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PCM Complete Notes



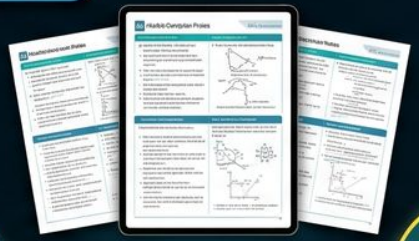
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Short Notes



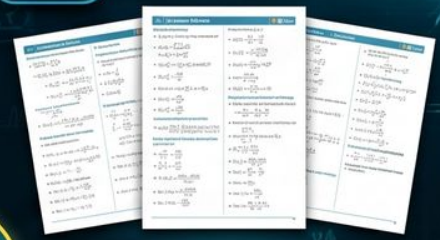
Solved PYQs (2010-2026)

Quick Revision Notes

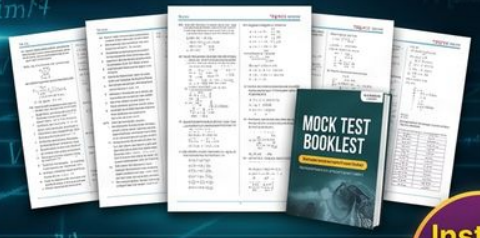


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* Coordination Compounds. *

→ Transition metals have tendency to form large no. of complex compound (coordination compd)

Central metal atom — Ligand

→ comp. in which [metal atom] is bonded to [anion/neutral] molecule by coordinate bond is coordinate comp.

Chlorophyll - Mg
Haemoglobin - Fe
Vitamin B12 - Co

* Werner's Theory.

First scientist to study about the structure

Postulates of werner theory

1) metals have 2 types of valency

i) Primary valency

→ (Oxidation no, ionizable, satisfied by -ve ligand. non-directional)

→ $CrCl_3 \rightarrow 3$
$CoCl_2 \rightarrow 2$
$PdCl_2 \rightarrow 2$

ii) Secondary valency

→ coordination no., non-ionizable, satisfied by -ve charge, neutral directional, gives definite geometry

→ $CoCl_3 \cdot 6NH_3$ $[Co(NH_3)_6]Cl_3$ ⟶ 3 mol AgCl Yellow
$CoCl_3 \cdot 5NH_3$ $[Co(NH_3)_5Cl]Cl_2$ ⟶ 2 mol AgCl Purple
$CoCl_3 \cdot 4NH_3$ $[Co(NH_3)_4Cl_2]Cl$ ⟶ 1 mol AgCl Green
$CoCl_3 \cdot 4NH_3$ $[Co(NH_3)_4Cl_2]Cl$ ⟶ 1 mol AgCl violet

isomer

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tris (ethane-1,2-diammine) cobalt (III) Sulphate

Co

$Co(en)_3$

$[Co(en)_3]^{+3} \; SO_4^{-2}$

$[Co(en)_3]_2 \; (SO_4)_3$

Naming of coordination compound

→ cation is named first than anion. then LCO Rule

→ Ligands are named first (alphabatical order) then central atom is named & then oxd^n no. is written in Roman

naming of ligand- Cl^- → chlorido

anionic ligand ends in O. $C_2O_4^{-2}$ → oxalato

O^{-2} → oxido

OH^- → Hydroxido

Neutral ligand. NH_3 - ammine

H_2O - aqua NO → nitrosyl

CO - carbonyl

→ Prefix. di, tri, tetra → denotes no. of ligand

→ if name of ligand already have (di, tri, tetra) then

bis - 2

tris - 3

tetrakis - 4

$[NiCl_2(PPh_3)_2]$

dichloro (triphenylphosphine) nickel (II)

Complex - cationic → central atom has same name

complex - anionic → central atom add ate / in el

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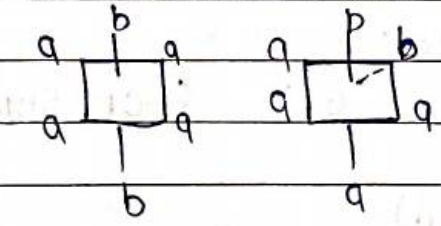
Geometrical & optical isomer in octahedral

	GI	pair of enantiomer	no. of stereoisomer (GI + OI)
Ma_6	0		
Ma_5b	0		
Ma_4b_2	2		
Ma_4bc	2		
Ma_3b_3	2		
Ma_3b_2c	3		
Ma_3bcd	4	1	5
$Ma_2b_2c_2$	5	1	6
Ma_2bcde	6	2	8
$Mabcdef$	9	6	15
$Mabcdef$	15	15	30

$[Pt(NH_3)_4Cl_2]$ or $[Co(NH_3)_4Cl_2]^+$

Ma_4b_2

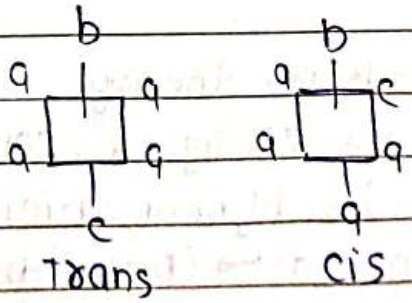
GI = 2 < cis, trans



$[Pt(NH_3)_4ClBr]$

Ma_4bc

GI = 2 < cis, trans

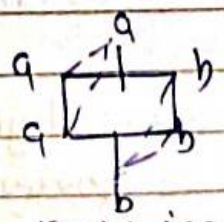


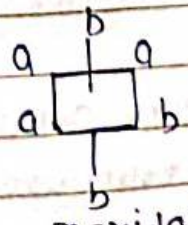
Ma_3b_3

$[Co(NH_3)_3Cl_3]$

$[Co(NH_3)_3(NO_2)_3]$

2 GI < fac, mer





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d^8 d^9 d^{10} | ↑↓ | ↑↓ | ↑↓ | ↑ | ↑ |

SFL / WFL → always sp^3d^2 outer orbital complex

↓

no pairing occurs

| ↑↓ | ↑↓ | ↑↓ | ↑↓ | ↑ |

| ↑↓ | ↑↓ | ↑↓ | ↑↓ | ↑↓ |

Magnetic moment

$[Mn(CN)_6]^{-3} \rightarrow Mn^{+3} \rightarrow 3d^4$

↳ Pairing

| 1 | 1 | 1 | 1 | |

| ↑↓ | 1 | 1 | | |

Unpaired $e^- = 2$

$MM = \sqrt{n(n+2)}$ BM

$= \sqrt{2 \times 4}$

$= \sqrt{8}$ BM

$[MnCl_6]^{-3} \rightarrow Mn^{+3} \rightarrow 3d^4$

para → WFL - no pairing

Unpaired $e^- = 4$

$MM = \sqrt{4(4+2)}$ BM

$= \sqrt{24}$ BM

$[Fe(CN)_6]^{-3} \rightarrow Fe^{+3} \rightarrow 3d^5$

para ↳ pairing

| 1 | 1 | 1 | 1 | 1 |

| ↑↓ | ↑↓ | 1 | | |

$n = 1$ $MM = \sqrt{3}$ BM

* CFT

→ electrostatic Model

→ central metal → +ve charge } electrostatical

→ Ligand → -ve charge

→ Degeneracy of d orbital breaks when ligand approaches central atom

repulsion betn e^- of d orbital & lone pair of ligand

So d orbital splitting occurs

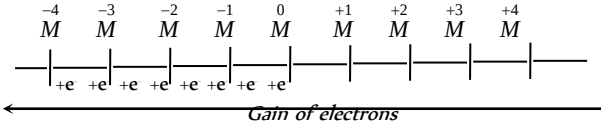
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(v) ***Decrease in +ve valency***

(a) $Fe^{3+} \rightarrow Fe^{2+}$ (*+ve valency decreases*)

(b) $[Fe(CN)_6]^{3-} \rightarrow [Fe(CN)_6]^{4-}$ (*–ve valency increases*)

(vi) ***Gain of electrons*** (also known as electronation)



(a) $Zn^{2+}(aq) + 2e^- \rightarrow Zn(S)$ (*Electronation of* Zn^{2+})

(b) $Pb^{2+} + 2e^- \rightarrow Pb^0$ (*Electronation of* Pb^{2+})

(c) $[Fe(CN)_6]^{3-} + e^- \rightarrow [Fe(CN)_6]^{4-}$

(*Electronation of* $[Fe(CN)_6]^{3-}$)

(vii) ***Decrease in oxidation number***

(a) $Mg^{2+} \rightarrow Mg^0$ (*From* +2 *to* 0)

(b) $[Fe(CN)_6]^{3-} \rightarrow [Fe(CN)_6]^{4-}$ (*From* +3 *to* +2)

(c) $Cl_2^0 \rightarrow 2Cl^-$ (*From* 0 *to* –1)

(3) **Redox-reactions**

(i) An overall reaction in which oxidation and reduction takes place simultaneously is called ***redox*** or ***oxidation-reduction reaction.*** These reactions involve transfer of electrons from one atom to another. Thus every redox reaction is made up of two ***half reactions;*** One half reaction represents the oxidation and the other half reaction represents the reduction.

(ii) ***Types of redox reaction***

(a) *Direct redox reaction* : The reactions in which oxidation and reduction takes place in the same vessel are called direct redox reactions.

(b) *Indirect redox reaction* : The reactions in which oxidation and reduction takes place in different vessels are called indirect redox reactions. Indirect redox reactions are the basis of electro-chemical cells.

(c) *Intermolecular redox reactions* : In which one substance is oxidised while the other is reduced.

For example, $2Al + Fe_2O_3 \rightarrow Al_2O_3 + 2Fe$

Here, *Al* is oxidised to Al_2O_3 while Fe_2O_3 is reduced to Fe.

(d) *Intramolecular redox reactions* : In which one element of a compound is oxidised while the other is reduced.

For example, $2KClO_3 \xrightarrow{\Delta} 2KCl + 3O_2$

Here, Cl^{+5} in $KClO_3$ is reduced to Cl^{-1} in *KCl* while O^{2-} in $KClO_3$ is oxidised to O_2^0.

Oxidising and Reducing agents

(1) **Definition :** The substance (atom, ion or molecule) that gains electrons and is thereby reduced to a low valency state is called an ***oxidising agent,*** while the substance that loses electrons and is thereby oxidised to a higher valency state is called a ***reducing agent.***

Or

An ***oxidising agent*** is a substance, the oxidation number of whose atom or atoms decreases while a *reducing agent* is a substance the oxidation number of whose atom increases.

(2) **Important oxidising agents**

(i) Molecules made up of electronegative elements.

Example : *O*, *O* and *X* (halogens).

(ii) Compounds containing an element which is in the highest oxidation state.

Example : $KMnO_4, K_2Cr_2O_7, Na_2Cr_2O_7, CrO_3, H_2SO_4,$

$HNO_3, NaNO_3, FeCl_3, HgCl_2, KClO_4, SO_3, CO_2, H_2O_2$ etc.

(iii) Oxides of elements, $MgO, CuO, CrO_3, CO_2, P_4O_{10}$, etc.

(iv) Fluorine is the strongest oxidising agent.

(3) **Important reducing agents**

(i) All metals *e.g. Na, Zn, Fe, Al,* etc.

(ii) A few non-metals *e.g. C, H, S* etc.

(iii) Hydracids : *HCl, HBr, HI, H*$_2$*S* etc.

(iv) A few compounds containing anelement in the lower oxidation state (ous).

Example : $FeCl_2, FeSO_4, SnCl_2, Hg_2Cl_2, Cu_2O$ etc.

(v) Metallic hydrides *e.g. NaH, LiH* etc.

(vi) Organic compounds like *HCOOH* and $(COOH)_2$ and their salts, aldehydes, alkanes etc.

(vii) Lithium is the strongest reducing agent in solution.

(viii) Cesium is the strongest reducing agent in absence of water. Other reducing agents are $Na_2S_2O_3$ and *KI.*

(ix) Hypo prefix indicates that central atom of compound has the minimum oxidation state so it will act as a reducing agent.

Example : H_3PO_2 (hypophosphorous acid).

(4) **Substances which act as oxidising as well as reducing agents**

Examples : $H_2O_2, SO_2, H_2SO_3, HNO_2, NaNO_2, Na_2SO_3, O_3$ etc.

(5) **Tips for the identification of oxidising and reducing agents**

(i) If an element is in its highest possible oxidation state in a compound, the compound can function as an oxidising agent.

Example : $KMnO_4, K_2Cr_2O_7, HNO_3, H_2SO_4, HClO_4$ etc.

(ii) If an element is in its lowest possible oxidation state in a compound, the compound can function only as a reducing agent.

Example : $H_2S, H_2C_2O_4, FeSO_4, Na_2S_2O_3, SnCl_2$ etc.

(iii) If an element is in its intermediate oxidation state in a compound, the compound can function both as an oxidising agent as well as reducing agent.

Example : $H_2O_2, H_2SO_3, HNO_2, SO_2$ etc.

(iv) If a highly electronegative element is in its highest oxidation state in a compound, that compound can function as a powerful oxidising agent.

Example : $KClO_4, KClO_3, KBrO_3, KIO_3$ etc.

(v) If an electronegative element is in its lowest possible oxidation state in a compound or in free state, it can function as a powerful reducing agent.

Example : I^-, Br^-, N^{3-} etc.

(6) **Equivalent weight of oxidising and reducing agents**

Equivalent weight of a substance (oxidant or reductant) is equal to molecular weight divided by number of electrons lost or gained by one molecule of the substance in a redox reaction.

$$\text{Eq. wt. of O. A.} = \frac{\text{Molecular weight}}{\text{No. of electrons gained by one molecule}}$$

$$= \frac{\text{Molecular weight}}{\text{Change in O. N. per mole}}$$

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molecular and ionic equations. The general procedure involves the following steps,

(i) Write the skeleton equation (if not given, frame it) representing the chemical change.

(ii) Assign oxidation numbers to the atoms in the equation and find out which atoms are undergoing oxidation and reduction. Write separate equations for the atoms undergoing oxidation and reduction.

(iii) Find the change in oxidation number in each equation. Make the change equal in both the equations by multiplying with suitable integers. Add both the equations.

(iv) Complete the balancing by inspection. First balance those substances which have undergone change in oxidation number and then other atoms except hydrogen and oxygen. Finally balance hydrogen and oxygen by putting H_2O molecules wherever needed.

The final balanced equation should be checked to ensure that there are as many atoms of each element on the right as there are on the left.

(v) In ionic equations the net charges on both sides of the equation must be exactly the same. Use H ion/ions in acidic reactions and OH ion/ions in basic reactions to balance the charge and number of hydrogen and oxygen atoms.

The following example illustrate the above rules,

Step **:** I $Cu + HNO_3 \rightarrow Cu(NO_3)_2 + NO_2 + H_2O$

(*Skeleton equation*)

Step: II Writing the oxidation number of all the atoms.

$$\overset{0}{Cu} + \overset{+1\ +5\ -2}{H\,N\,O_3} \rightarrow \overset{+2}{Cu}\,(\overset{+5\ -2}{N\,O_3})_2 + \overset{+4\ -2}{N\,O_2} + \overset{+1\ -2}{H_2O}$$

Step: III Change in oxidation number has occurred in copper and nitrogen.

$$\overset{0}{Cu} \rightarrow \overset{+2}{Cu}\,(NO_3)_2 \qquad \text{.....(i)}$$

$$H\,\overset{+5}{N}\,O_3 \rightarrow \overset{+4}{N}\,O_2 \qquad \text{.....(ii)}$$

Increase in oxidation number of copper = 2 units per molecule *Cu*

Decrease in oxidation number of nitrogen = 1 unit per molecule HNO_3

Step: IV To make increase and decrease equal, equation (ii) is multiplied by 2.

$$Cu + 2HNO_3 \rightarrow Cu\,(NO_3)_2 + 2NO_2 + H_2O$$

Step: V Balancing nitrate ions, hydrogen and oxygen, the following equation is obtained.

$$Cu + 4HNO_3 \rightarrow Cu\,(NO_3)_2 + 2NO_2 + 2H_2O$$

This is the balanced equation.

(2) **Ion-electron method** (half reaction method)

Jette and **LaMev** developed the method for balancing redox-reactions by ion electron method in 1927. It involves the following steps

(i) Write down the redox reaction in ionic form.

(ii) Split the redox reaction into two half reactions, one for oxidation and other for reduction.

(iii) Balance each half reaction for the number of atoms of each element. For this purpose,

(a) Balance the atoms other than H and O for each half reaction using simple multiples.

(b) Add water molecules to the side deficient in oxygen and H to the side deficient in hydrogen. This is done in acidic or neutral solutions.

(c) In alkaline solution, for each excess of oxygen, add one water molecule to the same side and 2OH ions to the other side. If hydrogen is still unbalanced, add one OH ion for each excess hydrogen on the same side and one water molecule to the other side.

(iv) Add electrons to the side deficient in electrons as to equalise the charge on both sides.

(v) Multiply one or both the half reactions by a suitable number so that number of electrons become equal in both the equations.

(vi) Add the two balanced half reactions and cancel any term common to both sides.

The following example illustrate the above rules

Step: I $I_2 + OH^- \rightarrow IO_3^- + I^- + H_2O$ *(Ionic equation)*

Step: II Splitting into two half reactions,

$I_2 + OH^- \rightarrow IO_3^- + H_2O$; $I_2 \rightarrow I^-$

(Oxidation half reaction) *(Reduction half reaction)*

Step: III Adding OH^- ions, $I_2 + 12OH^- \rightarrow 2IO_3^- + 6H_2O$

Step: IV Adding electrons to the sides deficient in electrons, (*Si*)

$I_2 + 12OH^- \rightarrow 2IO_3^- + 6H_2O + 10e^-$; $I_2 + 2e^- \rightarrow 2I^-$

Step: V Balancing electrons in both the half reactions.

$I_2 + 12OH^- \rightarrow 2IO_3^- + 6H_2O + 10e^-$; $5[I_2 + 2e^- \rightarrow 2I^-]$

Step: VI Adding both the half reactions.

$6I_2 + 12OH^- \rightarrow 2IO_3^- + 6H_2O + 10I^-$;

Dividing by 2, $3I_2 + 6OH^- \rightarrow IO_3^- + 5I^- + 3H_2O$

Autoxidation

(1) Turpentine and numerous other olefinic compounds, phosphorus and certain metals like *Zn* and *Pb* can absorb oxygen from the air in presence of water. The water is oxidised to hydrogen peroxide. This phenomenon of formation of H_2O_2 by the oxidation of H_2O is known as ***autoxidation.*** The substance such as turpentine or phosphorus or lead which can activate the oxygen is called ***activator.*** The activator is supposed to first combine with oxygen to form an addition compound, which acts as an ***autoxidator*** and reacts with water or some other acceptor so as to oxidise the latter. For example;

$$\underset{\text{(activator)}}{Pb} + O_2 \rightarrow \underset{\text{(autoxidator)}}{PbO_2} \; ; \; PbO_2 + \underset{\text{(acceptor)}}{H_2O} \rightarrow PbO + H_2O_2$$

(2) The turpentine or other unsaturated compounds which act as activators are supposed to take up oxygen molecule at the double bond position to form unstable peroxide called ***moloxide,*** which then gives up the oxygen to water molecule or any other acceptor.

$$RCH = CHR + O_2 \rightarrow RHC\!-\!CHR \quad (\text{O–O bridge})$$

$$RHC\!-\!CHR + 2H_2O \rightarrow RCH = CHR + 2H_2O_2 \quad (\text{O–O bridge})$$

$$2KI + H_2O_2 \rightarrow 2KOH + I_2$$

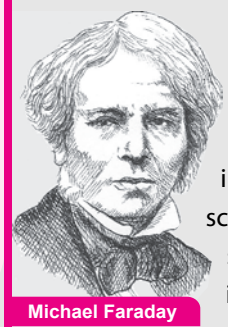
The evolution of iodine from KI solution in presence of turpentine can be confirmed with starch solution which turns blue.

(3) The concept of autoxidation help to explain the phenomenon of induced oxidation. Na_2SO_3 solution is oxidised by air but Na_3AsO_3

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BRAIN MAP

ELECTROMAGNETIC INDUCTION


Michael Faraday

A magnetic field can produce an electric field that can drive a current. This link between a magnetic field and the electric field is now known as Faraday's law of induction. The observations by Michael Faraday and other scientists which led to this law were at first just basic science. Today, however, applications of this basic science is everywhere.

MAGNETIC FLUX AND FARADAY'S LAW

- Magnetic flux $\phi_B = \vec{B}\cdot\vec{A} = BA\cos\theta$
- Faraday's Laws : It states that whenever magnetic flux linked with a coil changes, an emf is induced in the coil.

Induced emf, $\varepsilon = -N\frac{d\phi_B}{dt}$

Induced current, $I = \frac{\varepsilon}{R} = N\frac{(-d\phi_B/dt)}{R}$

Induced charge flow, $\Delta Q = I\Delta t = -N\frac{\Delta\phi_B}{R}$

- The polarity of induced emf is such that it tends to produce a current which opposes the change in magnetic flux that has produced it.

A wire moving in *B* produces emf

Motional emf

- On a straight conducting wire, $\varepsilon = Bvl$
- On a rotating conducting wire about one end $\varepsilon = \frac{B\omega l^2}{2}$

Here, $\vec{B}, \vec{v}(=\omega r\hat{v})$ and $\vec{l}$ are perpendicular to each other.

Special features of induced electric field

Induced Electric Field

- It is produced by change in magnetic field in a region. This is non-conservative in nature.

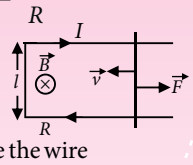
$$\oint \vec{E}\cdot d\vec{l} = -\frac{d\phi_B}{dt} = -A\frac{dB}{dt} \neq 0$$

- This is also known as integral form of Faraday's law.

Quantitative idea behind motional emf

Energy Consideration in Motional emf

- Emf in the wire = Bvl
- Induced current, $I = \frac{\varepsilon}{R} = \frac{Bvl}{R}$
- Force exerted on the wire, $F = \frac{B^2l^2v}{R}$



- Power required to move the wire $P = \frac{B^2l^2v^2}{R}$

It is dissipated as Joule heat.

Change in flux produces eddy current in a conducting plate

Eddy Current

- The currents induced in surface of bulk pieces of conductors when the magnetic flux linked with the conductor changes are known as Eddy currents. Slotting the conductor reduces the eddy current.
- Applications : Electromagnetic damping, Induction furnace, Electric power meter, Magnetic braking in trains.

Direction of induced emf/current is given by Lenz's law

Lenz's Law

- The direction of the induced current is such that it opposes the change that has induced it.
- If a current is induced by an increasing(decreasing) flux, it will weaken (strengthen) the original flux.
- It is a consequence of the law of conservation of energy.

Rotation of coil in *B* produces motional emf

Electric Generator

- Mechanical energy is converted into electrical energy by virtue of electromagnetic induction.
- Induced emf, $\varepsilon = NAB\omega\sin\omega t = \varepsilon_0\sin\omega t$
- Current, $I = \frac{NBA\omega}{R}\sin\omega t = I_0\sin\omega t$

Energy stored in the inductor (magnetic field)

Magnetic Energy

- Energy stored in an inductor $U_B = \frac{1}{2}LI^2$
- Energy stored in the solenoid, $U_B = \frac{1}{2\mu_0}B^2Al$
- Magnetic energy density, $u_B = \frac{U_B}{V} = \frac{B^2}{2\mu_0}$

Rate of change of current in the coil induces emf

Required inductance can be achieved by

Combination of Inductors

- Inductors in series, $L_S = L_1 + L_2 \pm 2M$
- Inductors in parallel, $L_P = \frac{L_1L_2 - M^2}{L_1 + L_2 + 2M}$
- If coils are far away, then $M = 0$

So, $L_S = L_1 + L_2$ and $L_P = \frac{L_1L_2}{L_1 + L_2}$

Inductance

- Emf induced in the coil/conductor, $\varepsilon = -L\frac{dI}{dt}$
- Coefficient of self induction $L = \frac{N}{I}\phi_B = \frac{-\varepsilon}{dI/dt}$
- Self inductance of a coil $L = \frac{1}{2}\mu_0\pi N^2R$
- Self inductance of a long solenoid $L = \mu_0\mu_r n^2Al = \frac{\mu_0\mu_r N^2A}{l}$
- Mutual inductance, $M = \frac{N_2\phi_2}{I_1} = \frac{-\varepsilon_2}{(dI_1/dt)} = \frac{-\varepsilon_1}{(dI_2/dt)}$
- Mutual inductance of two closely wound circular coils, $M \propto N_1N_2$
- Mutual inductance of two long coaxial solenoids

$$M = \mu_0\mu_r\pi r_1^2 n_1 n_2 l = \frac{\mu_0\mu_r N_1N_2A_1}{l}$$

- Coefficient of coupling, $k = \frac{M}{\sqrt{L_1L_2}}$

For perfect coupling, $k = 1$ so, $M = \sqrt{L_1L_2}$

Growth/decay of current in an inductor

L – R Circuit

- Current growth in L–R circuit $I = I_0(1 - e^{-t/\tau_L})$
- Current decay in L–R circuit $I = I_0(e^{-t/\tau_L})$

Here, τ_L = Time constant = $\frac{L}{R}$

$I_0 = \frac{\varepsilon}{R}$

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