrium, the concentration of A is 0.381 M and that of B is 0.154 M. The value of equilibrium constant, **0.0622 Keq**.
22. The reaction $\text{N}_2(\text{g}) + 3\text{H}_2(\text{g}) \longrightarrow 2\text{NH}_3(\text{g})$ is at equilibrium. The equilibrium position will shift to the right when the effect of hydrogen gas is added.

23. Exothermic reactions are often, but not always, **spontaneous**.
24. A reaction that leads to a **decrease** in the free energy of the system is always spontaneous.
25. In general, a liquid state will have **higher** entropy than a solid state.

Homework CH.9

1. Another name for protonated water molecule is **hydronium ion**.
2. **In the Bronsted-Lowry theory**, an acid donates a proton and in the process becomes a base. Thus, a conjugate acid-base pair is a pair of species which differ by one proton. E.g.; H₂O and OH⁻.
3. The conjugate acid of NH₃ is NH₄⁺.
4. The conjugate base of HNO₃ is NO₃⁻.
5. The fundamental difference between a strong acid and weak acid is a strong acid **dissociates completely in solution**. A weak acid **dissociates only partially**, forming relatively fewer hydronium ions than a strong acid.
6. **The auto-ionization of water** is the transfer of a proton from one water molecule to another, producing a hydronium ion and a hydroxide ion.
7. **Neutralization** is the reaction between an acid and a base.
8. **An indicator** is a substance which is used to show changes in pH by its change in color.
9. The name of the process in which we carefully measure the volume of a solution of known concentration needed to neutralize a solution of unknown concentration is **titration**.
10. A diprotic acid is capable of donating two protons, e.g., **H₂SO₄**.

11. **Buffer capacity** refers to the amount of added acid or base which the buffer solution can neutralize without undergoing a substantial change in pH.
12. A **voltaic cell** utilizes a spontaneous oxidation-reduction reaction to produce electrical energy; an electrolytic cell reverses this process, using electrical energy to drive a non-spontaneous oxidation-reduction reaction.
13. In a voltaic cell, oxidation occurs at the **anode** while reduction occurs at the **cathode**.
14. A **base** tastes bitter, feels slippery, is corrosive, and causes many metal ions to precipitate.
15. The hydronium ion concentration of pure water at 25°C is **1.0×10^{-7}** .
16. The hydronium ion concentration of a solution with a pH of 6.0 is 1×10^{-6} .
17. The pH of a 1.0×10^{-4} M solution of KOH is **10.00**.
18. The pH of a solution that has $[\text{H}_3\text{O}^+] = 6.0 \times 10^{-3}$ M is 2.22.
19. If 14.8 mL of 0.100 M NaOH solution are needed to react with 25.0 mL of an unknown HCl solution, the molar concentration of the HCl is **0.0592 M**.
20. The reaction of an acid with a base will produce a **salt**, but it will also produce **water**.
21. Electrons are transferred from one reactant to another in **oxidation-reduction**.
22. Carbon dioxide is not a good oxidizing agent.
23. All strong bases are **metal hydroxides**.
24. Carbonic acid and the bicarbonate ion form the **main buffer system** found in blood.
25. **Emphysema** can cause blood acidosis.

26. Metals such as sodium are good **reducing agents**.
27. In **electrolysis**, electrical energy is used to drive a non- spontaneous chemical reaction.
28. The Bronsted- Lowry theory describes an acid as a proton **donor** and a base as a proton **acceptor**.

Chapters 11,12,13,14,15. &16

1. **Butane** is formed from four carbons and ten hydrogens. Butane is the first alkane to have isomers. There are two possible arrangements of the carbon skeleton- all four carbons in a row (butane) or three in a row with another carbon bound to the center carbon. (Isobutene, or 2-methyl propane)
2. The three isomers of pentane are: **pentane, 2-methylbutane, and 2,2-dimethylpropane.**
3. **Diamond (3-dimensional structure), Graphite, and Buckminster Fullerene (Bucky Balls)** are three allotropic forms of carbon.
4. The three families of aliphatic hydrocarbons are **alkane, alkenes, and alkynes.**
5. **Double and triple bonds** are never found in a saturated hydrocarbon.
6. **Functional Group** is the term for an atom or group of atoms in an organic compound that is primarily responsible for the chemical and physical properties of that compound.
7. "OH is the name for the functional group found in all alcohols.
8. CHO is the name for the functional group found in all aldehydes.
9. **Aromatic or Arene** is the major family of organic compounds that contain a benzene structure.
10. The family of organic compounds that contain hydroxyl group is **CH₃CH₂OH (Alcohols.)**

11. The main functional group present in acetone is **carbonyl**.
12. **COOH** is Acid's and **OH** is Alcohols functional group name.
13. A waxy alkane consists of molecules containing 23 **carbon atoms** and there are **C_nH_{2n+2}** many molecules in hydrogen atoms.
14. The molecular formula for butane is **C_4H_{10}** .
15. The kind of formula for organic compounds that explicitly shows all of the atoms and bonds in a molecule is **structural**.
16. **Pentane** is a linear alkane that contains five carbon atoms.
17. By **one hydrogen** is how an alkyl group differs from its parent alkane.
18. The I.U.P.A.C. name for CCl_4 is **Tetra chloromethane**.
19. $HC\equiv CH$ Acetylene is the come name for **Ethyne**.
20. The I.U.P.A.C. name for a three carbon alkene is **Propene**.
21. The I.U.P.A.C. name for a three carbon alkyne is **propyne**.
22. The class of compounds that is represented by the general formula of ROH is **alcohols**.
23. The class of compounds that is represented by the general formula RSH or is **Thiols or Mercaptans**.
24. The common name for the simplest alcohol is **methylcohol**.

25. The common name for the simplest ether is **dimethyl ether**.
26. The I.U.P.A.C. name for the compound $\text{CH}_3\text{CH}_2\text{CH}_2\text{OH}$ is **Propan-1-ol**.
27. The I.U.P.A.C. name for the alcohol produced by the fermentation of sugars and starches is **Ethanol**.
28. The common name for the simplest aldehyde is **formaldehyde**.
29. The **liver** is where in the body that ethanol is oxidized to produce acetaldehyde.
30. The compound that is produced in the liver by oxidation of methanol is **Formaldehyde or methanol**.
31. The name of the aqueous solution of formaldehyde that is available commercially as tissue preservative is **Formalin**.
32. Acetic acid is the common name for **ethanoic acid**.
33. $\sim\text{COOH}$ is the name of the functional group found in all **carboxylic acids**.
34. The I.U.P.A.C. name for and the common name of the simplest carboxylic acid is HCOOH , Formic Methanoic.
35. **Lactic Acid** is the carboxylic acid produced in muscle cells during strenuous exercise.
36. Two important medical uses of amphetamines are **diet pills and epilepsy**.