

Chem 11 Study Guide SCH3U

Unit 1

Definitions:

- **SATP:** Standard Ambient Temperature and Pressure (100kpa and 25 degrees C)
- **STP:** Standard Temperature and Pressure (101.325kp and 0 deg C)
- **IUPAC:** Intn'l Union of Pure and Applied Chem (approves, makes chem names symbols, etc.
- **Representative Elements:** an element in any of groups 1,2,13-18
- **Transition Metal:** element of groups 3-12
- **Energy level:** a space with definite and fixed energy in which an electron is allowed to move
- **Orbit:** circular/spherical path in which the electron can move around the nucleus
- **Electron Cloud:** the region of an atom in which there is a probability that an electron exists
- **Quantum Mechanics:** theory of the atom in which electrons are described in terms of their energies/probability patterns (model looks like smoke around a circle)
- **Principal Quantum # (n from $2n^2$ formula for max # e- per level):** # specifying the theoretical energy lvl of an elctron in the atom
- **Avg Atomic Mass:** weighted avg of masses of an atom's isotopes
- **Isotope:** atoms with same # of protons but diff # of neutrons
- **Isotopic Abundance:** % of an isotope in a sample of an element
- **Nuclear Charge:** positive charge on nucleus (# of protons/atomic number)
- **Position of Outermost Orbital:** dist. Btwn nucleus and valence e- (increases down cuz more lvls outweigh nuc charge, decreases across cuz same # o lvls, more nuc charge pulls elec in)
- **Shielding Effect:** extent to which elec in inner orbits affect bonding of valence elec
- **Atomic radius:** size o atom (picometrs) (increases down cuz more levels outweigh nuc charge, increases to the left cuz of less nuc charge for same number of lvls)
- **Ionic Radius:** radius as an atom becomes an ion (decreases since an element has one fewer energy level as it becomes a positive ion loses electrons) *increases downwards cuz more energy lvls to begin with even after losing, *increases to the left cuz of less nuclear charge yet same number of lvls
- **Ionization Energy:** amount of energy it takes to remove an electron from atom/ion in gas state (1st~ = farthest elec, 2nd~ = 2nd farthest, etc.) (increases upwards (requires more energy/is harder) cuz less lvls, less shielding effect = does NOT stop nuc charge making it harder and increases to right cuz more nuc charge, no increase in lvls= shielding effect = electrons pulled closer and made HARDER to remove)
- **Electronegativity:** number that describe the relative ability of an atom when bonded to attract electrons (increases upwards cuz less shielding effect and increases to right cuz less shielding effect)

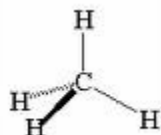
- **Electron Affinity:** the energy change that occurs when an electron is accepted by an atom in the gaseous state (same as ionization energy!)
- **Reactivity:** how reactive the element is (metals = increases down and to left cuz shielding elec block nuc charge making elec easier to remove so closer to losing enough to get an empty valence shell. Nonmetals = increases up cuz less nuclear charge, increases to right cuz more elec but no increase in shielding effect so closer to achieving full valence shell)
- **Intramolecular Forces:** the attractive forces between atoms and ions WITHIN a compound
- **Intermolecular forces:** the attractive forces BETWEEN molecules
- **Ionic Compounds Properties:** high melting point, solids at SATP, electrolyte, strong bond (crystal lattice)
- **Covalent Compounds Properties:** low melting point, liquid/gas at SATP, don't conduct/form electrolytes, flexible
- **Excited State:** any state other than ground state
- **Ground State:** lowest allowed energy state of an atom/molecule/ion
- **Catalyst:** substance that speeds up the rate of a reaction without undergoing permanent change itself (does not get consumed/changed)

Notes:

- **Diagnostic test for an ionic compound:** check if it dissolves and becomes an electrolyte
- Anything above and below any horizontal line segment of staircase is metalloid except obvious Aluminum
- **VSEPR Theory (Valence Shell Electron Pair Repulsion Theory):**
 - REMEMBER THE FOLLOWING RULES:**
 - Electron pairs repel one another and molecules will adjust their shape so that the electron pairs are as far apart as possible
 - ****Electron pairs around the CENTRAL ATOM ONLY** (both bonded and lone) influence the molecular shape! So rnr this one in particular! If it's not near the central, it's NOT counted!!!**
 - Lone pairs and bonded repel each other and themselves

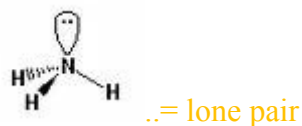
SHAPES FOR VSEPR

- **Tetrahedral:** 4 bond sites, 0 lone pairs, 109 degree angle, i.e. CH₄

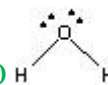


--- = behind --- = sticking out each line in general tho is a bond site ☺

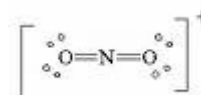
- Trigonal Pyramidal: 3 bs, 1 lp, 107 degree angle, i.e. NH₃



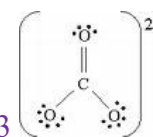
- Bent: 2 bs, 2lp, 105 degree angle i.e. H₂O



- Linear: 2 bs 0 lp, 180 deg angle, i.e. NO₂ **A
double bond, triple bond etc. is STILL ONLY ONE BOND SITE!!!
**



- Trigonal Planar: 3bs, 0 lp, 120deg angle i.e. CO₃²⁻
- ****: any diatomic molecules = linear
- ****: if more than one central atom, predict shape at each central atom SEPERATELY!



➤ POLAR BONDS:

- Using electronegativity number and VSEPR theory, you can figure out if a molecule is polar covalent, non-polar covalent, or ionic.
- If the difference in elec-ity is <0.4 (less than), it is non-polar covalent (if you already checked it's covalent in the first place lol)
- Btwn 0.5 and 1.7 AND a VSEPR shape that causes the bond polarities to cancel out, it's polar covalent. ** label the element that had the highest elec- number with "s-" and label the lower one with "s+" (meaning slightly negative/positive) and make little vector arrows on VSEPR shape pointing from s+ to s-. if they cancel, it's polar. (***there are also specific guidelines on P.85 in txt which can be used INSTEAD of this method to determine if polar bonds)
- Any elec-ity difference more than 1.7 and it's ionic

Equations/Anagrams/ Memory:

- Avg Atomic Mass = Sigma (Atomic Mass)(% Abundance)
- Carry RayBert's New Toothbrush Down = atomic theory scientists in reverse chrono order: Chadwick (Neutron), Rutherford (Nucleus), Bohr (liNe spectrum, technically Bohr came at the end but doubt this'll be asked ever so no worries about time order just rmr what each discovered), Nagaoka (saturN), Thomson (raisin buN), Dalton (Nothing (ATOM in general))
- Diagnostic Tests:
 - CO₂ = limewater turns cloudy
 - O₂ = glowing splint relights

- H₂ = lit splint makes a popping noise
- Mass of Proton and Neutron = 1, Electron = 1/1837
- Atomic # = # of protons and electrons
- Atomic Mass # - Atomic # = # of neutrons
- To draw bonding:
 - Ionic bonds: metal → nonmetal transfer
i.e. CaF₂ Ca . . . 2F (** ** ** *) [Ca]²⁺ [** **F ** **]⁻
!!!!= DRAW FINAL NUMBER OF VALENCE ELECS INSIDE SQ.
BRACKETS AND WRITE ION # ON OUTSIDE FOR MARKS
 - Covalent bonds: nonmetal + nonmetal (shared)
i.e. H₂O = H* H* . .O. . becomes H..O..H (lone pair elec's are drawn and each pair of joined elec's is a line)
!!!!= NO SQ. BRACKETS! NO ION NUMBERS!
- HOFBrINCl = Hydrogen, Oxygen, Fluorine, Bromine, Iodine, Nitrogen, Chlorine
ALL DIATOMIC MOLECULES!

Nomenclature

Naming Binary Molecular Compounds

Molecular compounds are formed from the covalent bonding between non-metallic elements. The nomenclature for these compounds is described in the following set of rules.

The more positive atom is written first (the atom which is the furthest to the left and to the bottom of the periodic table)

1. The more negative second atom has an "-ide" ending.
2. Each prefix indicates the number of each atom present in the compound.

Number of Atoms	Prefix	Number of Atoms	Prefix
1	mono	6	hexa
2	di	7	hepta
3	tri	8	octa
4	tetra	9	nona
5	penta	10	deca

3.

Examples:

CO₂ = carbon dioxide

P₄S₁₀ = tetraphosphorus decasulfide

Naming Anions

1. **Monatomic anions** have the suffix "ide added to the stem of the non-metal's name i.e., chloride (Cl^-), sulfide (S^{2-}), bromide (Br^-), oxide (O^{2-}), etc.
Group 17 anions (of halogens) are called **halides**
2. Polyatomic anions most often contain oxygen and must be learned.
 - ate indicates the base number of oxygens (typically 3 but sulfate and phosphate have 4)
 - CO_3^{2-} , NO_3^- , PO_4^{3-} , SO_4^{2-} , ClO_3^- , BrO_3^- , IO_3^- , .
carbonate, nitrate, phosphate, sulfate, chlorate, bromate, iodate

These can form a series,

- ClO_4^- perchlorate ion per means one more "O" than "ate"
- ClO_3^- chlorate ion
- ClO_2^- chlorite means one less "O" than "ate"
- ClO^- hypochlorite hypo means one less "O" than "ite"

Other polyatomic ions you should know:

Naming Inorganic Acids

1. Binary acids (H plus a nonmetal element) are acids that dissociate into hydrogen atoms and anions in water. Acids that only release one hydrogen atom are known as *monoprotic*. Those acids that release more than one hydrogen atom are called *polyprotic* acids. When naming these binary acids, you merely add "hydro-" (denoting the presence of a hydrogen atom) to the beginning and "-ic acid" to the end of the anion name.

Examples:

HCl = hydrochloric acid

HBr = hydrobromic acid

2. Ternary acids (also called oxoacids, are formed by hydrogen plus another element plus oxygen) are based on the name of the anion. In this case, the *-ate*, and *-ite* suffixes for the anion are replaced with *-ic* and *-ous* respectively. The new anion name is then followed by the word "acid." The chart below depicts the changes in nomenclature.

Anion name	Acid name
hypo__ite	hypo__ous acid
__ite	__ous acid
__ate	__ic acid
per__ate	per__ic acid

Example:

ClO_4^- to HClO_4 => perchlorate to perchloric acid

ClO^- to HClO => hypochlorite to hypochlorous acid

Reactions

- Synthesis: $A + B \rightarrow C$ (like cooking, many ingredients but only one result)
- Decomposition: $C \rightarrow A + B$
- Single displacement: $A + BC \rightarrow B + AC$ (the cations moves)
- Double displacement: $AB + CD \rightarrow CB + AD$ (the cations moves)
- Complete combustion: $A + \text{O}_2 \rightarrow \text{CO}_2 + \text{H}_2\text{O}$
- Incomplete combustion: $A + \text{O}_2 \rightarrow \text{CO}_2 + \text{H}_2\text{O} + \text{CO} + \text{C}$
- Neutralisation acid + base \rightarrow salt + water (acids have H; bases have OH)
- No Reaction: $A + B \rightarrow$ No Reaction (in certain cases, the cations may not be able to move; to find out when, observe the metal reactivity series)

Metals and alloys**Metals**

- Solids are malleable and ductile
- Conductors of heat + electricity
- High melting and boiling point

As you go down a group, the elements become more metallic

Alloys

Combinations of 2 or more metals i.e. steel or brass Rarely found in nature because metals normally require some sort of man-made chemical reaction

Quantities in Chemical Reactions**Basic Concepts****Mole**

Number of moles = n

Avogadro's Number

The number of atoms (or molecules) in one mole of a substance

6.02×10^{23}

Molar mass

The mass of one mole of a substance (element or compound)

Molar Mass = Mr

Mr = add the mass of every element in the compound

Do not count the coefficient

Unit =

$$\frac{g}{moles}$$

Law of Constant Composition

The law stating that compounds always have the same percentage comp by mass

Law of Conservation of Mass

The law stating that during a chem. Rxn, matter is neither created or destroyed

Stoichiometry

The study of quantitative (measurable) relationships of the reactants and products in balanced chemical reactions.

Percentage yield

Measure the effectiveness of the experiment.

Percentage Composition

~ is when the composition of a substance is found by using the mass and then converting the mass to a percentage

**The laboratory instrument used to determine percent comp is a combustion analyzer

$$\%part = \frac{masspart(g)}{masswhole(g)} \times 100\%$$

Empirical and Molecular formulas

Empirical formula: Simplest form of atoms

Empirical formula	Molecular formula
H ₂ O	H ₂ O
HO	H ₂ O ₂

Molar volume

The volume occupied by 1 mole of any gas
@ STP is 22.4L (called molar volume)

The volume occupied by 1 mole of any gas @ SATP is 24.8 L

Limiting reactant

A reactant completely consumed in a chemical reaction

Excess reactant

The reactants that are left behind in a chemical reaction

Problems

General Problem

1 mole of CO₂ is equal to

- a) 6.02×10^{23} molecules
- b) **2** moles of O
- c) **1** moles of C
- d) **44.0095** grams
- e) **22.4** L @ STP
- f) **31.9988** grams of O
- g) **12.0107** grams of C

Empirical Formula Problem

A compound consists of 72.2% Mg (Magnesium), 27.8% N (Nitrogen). What is the empirical formula? The Molar Masses:

1. $n(\text{Mg}) = 24.3050$

$$\frac{g}{\text{moles}}$$

2. $n(\text{N}) = 14.0067$

$$\frac{g}{\text{moles}}$$

Elements	%	Mass (g)	$n = \frac{m}{Mr}$	Divide by small	Multiply till whole
Mg	72.2	72.2	2.97	$\frac{2.97}{1.98} = 1.5$	$1.5 \times 2 = 3$
N	27.8	27.8	1.98	$\frac{1.98}{1.98} = 1$	$1 \times 2 = 2$

Answer: Mg₃N₂

Molecular Formula Problem

If the molar mass of the compound is 100.9, what is the molecular formula?

$$\frac{g}{\text{moles}}$$

Empirical formula = Mg₃N₂

Empirical mass

$$= (3 \times 24.3050) + (2 \times 14.0067)$$

$$= 100.99 \text{ moles}$$

Molecular formula = (Empirical Formula) \times

X

$$\text{Molecular Formula} = (\text{Empirical Mass}) \times X$$

$$= (\text{Mg}_3\text{N}_2) \times 1$$

$$= \text{Mg}_3\text{N}_2$$

$$= \frac{\text{Molecular mass}}{\text{Empirical Mass}}$$

$$= \frac{100.99 \text{ moles}}{100.99 \text{ moles}}$$

$$= 1$$

Answer: Mg₃N₂

Limiting and Excess Reactant Problem

Silver nitrate and sodium phosphate have reacted in equal amounts of 200 grams each. What is the limiting reactant?

Steps	Sodium nitrate	Sodium phosphate
<i>Chemical</i>	3AgNO ₃	Na ₃ PO ₄
<i>Mass (g)</i>	200g	200g
<i>Molar Mass</i> $\frac{\text{g}}{\text{moles}}$	169.87	163.94
<i>Ratio moles</i>	1.18 mol $\frac{200}{169.87} =$	1.22 moles $\frac{200}{163.94} =$
<i>Divide by coefficients</i>	0.392 $\frac{1.18}{3} =$	1.22 $\frac{1.22}{1} =$

Answer: Limiting reactant is Sodium nitrate

Formulas

$$n = m/MM$$

$$\text{Percentage yield} = \frac{\text{Actual yield}}{\text{Theoretical yield}} \times 100$$

Solution and Solubility

General terms

- The higher the temperature, the faster the reaction
 - Miscible = liquids that mix in all proportions and have no max concentration
 - Immiscible = 2 liquids that form separate layers instead of dissolving
- ****The solubility of a solid or liquid increases with temperature.
**However for a gas it decreases with temperature

Water as a solvent

Water is able to dissolve other substances. There is hardly a substance known which has not been identified in solution in water.

Concentration of a solution

How much solute and solvent there is. A solution can be *concentrated* or *dilute*.

Concentration = amount of solute/amount of solution X 100%

Solute

The substance to be dissolved

Polar Solvent

The substance doing the dissolving the dissolving (water = polar, grease/etc. = non-polar)

Solution

A homogenous mixture composed of two or more substances

Supersaturated

Completely not dissolved
dissolved anything dissolved

	Electrolyte (Y or N)	Litmus Result
Ionic Salt	Y	Neutral
Acid	Y	Red
Base	Y	Blue
Molecular	N	Neutral

Intermolecular Forces

Three types of force can operate between covalent molecules:

LDF for ALL (London Dispersion forces) weak

Dipole-dipole for polar molecules only! Usually stronger than LDF

Hydrogen Bonding: an extreme version of DDF occurs for polar molecules that contain Hydrogen AND one of F, O, or N

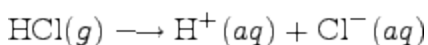
** LIKE DISSOLVES LIKE

Acids and Bases

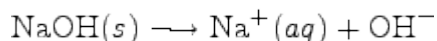
Acids	Bases
Ph [potential hydrogen] is under 7	Ph [potential hydrogen] is over 7
H ⁺ ion	OH ⁻ ion
Acid + Phenolphthalein = colourless	Base + Phenolphthalein = pink
HCl; H ₂ SO ₄ ; H ₃ PO ₄	NaOH; LiOH

Arrhenius acids and bases

According to the Arrhenius definition, an acid is any substance, which when dissolved in water, tends to increase the amount of H⁺. An example is HCl:



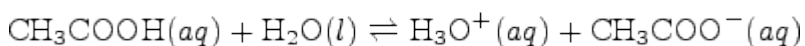
An Arrhenius base is any substance, which when dissolved in water, tends to increase the amount of OH^- . An example is NaOH:



These definitions are correct but not general enough to include the wide range of acid and base substances which are known to exist. In addition, they rely on the use of water as a solvent, which is also too narrow.

Bronsted-Lowry acids and bases

A Bronsted-Lowry (BL) acid is defined as any substance that can donate a hydrogen ion (proton) and a Bronsted-Lowry base is any substance that can accept a hydrogen ion (proton). Thus, according to the BL definition, acids and bases must come in what is called *conjugate pairs*. For example, consider acetic acid dissolved in water:



Notice that we have written $\text{H}_2\text{O}(l)$ explicitly in these reactions. The reason is that acid/base dissociation occurs by a *proton transfer reaction* from an acid species to a specific water molecule. The transfer occurs through a hydrogen bond between the acid molecule and a solvating water molecule.

Here, CH_3COOH is a BL acid because it can donate a proton, and CH_3COO^- its *conjugate* base because it can accept a proton. Note that H_2O and H_3O^+ also form such a conjugate pair.

H_3O^+
 **Note that the H_3O^+ can be called a hydronium ion. The hydronium ion is what really makes acids acids.

** A hydrogen ion (H^+) is the same thing as a single proton

Theory	Acids	Bases
Arrhenius	H^+ produced from the ionization of the molecule	OH^- produced from the dissociation of the compound
Revised	H_3O^+ produced by acid reacting with water	Base reacts with water to produce OH^-
Bronsted-Lowry	H^+ donors	H^+ acceptors

Formulas

DILUTION:

$$\frac{\text{mass}}{\text{volume}} \text{ percent} = \frac{\text{mass of solute (g)}}{\text{volume of solution (ml)}} \times 100$$

$$C_1 V_1 = C_2 V_2$$

$$\frac{\text{volume}}{\text{volume}} \text{ percent} = \frac{\text{volume of solute (L)}}{\text{volume of solution (L)}} \times 100$$

$$\text{Concentration in parts per million} = \frac{\text{mass of solute (g)}}{\text{mass of solution (g)}} \times 10^6$$

$$\frac{\text{mass}}{\text{mass}} \text{ percent} = \frac{\text{mass of solute (g)}}{\text{mass of solution (g)}} \times 100$$

$$\text{Concentration in parts per billion} = \frac{\text{mass of solute (g)}}{\text{mass of solution (g)}} \times 10^9$$

$$C = \frac{n}{V(L)}$$

Gases and Atmospheric Chemistry

Standard conditions

$$0^\circ\text{C} = 273 \text{ K}$$

$$1 \text{ atm} = 101.352 \text{ kPa} = 101,325 \text{ Pa} = 760 \text{ Hg} = 760 \text{ torr}$$

$$1 \text{ L} = 1000 \text{ ml} = 1000 \text{ cm}^3 = 1 \text{ dm}^3$$

- Temperature is always in Kelvin and is represented by a capital "T"

Laws and Formulas

Boyle's Law

- *Pressure* and *volume* are *inversely proportionate* when temperature is at a constant
- If pressure increases, volume decreases [*temperature is constant*]

$$P_1 \times V_1 = P_2 \times V_2$$

Charles' Law

- *Volume* and *temperature* are *proportionate* when pressure is at a constant

- If volume increases, temperature increases [*pressure is constant*]

$$V_1 \times T_2 = V_2 \times T_1$$

Gay-Lussac's Law

- *Pressure* and *temperature* are proportionate when volume is at a constant
- If pressure increases, temperature increases [*volume is constant*]

$$\frac{P_1}{T_1} = \frac{P_2}{T_2}$$

Combined gas Law

- The ratio between the pressure-volume product and the temperature of a system remains constant

$$P_1 \times V_1 \times T_2 = P_2 \times V_2 \times T_1$$

Dalton's law of partial pressure

- The sum of the individual pressures of all the gases that make up a mixture is equal to the total pressure

$$P_T = P_1 + P_2 + P_3 \dots_1$$

- The partial pressure of each gas is equal to the mole fraction of each gas multiplied by the total pressure

$$\frac{\text{moles gas}}{\text{total moles}} \times P_T = P_x$$

Ideal Gas Law

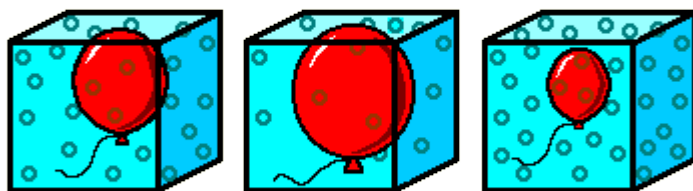
$$PV = nRT$$

P = pressure in atmosphere n = Number of moles of gas

V = volume [litres] R = Universal Gas Constant
0.0821 L/(mol.K) T = Temperature [Kelvin]

Figure 2. Volume of One Mole of Gas Under Different Conditions

All Balloons contain one mole of gas (6.02×10^{23} molecules)



A
T = Medium
P = Medium

B
T = High
P = Low

C
T = Low
P = High

Hydrocarbons and Energy

Hydrocarbons are compounds that contain only hydrogen and carbon. There are three types: alkanes, alkenes and alkynes.

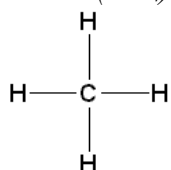
- 1) Alkanes: their general formula is C_nH_{2n+2} → single bond
- 2) Alkenes: their general formula is C_nH_{2n} → double bond
- 3) Alkynes: their general formula is C_nH_{2n-2} → triple bond

Types of alkanes

Number of Carbon atoms	Name
1	Methane
2	Ethane
3	Propane
4	Butane
5	Pentane
6	Hexane
7	Heptane
8	Octane
9	Nonane

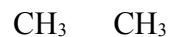
Alkanes have *single bonds* only. Diagrams of alkanes are:

Methane (CH_4)



Ethane (C_2H_6)





Structural

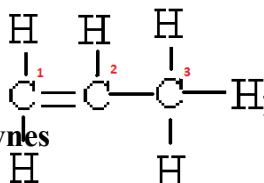
Condensed

Types of alkenes

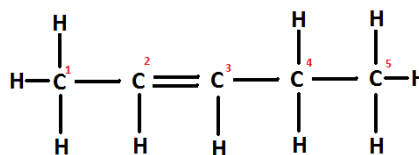
Number of Carbon atoms	Name
1	Methene
2	Ethene
3	Propene
4	Butene
5	Pentene
6	Hexene
7	Heptene
8	Octene
9	Nonene

Alkenes have *double bonds* as well as single bonds. When naming alkenes, it is important to number the carbon atoms, giving the **double bond the lowest number**:

prop-1-ene:



pent-2-ene:

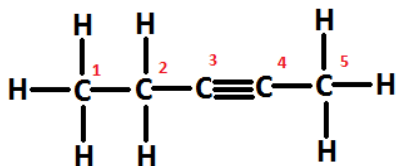


Type of alkynes

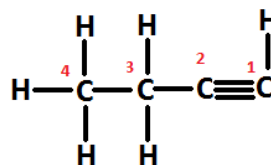
Number of Carbon atoms	Name
1	Methyne
2	Ethyne
3	Propyne
4	Butyne
5	Pentyne
6	Hexyne
7	Heptyne
8	Octyne
9	Nonyne

Alkynes have *triple bonds* as well as single bonds. When naming alkynes, as with alkenes, it is important to number carbon atoms, giving the **triple bond the lowest number**:

pent-3-yne:



but-1-yne:



Naming hydrocarbons

- 1) Find the largest continuous chain.
- 2) Identify the branches of the longest chain.
- 3) Name the branches.
- 4) If there are 2 or more branches, the carbon closest to the branch gets the lowest number.
- 5) Branches of the same type use prefixes. Commas are used between numbers and hyphens between numbers and letters.
- 6) If there is more than one type of branch, name the branches in alphabetical order.
- 7) If there is more than one chain that could be the main chain, select the chain with the most branches attached.

Naming branches

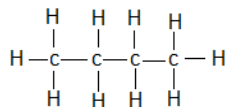
Number of Carbon atoms	Name
1	Methy
2	Ethyl
3	Propyl
4	Butyl
5	Pentyl
6	Hexyl
7	Heptyl
8	Octyl
9	Nonyl

The **prefixes for when there are more than one branch of the same type** are:

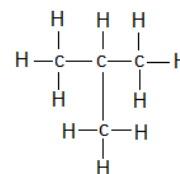
Number of branches	Prefix
1	-
2	Di
3	Tri
4	Tetra
5	Penta
6	Hexa
7	Hepta
8	Octa
9	Nona

Isomers

Isomers organic compounds with *the same molecular formula but different structural formulas*



They have the same number of carbons and hydrogen's (molecular formula), but they have different configurations

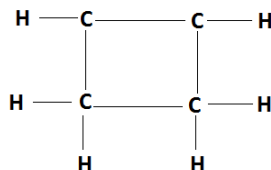


2-

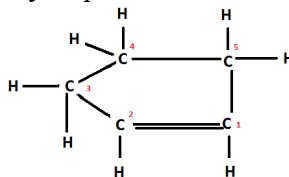
Cyclos

Cycloalkanes are hydrocarbons that form geometric shapes when drawn out into their structural diagrams. There are also cycloalkenes and cycloalkynes. The same rules apply.

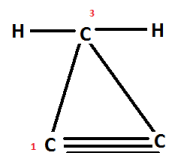
Cyclobutane:



Cyclopentene:



Cyclopropene:



Combustion

Complete combustion of *alkanes* equations are: $C_nH_n + O_2 = CO_2 + H_2O$.

Incomplete combustion of *alkanes* equations are: $C_nH_n + O_2 = CO_2 + H_2O + CO + C$

Endothermic and Exothermic reactions

Endothermic reactions absorb heat (energy) [*cold*]. Examples are dissolving a chemical in water (a medical cold pack) and any reaction where a bond is broken. *Exothermic reactions release heat* (energy) [*hot*]. Examples are combustion equations and any reaction where a bond is made.

Types of Calorimeters

Calorimeter

A calorimeter measures energy change

$$Q = m \times c \times \Delta t$$

Q is the heat content (thermal energy)
 m is the mass
 c is the specific heat capacity
 Δt is the temperature change

Q is measured in Joules (J) or Kilojoules (KJ).
 m is measured in grams (g).
 c is measured in

$$\frac{\text{Joules}}{\text{grams} \times \text{Celsius}}$$

Bomb calorimeter

Bomb calorimeters measure energy change much more accurately than the calorimeter. They are also much more expensive.

Δt is measured in °C.

$$Q = C \times \Delta t$$

Q is the heat content (thermal energy)

Q is measured in Joules (J) or Kilojoules (KJ).
 C is measured in

$$\frac{\text{Joules}}{\text{grams} \times \text{Celsius}}$$

Δt is measured in °C.

C is the heat capacity
 Δt is the temperature change

Enthalpy

$$\Delta H = \frac{-Q_w}{n}$$

"n" can change depending on the
..

Problems

Calorimeter Problem

100g of water ($c = 4.184$) are placed in a calorimeter. The heat of the water

$$\frac{\text{Joules}}{\text{grams} \times \text{Celsius}}$$

is measured to be 15.6°C . 0.568 g of calcium is added to the water, and a chemical reaction takes place. The final temperature of the solution is 26.3°C . How much heat is released per mole of calcium?

In this case it is n because the question asks how much heat is released per mole of calcium

For this question, you will use the following formulas:

$$\Delta T = T_f - T_i$$

$$Q_w = m \times c \times \Delta T$$

$$\Delta H = \frac{-Q_w}{n}$$

$$n = \frac{m}{M_r}$$

Step 1

$$\Delta T = 26.3 - 15.6$$

$$\Delta T = 10.7^\circ\text{C}$$

Step 2

$$Q_w = m \times c \times \Delta T$$

$$Q_w = 100\text{g} \times 4.18 \frac{\text{J}}{\text{g} \times ^\circ\text{C}} \times 10.7^\circ\text{C}$$

$$Q_w = 4476.88\text{J}$$

Step 3

$$n = \frac{m}{M_r}$$

$$n = \frac{0.568}{40.08}$$

$$n = 0.014$$

Step 4

$$\Delta H = \frac{-Q_w}{n}$$

$$\Delta H = \frac{-4476.88\text{J}}{0.014}$$

$$\Delta H = -315903$$

Answer:

$$\Delta H = -315903 \frac{\text{J}}{\text{mole}}$$

-315903 Joules are released per mole

