UNIT 1: MATTER MATTERS

Particle Theory: Particles are always moving, they attract each other, all matter is made up of particles, their speed depends on external factors (e.g.: temperature, pressure, etc.)
Solid: Tightly packed particles, Fixed shape and volume.
Liquid: Not tightly or loosely packed particles. No fixed shape but has a fixed volume.
Gas: Loosely packed particles. No fixed shape or volume.
Plasma: Group of charged ionized particles. It's produced when the atoms in a gas become ionized after undergoing extreme heat.
Changes of state: Melting, Freezing, Boiling, Condensing, Sublimation.

Pure substances include elements or chemically bonded compounds/molecules. Mixtures are two atoms that are not chemically bonded.

There are certain methods used to separate mixtures.
Filtration: Separating solid and liquid. Using a funnel to separate them by their particle's properties.
Distillation: Separating liquids based on their boiling points. Heating a liquid to a certain temperature so that it boils ensuring the other one does not boil. Condensing the liquid and transferring it elsewhere to separate the two liquids.
Evaporation: Boils liquid into gas and leaves behind the dissolved solid.
Chromatography: Separates components of a mixture based on their ability to move through a certain medium. E.g.: different types of ink on wet paper.

UNIT 2: SMALL MATTERS

An atom is the basic unit of matter consists of proton neutron and electron. Electrically neutral atoms have equal amount of protons and electrons. An element is a substance made out of one type of atom. A compound is made of different types of atoms. Molecules can be an element or a compound, depending on the type of atoms.
Atomic number = number of protons.
Mass number = protons + neutrons. In a compound it is referred to as atomic mass or molecular mass.
Isotope has same amount of protons as its base element but different number of neutrons.
Electrons are negatively charged particles that orbit the nucleus in its respective shell. The first shell can consist of maximum of 2 electrons and the following shells can consist of 8. To become stable, an atom must fill its shell. It can do so by gaining or losing an electron and becoming charged.

The periodic table is arranged based on increasing atomic numbers. Vertical (groups), Horizontal (periods). Periods represent energy levels of electrons, groups represent the number of valence electrons.
The elements in the first three groups are metals. They are malleable and good conductors. The elements from group 4 onwards are non-metals and are bad conductors. The groups in between are transition metals.
Group 1 consists of alkali metals, they are basic in nature and reactive. Their reactivity increases down the group (1 valence electron).
Group 7 are halogens they are reactive non-metals, their reactivity decreases down the group.
Group 8 are the noble gases with a complete outer shell, they do not tend to react with other atoms.

UNIT 3: GLUED TOGETHER:

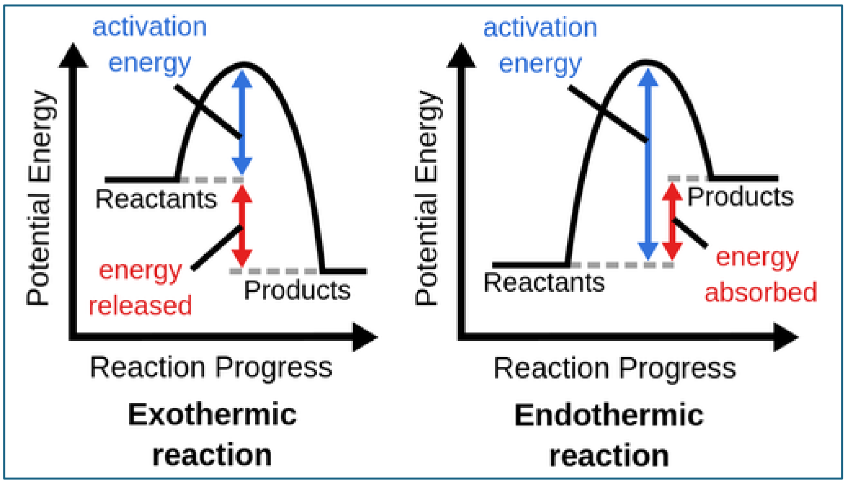
In ionic bonding (between a metal and non-metal) elements bond to form molecules and compounds with the aim of having a complete outer shell. Typically, metals lose electrons to form positive ions (cations). Non-metals gain electrons to form negative ions, anions. When looked at 3D. Ionic bonds have a crystal lattice structure. It is held together by strong electrostatic force between the protons and electrons. The structure is rigid and ordered. Ionic compounds have a high boiling point, they are conductive when molten or dissolved but never solid and require heat to break electrostatic bonds, due to their lattice structure. One shared electron forms a single bond, two shared electrons forms a double bond, etc.
Practice: Mg + Cl = $MgCl_2$, Al + Cl = $AlCl_3$, Na + 0 = Na_2O, Mg + 0 = MgO, Al + O = Al_2O_3.

In covalent bonding elements share electrons with each other to have a balanced outer shell. They can form simple molecules such as water or carbon dioxide, or giant structures such as diamond and graphite. Covalent bonds do not have a 'crystal' lattice structure, but do have a covalent lattice/covalent network which is a continuous network of atoms held together by covalent bonds. More complex compounds have larger covalent lattices. They have a low melting point and are bad conductors (except graphite). One pair of shared electrons forms a single bond, two pairs form a double bond, etc.

Metallic bonding occurs in metal atoms. It occurs between the atom's positive ions and other delocalised electrons. Metallic bonds have a crystal lattice structure like ionic bonds. They have generally high boiling points and are good conductors in any condition and are very rigid and strong.

Ionic	Covalent	Metallic
Complete transfer of electrons from metal to non-metal	Sharing of electrons between nonmetals	Force of attraction between metal ions and electrons
Strong electrostatic bond	Weak force of attraction	Strong electrostatic force
High melting and boiling point	Low melting and boiling point, except diamond and graphite	High melting and boiling points.
Ions involved, hence charges	Atoms involved, hence no charges	Cations and electrons involved, hence charges

UNIT 4: BREATHIG HOT AND COLD:



Breaking bonds requires energy, forming bonds releases energy. Exothermic reactions release heat into the surroundings, the change in heat is negative so the enthalpy change will be negative as well. Examples of these reactions include combustion, burning fuels in the presence of oxygen, neutralisation reactions.
Endothermic reactions absorb heat from the surroundings, the change in heat is positive, so the enthalpy change will be positive as well. Examples of such reactions include photosynthesis.
Activation energy is the vertical distance from the reactants to the peak energy point. It is the amount

of energy needed to start the reaction. The required energy needed to start the reaction can be reduced with a catalyst. The value of energy released/absorbed is the enthalpy change.
According to the conservation of energy, energy cannot be created or destroyed only transferred. Energy of the system and surroundings remain constant throughout.

UNIT 5: HOW FAST? HOW FAR?:

Rate of reaction is the time taken for the reaction to finish against the concentration of product formed. Formula = rate of reaction = change in concentration/time, thus measured in n/s (moles per second). This can be displayed on a graph with the time taken on the 'x' axis, and concentration of product on the 'y' axis.

There are multiple factors that affect the rate of reaction: Temperature, pressure, volume, surface area, concentration and catalyst.
Temperature increases the kinetic energy of particles resulting in more collisions and more rate of reaction (directly proportional)
Pressure only affects the rate of reaction of gases. The increase of pressure compresses the gas particles, which makes them collide more, increases the rate of reaction (directly proportional).
Volume only affects the rate of reaction of gases. More volume, means that the particles have more space to roam, and do not collide, reducing the rate of reaction (inversely proportional).
Surface area increase results in the particles being exposed to more collisions, resulting in an increase of rate of reaction (directly proportional).
Concentration increase results in more particles being present, which increases the chance of collisions, which increases the rate of reaction (directly proportional).
Catalyst reduces the activation energy and increases the rate of reaction (directly proportional).

For a reaction to occur particles must collide, collisions must have enough energy (activation energy).

UNIT 6: LET'S COUNT:

Mole (n) is used to measure the number of particles in a substance. Avogadro's constant: 1 mole = 6.022 x 10^23 particles. Formulas: n = mass/molecular mass (for solids), n = concentration/volume, mol/L (for liquids). n = volume (dm^3)/24 = volume(cm^3)/24000.

The number of atoms in a compound = the no.of moles x the subscript number x Avogadro's constant. E.g.: 0.2 moles of c6h1206. No. of carbon atoms = 0.2 x 6 x 6.022 x 10^23.
The number of molecules in a compound = molecular formula mass/formula of compound x moles x Avogadro's constant. Same example: No.of C3H603 molecules = c6h1206/c3h603 x 0.2 x 6.022x10^23 = 2 x 0.2 x 6.022 x 10^23. Or, number of CH2 molecules = c6h12/ch2 x 0.2 x 6.022x10^23 = 6 x 0.2 x 6.022x10^23.

Stoichiometry is the relationship between reactants and products in a reaction. Firstly, an equation is provided, with the mass of one reagent. Balance the equation and then find the amount of moles of the given reagent with the mass given and the molecular mass. Compare the ratio of the no.of moles of the balanced equation and the moles calculated with the given mass, between the given reagent and the one that you are trying to find. Cross multiply and find the moles of the other reagent. Use its moles and molecular mass to find its actual mass.

E.g.: Al + 02 = Al203.
Balanced: 4Al + 302 = 2Al203.

Question: 30g of Al is given find mass of oxygen needed and Al203 produced.
Number of moles of aluminium = 30/27 = 1.1
Comparing ratio: In the equation 4.4 moles of aluminium reacts with 3 moles of oxygen, with 30g of aluminium, 1.1 moles aluminium reacts with x moles of oxygen.
Cross multiplying: 4.4x – 3 x 1.1, x = 3x1.1/4 = 0.83 oxygen moles.
Mass of oxygen = 0.83 x 32(02) = 26.67g

In the equation 4 moles of aluminium forms 2 moles of aluminium oxide, with 30g of aluminium, 1.1 moles of aluminium forms y moles of aluminium oxide.
Cross multiplying: 4y = 2x1.1, y = 2x1.1/4 = 0.55. Mass of al203 = 102x0.55 = 56.10g.

Percentage composition is the percentage of each singular element in a compound. The formula is: molecular mass of element in compound/molecular mass of compound x 100.
Simple example: H20. First calculate molecular mass of compound = 2 + 16 = 18g
Calculating for hydrogen: 2/18 x 100 = 11.11%. Calculating for oxygen: 16/18 x 100 = 88.89%. Simpler way is subtracting the found one with 100, but in many cases the compound can have more than one atom.

Empirical formula is the smallest whole ratio of a compound. It could be the actual number of atoms in a compound but is not always. Molecular formula is the actual atoms in a compound.
In these questions the percentage composition is provided and the molecular formula mass. We need to assume that the compound weighs 100g. Calculate the percentage of each element easy assuming the weight is 100g, the percentage will be = the mass. Using the assumed mass calculate the moles for all of them using their respective masses and molecular masses. We will have the moles for all of the elements in the compound, some may be equal, higher, or lower. What we need to do the mole value of all of the elements by the lowest mole value found. This will leave few whole numbers for each element. Each whole number will be a subscript for the whole number value found. This is the empirical formula. It may be equal to the molecular formula, but not for sure. Find the empirical formula mass which is simply the molecular mass of the empirical formula of the compound. Use the equation molecular formula mass (given in question) = K x empirical formula mass. Find the value of K (whole number). Multiply K with the empirical formula to find the molecular formula. If K = 1, the molecular formula and empirical formula are the same.

E.g.: Glucose: c = 40%, H = 6.67%, 0 = 53.3% and the molecular formula mass = 180.156g.
Mass (assuming glucose is 100g). Carbon mass = 40g, Hydrogen mass = 6.67g, Oxygen mass = 53.3g.
Calculate moles of each:
Carbon mole = 40/12 = 3.3. Hydrogen mole = 6.67/1 = 6.67. Oxygen mole = 53.3/16 = 3.3
Dividing all of them by the lowest mole value (3.3):
Carbon = 3.3/3.3=1, Hydrogen mole = 6.67/3 = 2, Oxygen = 3.3/3.3 = 1.
Glucose Empirical Formula = ch2o.
Calculating empirical formula mass: 12 + 2 + 16 = 30g
30g x constant = molecular formula mass (180g)
constant = 60g, K= 60g.
K x Empirical Formula = Molecular formula, 6 X ch2o = c6h12o6.
Glucose Molecular Formula = c6h1206

E.g.2: 1.24 grams of phosphorus was burned completely in oxygen to give 2.84 grams of phosphorus oxide, calculate empirical and molecular formula of the oxide. 1 mole has mass of 284g

Mass of oxygen = 2.84-1.24 = 1.6g
Calculating moles of oxygen and phosphorus:
Phosphorus: 1.24/30.9 = 0.04 moles
Oxygen = 1.6/16 = 0.1
0.1/0.04, 0.04/0.04
2.5, 1
Multiply them both by 2 if it is not a whole number
5: Oxygen, 2: Phosphorus: P205 is the empirical formula.
Empirical formula mass = 142
142 x K = 284
K = 2
2 x p2o5 = P4010

E.g.3: C = 66.7%, H = 11.1%, 0 = 22.2%
66.67/12, 11.1/1, 22.2/16
5.5, 11.1, 1.3875
4, 8, 1
= C4H801
mfm = 72
k =1 ef = mf

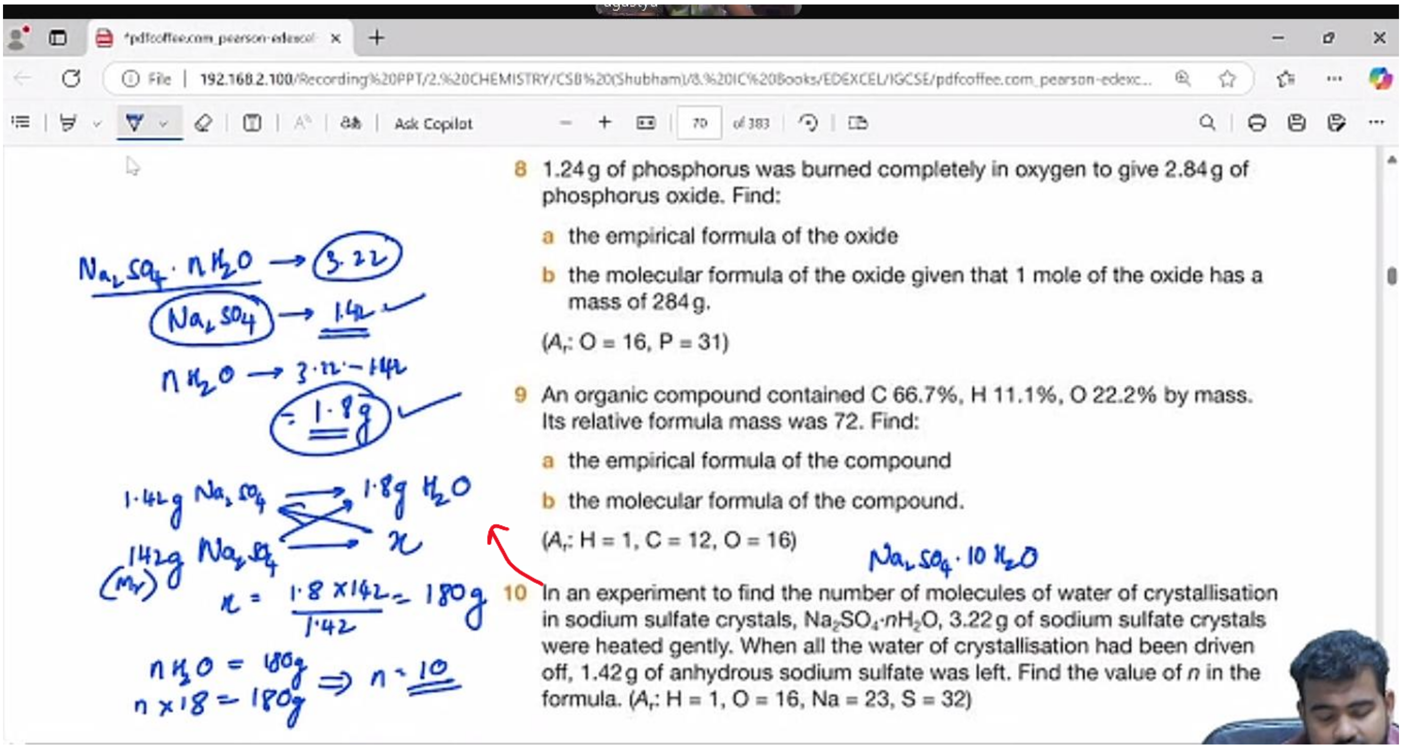
The limiting reagent in the reaction runs out first and decides the amount of product formed, the excess reagent remains after the reaction ends and the product forms. In these questions the chemical equation is given and the mass of the two reactants are provided. We are supposed to find 1) which is the limiting reagent 2) which is the excess reagent 3) amount of excess reagent, and 4) amount of product formed.
The mass of two reagents are given, ignore the second one and pretend it does not exist. Compare their molecular mass and their actual mass (still ignore the second reagent), and cross multiply and solve like you would do in normal stoichiometry. Upon doing so we will find the value for the second reagent. It will be more or less then the value given in the question. If the value is more than in the question, the second reagent is excess, if it is lesser than in the question, the second reagent is limiting. The amount of excess reagent is the mass of it (could be reagent 1 or 2) – the required mass.

To find the amount of product formed. Compare the molecular mass and mass of the limiting reagent and product, cross multiply and solve for the unknown value of the mass of the product formed.

E.g.: Haber process: $N_2 + 3H_2 = 2NH_3$. The mass of nitrogen and hydrogen is 20g and 5g.
Calculating molecular mass of all reagents and product: 28 = nitrogen, 6= hydrogen, 34 = ammonia.
Mole to mole ratio: 28g reacts with 6g (molecular mass), so 20g reacts with x grams (actual mass).
Cross multiply: 28x = 120, x = 4.28g. The required mass of hydrogen is 4.28 grams, the given mass is 5g. Thus, hydrogen is the excess reagent and nitrogen is the limiting reagent. The amount of excess reagent = 5-4.28 = 0.72g.
Calculating amount of product formed, mole to mole ratio with limiting reagent and product (normal stoichiometry):
28g reacts with 34 grams (molecular mass), 20g reacts with y grams (actual mass). Cross multiply:
28x = 34 X 20, x = 34 X 20/28, x = 24.28 grams. Amount of product formed with 20g of nitrogen in the Haber process reaction is 24.28g of ammonia.

Percentage yield is a measure of the success of the reaction, it can never be 100% due to external factors. Percentage yield = Actual Yield/Theoretical Yield X 100. Actual yield of product is measured after completing the reaction, theoretical yield is assuming the mass of product formed by doing limiting reagent and stoichiometry calculations; after plugging in the values we can get the percentage yield. The higher the percentage yield the more successful the reaction.
Simple example: 2H2 + 02 = 2H2O, the actual yield of water is 35g, oxygen is the excess reagent and 4g of hydrogen is given.
Solving: Hydrogen (h2) molar mass = 2g, molar mass of Water = 18g. Mole to mole ratio: 2g reacts with 18g, 4g reacts with x grams. Cross multiply = 2x = 4x18, x =36g (theoretical yield), Actual yield = 35g. Calculating percentage yield: 35/36 x 100 = 97.22%.

MOLES PRACTICE QUESTIONS:


8 1.24 g of phosphorus was burned completely in oxygen to give 2.84 g of phosphorus oxide. Find:

a the empirical formula of the oxide

b the molecular formula of the oxide given that 1 mole of the oxide has a mass of 284 g.

(A_r: O = 16, P = 31)

9 An organic compound contained C 66.7%, H 11.1%, O 22.2% by mass. Its relative formula mass was 72. Find:

a the empirical formula of the compound

b the molecular formula of the compound.

(A_r: H = 1, C = 12, O = 16)

10 In an experiment to find the number of molecules of water of crystallisation in sodium sulfate crystals, $Na_2SO_4 \cdot nH_2O$, 3.22 g of sodium sulfate crystals were heated gently. When all the water of crystallisation had been driven off, 1.42 g of anhydrous sodium sulfate was left. Find the value of n in the formula. (A_r: H = 1, O = 16, Na = 23, S = 32)

$Na_2SO_4 \cdot 10H_2O$

$Na_2SO_4 \cdot nH_2O \rightarrow 3.22$

$Na_2SO_4 \rightarrow 1.42$

$nH_2O \rightarrow 3.22 - 1.42$

$= 1.8g$

1.42g Na_2SO_4 → 1.8g H_2O

142g Na_2SO_4 (Mr) → x

$x = \frac{1.8 \times 142}{1.42} = 180g$

$nH_2O = 180g$

$n \times 18 = 180g \Rightarrow n = 10$

11 Gypsum is hydrated calcium sulfate, $CaSO_4 \cdot nH_2O$. A sample of gypsum was heated in a crucible until all the water of crystallisation had been driven off. The following results were obtained:

Mass of crucible = 37.34 g

Mass of crucible + gypsum, $CaSO_4 \cdot nH_2O$ = 45.94 g

Mass of crucible + anhydrous calcium sulfate, $CaSO_4$ = 44.14 g

11) Mass of caso4: 6.8g, molar mass: 136g
mass of h20: 1.8g, molar mass: 18g
caso4 . nh20
6.8 reacts with 1.8 (actual mass), 136 reacts with x (molar mass)
x = 36
n X 18 (h20 molar mass) = x (36)

n = 2
caso4 . 2h2o

12 a Calculate the amount in moles of SO_3 formed when 0.36 mol SO_2 reacts with excess O_2.

$$2SO_2 + O_2 \rightarrow 2SO_3$$

b Calculate the amount in moles of HCl that reacts with 0.4 mol $CaCO_3$.

$$CaCO_3(s) + 2HCl(aq) \rightarrow CaCl_2(aq) + H_2O(l) + CO_2(g)$$

c Calculate the amount in moles of H_2S formed when 0.4 mol of HCl reacts with excess Sb_2S_3.

$$Sb_2S_3 + 6HCl \rightarrow 2SbCl_3 + 3H_2S$$

d Calculate the amount in moles of iron formed when 0.9 mol carbon monoxide reacts with excess iron(III) oxide.

$$Fe_2O_3 + 3CO \rightarrow 2Fe + 3CO_2$$

e Calculate the amount in moles of hydrogen that would be required to make 0.8 mol NH_3.

$$N_2 + 3H_2 \rightarrow 2NH_3$$

12a) so2 molar mass = 64, so3 molar mass = 80
so2 moles = 0.36, so3 moles = x
x = 0.45 moles of so3.
b) caco3 molar mass = 100, hcl molar mass = 36.6
caco3 moles = 0.4, hcl moles = x
x = 0.1464 moles of hcl
c) h2s molar mass = 34, hcl molar mass = 36.6
h2s moles = x, hcl moles = 0.4
x = 0.371 moles of h2s.
d) co molar mass = 28, co moles = 0.9
fe molar mass = 55.8, fe moles = x
x = 1.8 moles.
e) h moles = 3, nh3 moles = 2
h moles given = x, nh3 moles given = 0.8
h moles = 1.2.

UNIT 7: pHun Meeting at the End Point:

Acids: Donate H+ ions, sour, pH7.
Bases: Donate 0H- ions, bitter, pH>7.

A polyatomic ion is two or more different type of atoms, covalently bonded with a charge:

Summary Table

Charge	Polyatomic Ions
-1	OH^-, CN^-, SCN^-, NO_3^-, NO_2^-, HCO_3^-, HSO_4^-, $C_2H_3O_2^-$, MnO_4^-, ClO^-, ClO_2^-, ClO_3^-, ClO_4^-, $H_2PO_4^-$
-2	CO_3^{2-}, SO_4^{2-}, SO_3^{2-}, CrO_4^{2-}, $Cr_2O_7^{2-}$, O_2^{2-}, HPO_4^{2-}, $C_2O_4^{2-}$
-3	PO_4^{3-}, PO_3^{3-}
+1	NH_4^+
+2	Hg_2^{2+}

Yes, there are patterns for using the suffixes "-ide", "-ate", and "-ite" in chemical nomenclature:

-ide
Used for non-metal compounds, such as the chloride ion in NaCl (sodium chloride)

-ate
Used for polyatomic ions that contain more oxygen atoms than those ending in "-ite". For example, the nitrate ion (NO3-) has more oxygen atoms than the nitrite ion (NO2-).

-ite
Used for polyatomic ions that contain fewer oxygen atoms than those ending in "-ate". For example, the nitrite ion (NO2-) has fewer oxygen atoms than the nitrate ion (NO3-).

Ionic compounds
Use "-ide" if the compound contains only two elements, and "-ate" if it contains three or more elements, including oxygen.

Oxidation state
The suffix "-ite" indicates a low oxidation state, while "-ate" indicates a high oxidation state.

Halogens and transition metals
These elements can form series of oxoanions with up to four members. The prefix "per-" is used to identify the oxoanion with the most oxygen, while the prefix "hypo-" is used to identify the anion with the fewest oxygen.

Universal indicator is used to identify acids and bases. Acids red, bases bluish purple, neutral is green, (7), water. There are different levels of acids and bases. Some acids are more acidic then others (if they are lower in the pH scale), some bases are more basic than others, (they are higher in the pH scale).

When an acid and base react it is known as a neutralisation reaction, the products of the reaction can be neutral, basic, or acidic. If the reagents are equally basic and acidic on the pH scale the product will be neutral, if one reagent is more basic, then the other is acidic, the product will be basic, and vice-versa.

Stronger acids, lower in the pH scale disassociate/ionize more than weaker ones. For example, HCl, is a strong acid as the H+ and Cl- dissociate easily and readily react, however CH3COOH does not disassociate completely, the three hydrogens bonded with carbon do not disassociate easily due to how they are bonded.
Acids can donate 1 H+ ion, 2 H+ ions, or 3H+ ions. E.g.: HCl hydrogen +1 charge, chlorine -1 charge, H2S04 (sulphuric acid) hydrogen +2 charge, S04 (sulfate), -2 charge. H3P04 hydrogen +3 charge, phosphate 3- charge.

There are four main types of neutralisation reactions:
1. Acid + Base = Salt + Water. In this water will be a constant product and won't vary based on the reagents, the salt will differ according to the nature of the reagent.
Simple example: HCl + NaOH = H2O + NaCl.
2. Acid + Metal = Salt + H2. In this reaction hydrogen is constant, just as water was in the previous reaction.
E.g.: H2SO4 + Na: To solve this use the valency, Na has a 1+ charge S04 has a 2- charge, so two sodium's must bond with sulphate to be stable which will create: Na2SO4.
H2S04 + Na = Na2S04 + h2 (Not balanced)
3. Metal oxide (has a metal and oxygen atom) + Acid = Salt + Water. Water remains constant.
E.g.: Na2O + H2S04: Na2 has a 2+ charge, S04 has a 2- charge, they bond to form Na2S04, and water

is a byproduct.
$Na_2O + H_2S04 = Na_2S04 + H_20$. (Not balanced)
4. Metal hydroxide (has a OH ion) + Acid = Salt + Water. H_20 is constant.
E.g.: Ca $(OH)_2 + H_2S04$: Ca has a 2+ charge, S04 has a 2- charge. They bond to form $CaS04$ with H_2O as a byproduct.
Ca $(OH)_2 + H_2S04 = CaS04 + H_20$.
These are few examples there are multiple other atoms that are possible.

UNIT 8: Redox Reactions:

Redox = Reduction + Oxidisation
Reduction is the addition of hydrogen atoms/electrons or loss of oxygen atoms. The atom undergoing reduction is the oxidising agent.
Oxidisation is the loss of hydrogen atoms/electrons or gain of oxygen atoms. The atom undergoing oxidation is the reduction agent
Both oxidation and reduction happen in a reaction only one cannot possible happen

Rules for oxidation number:
1. Elemental form charge = 0
2. Compound form charge = 0 (two atoms may have different charges but will add up to 0)
3. Group 1 elements = +1 charge, group 2 elements = +2 charge.
4. Aluminium has +3 charge
5. Oxygen has -2 charge, (except for -1 in h_202, and +2 in of_2)
6. Fluorine has -1 charge
7. Hydrogen has +1 charge except for hydrides (metal and oxygen bonded).

E.g.:
$C + O_2 = CO_2$
Carbon and O_2 have 0 charge as they are in elemental form. CO_2 has 0 charge as a whole when adding the charges of the both atoms. Oxygen charge = 2-. There are to oxygen atoms in CO_2, so the overall oxygen charge = -4, which will make the carbon charge 4+
Carbon went from 0 to 4 and oxygen went from 0 to -4. Which is why carbon underwent oxidisation and oxygen underwent reduction.
E.g.2:
$Zn + Cu(SO_4) = Zn(SO_4) + Cu$
In the reactants zinc is 0, SO_4 is 2-, which will make copper 2+. In the products S04 is 2- which will make zinc 2+, and copper 0 as it is in elemental form. Zinc goes from 0-2, copper goes from 2-0. Zinc lost electrons and gained charge and underwent oxidation where as copper underwent reduction which is why copper is oxidising agent and zinc is reducing agent.
E.g.3:
Ionic equation without spectator ions (like so4)
Zn + Cu = Zn + Cu
Zinc charge in reactants is 0 and copper is 2, zinc in products is 2+ charge and copper is 0.
Zinc = oxidation, copper = reduction

Electrolysis is the process of breaking bonds of a compound with electricity and then stabilizing them (in terms of outer shell electrons). The electrolysis setup consists of a cathode, anode, connected to a battery, and a container for the electrolyte. The electrolyte can be in molten or aqueous state. Molten is when the compound is heat to liquid, aqueous is when the compound is dissolved to be a liquid. When electricity passes through the electrolyte the bonds break and the atoms of different elements separate and turn into ions. The cation (positive) goes to the cathode and anion (negative) goes to the anode. In the anode oxidation occurs, and at the cathode reduction occurs. The anion at the anode undergoes oxidation and loses electrons, the electrons pass to the external circuit and go to the cation at the cathode that is undergoing reduction. The gain and loss turn the ions into stable elements as they are seen forming over the electrodes.

Example of molten electrolysis with NaCl. Na+1, and Cl-1 break bonds. Na+1 goes to the cathode, and Cl1- goes to the anode.
Anode (oxidation equation): $2Cl^{-1} – 2e = Cl_2$(gas) (Use eudiometer to store the gas)
Cathode (reduction equation): $2Na^{+1} + 2e = 2Na$(solid) (Solid forms around the cathode)

When the solution is aqueous then the water breaks down into H+ and OH- ions. These interfere in the electrolysis reaction as these ions can also go to the electrodes instead of the compound of the electrolyte. There is a series in which what element will go to the electrode due to preference. In the cathode hydrogen atoms are preferred over group 1, etc. In the anode group 7 elements are preferred the most and then hydroxides, etc. So in an aqueous salt solution instead of molten hydrogen will go the cathode and chlorine (group 7) will go to the anode.

Electrolysis with bauxite

Electroplating is the process of plating the cathode with the anode material. For example the cathode is a metal conducting spoon, and the anode is copper that needs to be plate on the cathode. For this the cathode must be conductive of electricity, and the electrolyte compound should have an element of the anode. For example, the anode is copper so the electrolyte would be $CuSO_4$ or $Cu_3(PO_4)_2$. What would happen is that the electrolyte would break into ions, copper 2+ and sulphate 2-. In electroplating the ion that is not the same as the anode material (in this case sulphate) is a spectator ion. The copper 2+ would go to the cathode and sulphate would stay at the electrolyte. This would create a vacancy in the electrolyte so the copper from the anode would lose electrons (oxidation) (these electrons would travel through the external circuit) and turn into an ion and fill the vacancy in the electrolyte. These copper ions would once again go to the cathode and the process would repeat until there is nothing at the anode, and when the cathode is plated by the anode material.

Electrorefining is the process of purifying the anode material and plating it the clean pure version on the cathode. It uses the same concept with the anode and electrolyte. When the anode material turns into an ion to fill the vacancy in the electrolyte the extra dust/impure material follows it and sinks at the bottom of the container. This process repeats and then impure metal is fully separated and can be seen floating or sinking in the container and easily separable using filteration.

Finding Oxidation number

KMnO_4 (Mn: +7)

$+1 \quad x \quad 4(-2) = 0$

$+1 + x + 4(-2) = 0$

$x - 7 = 0$

$\boxed{x = +7}$

$K_2Cr_2O_7$

$2(+1) \quad 2x \quad 7(-2) = 0$

$2 + 2x + 7(-2) = 0$

$2x - 12 = 0$

$\boxed{x = 6}$

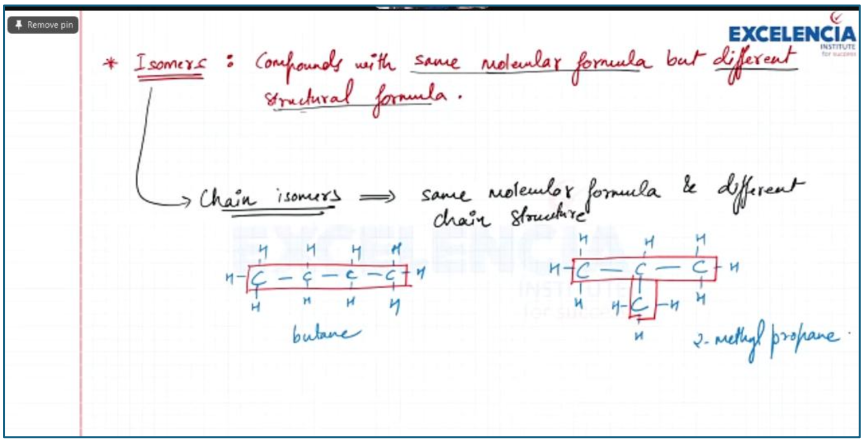
UNIT 9: Organic Chemistry:

Carbon can form multiple bonds (1-4) and form long chain structures and it has 4 valence electrons.

Homologous series are series of compounds with a common general formula but each successive member increases by ch2. All carbons in the compound must make 4 bonds.
Alkanes general formula: CnH(2n+2). (all bonds are single bonds that's how to distinguish)
Alkenes general formula: CnH2n. (can have double bond)

When C=1: meth, C=2, eth, C=3, prop, C=4, but, C=5, peut, C=6, hex. (ending is the last three letters of the homologous series).
e.g.: alkane when n=1 ch4, 1 carbon and it is an 'ane' so its methane.

Displayed formula is when you draw and show all carbon bonds. Structural formula is writing the chemical bonds in the reaction. (In book good examples). When 'n' increases the number of ch2 bonds increases for all compounds in the homologous series.



(carbons not in straight chain).