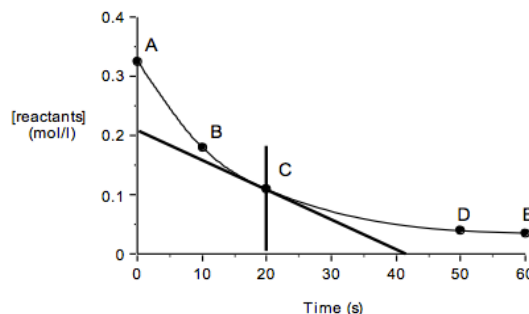# UNIT 1: REACTION KINETICS

REACTION KINETICS & RATE OF REACTION

- **reaction rate** = speed at which a reaction occurs
- compare rates of reactants and products using mole ratios

$$\text{Rate} = \frac{\Delta \text{ Amount of reactants or products}}{\Delta \text{ Time}}$$

- reaction rate can be measured by:
  a) **mass change** - open system where gas escapes; use balance & stopwatch
  b) **colour change** - change in intensity; use spectrophotometer & stopwatch
  c) **pressure change** - closed system, with a gas; use pressure gauge & stopwatch
  d) **temperature change** - exothermic, T $\uparrow$; endothermic, T $\downarrow$; use thermometer & stopwatch
  e) **pH change** - acid/base produced; use pH meter & stopwatch
- reaction rate does not remain constant
  - initially fast because [reactants] high
  - $\downarrow$ as rxn proceeds since [reactants] $\downarrow$
- rate = slope of line tangent to concentration-time graph at given point

FACTORS AFFECTING REACTION RATES

- **temperature** - T$\uparrow$, $\uparrow$RR
  - for slow rxns, 10°C$\uparrow$ doubles RR
- **concentration** - $\uparrow$ [reactant], $\uparrow$RR
- **pressure** - $\uparrow$P=$\uparrow$[ ], $\therefore\uparrow$RR
- **nature of reactants** - due to bonding, reactivity; can not control
- **surface area** - $\uparrow$SA, $\uparrow$RR (only in reactions with different phases)
- **catalyst** - $\uparrow$RR but is regenerated in original form at end of rxn
- **inhibitor** - $\downarrow$RR by combining with catalyst or reactant to prevent rxn
- **phase of reactants** - Aqueous Ions > Gases or Liquids > Solids
- few/weak bonds faster
- fewer particles faster

REACTION RATES & COLLISION THEORY

- not all collisions lead to reaction
- effectiveness of collision depends on:
  a) **geometry** - must be favourable to allow atoms to rearrange
  b) **activation energy** = minimum amount of E needed for rxn to occur
- collision theory explains:
  - effect of $\uparrow$ [ ] - closer molecules, more collisions, faster rate
  - effect of $\uparrow$ T - faster molecules, more collisions, faster rate
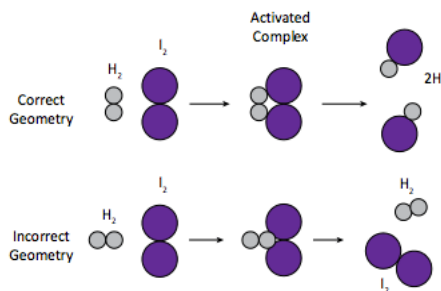

ENERGY CHANGES IN CHEMICAL REACTIONS

- **potential energy** = stored energy
  - energy of e- in bonds
  - $\uparrow$ when bonds broken; $\downarrow$ when new bonds formed
- **kinetic energy** = energy of motion
  - result of movement of molecules
  - can be related to T
- **enthalpy** (H) = heat of reaction = total KE & PE in system at constant P

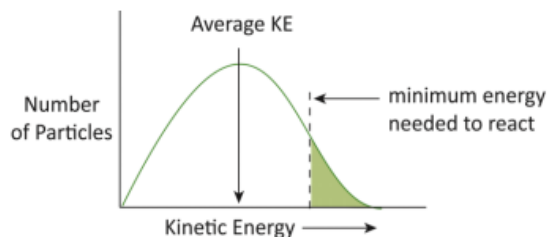**EXOTHERMIC** $\rightarrow$ H$_{products}$ < H$_{reactants}$ $\rightarrow$ ΔH is negative

ENDOTHERMIC $\rightarrow$ H$_{products}$ > H$_{reactants}$ $\rightarrow$ ΔH is positive

- **potential energy diagram**
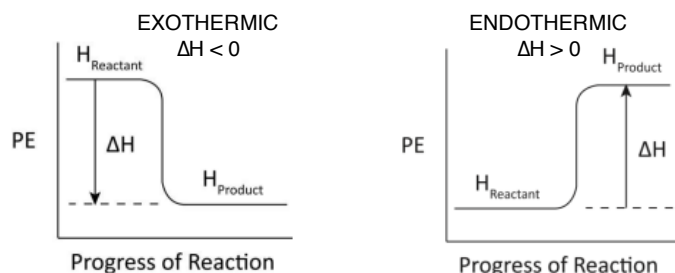

KINETIC ENERGY DISTRIBUTIONS
- only molecules with sufficient energy will react

ACTIVATION ENERGY(Ea)
= minimum amount of E required for reactants to form the activated complex
- **activated complex** = high energy, unstable arrangement of atoms which occurs when reactants are in the process of rearranging to form products
- total energy = potential energy + kinetic energy
- the higher the Ea, the slower the reaction
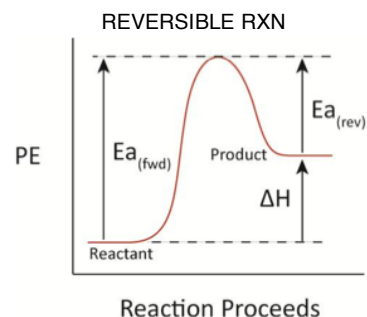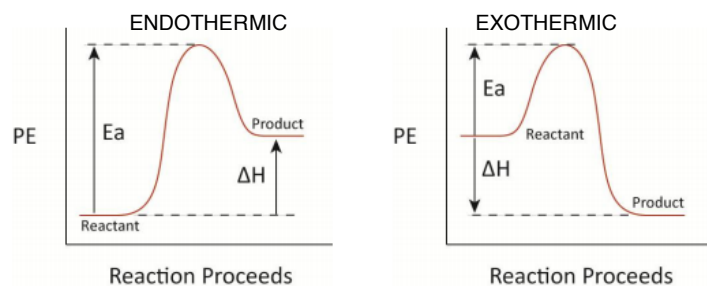
REACTION MECHANISMS
= sequence of steps which makes up an overall reaction
- **rate determining step** = the slowest step in a mechanism

EFFECTS OF CATALYSTS ON ACTIVATION ENERGY
- catalyst provides alternate mechanism having lower activation energy
- catalyst is regenerated (reactant then product)
- rxn intermediate - product then reactant
- ΔH for overall reaction remains same
- common catalysts:
  - biological catalysts = enzymes
  - platinum - common industrial catalyst
  - $H^+(aq)$ common catalyst in chemistry
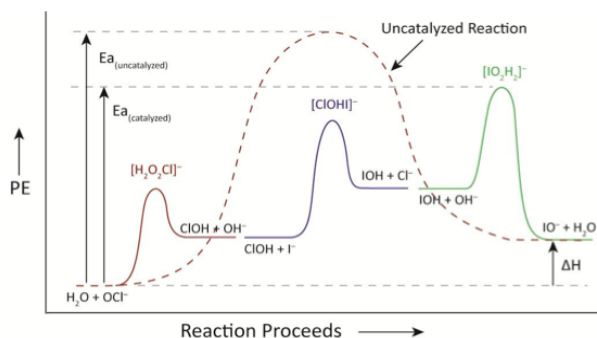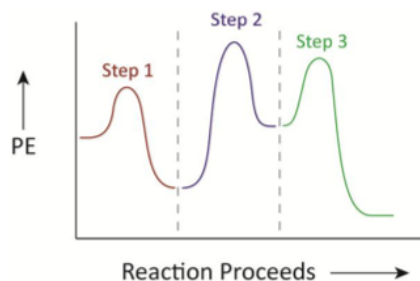
UNIT 2: REACTION EQUILIBRIUM

REVERSIBLE REACTIONS
= reactions that proceed in the forward (R→P) and reverse (P→R) direction simultaneously
- at beginning, only forward rxn; as rxn proceeds, fwd rxn slows and reverse speeds up; eventually, the system reaches **equilibrium**
- *A reversible reaction is said to be at* **equilibrium** *when the rate of the forward reaction equals the rate of the reverse reaction.*
- characteristics of equilibrium system:
  - system is closed
  - opposite reactions occur at same rate
  - equilibrium was reach starting with either reactants or products
  - T is constant
  - no change in macroscopic properties
  - rate of forward and revers reactions are equal

PREDICTING SPONTANEOUS REACTIONS
- spontaneous = occurs without outside assistance
- rxns favour side with **lower enthalpy** (ΔH)
  - exothermic - products favoured
  - endothermic - reactants favoured
  - exothermic reactions favoured - move to minimum enthalpy
- **entropy** = amount of disorder or randomness
  - rxns favour side with **maximum entropy** (ΔS)
  - gases (g) >> solutions (aq) > liquids (l) >> solids (s)
  - look also at moles on each side

LE CHATELIER'S PRINCIPLE
= when a stress is applied to a system at equilibrium, the system readjusts to relieve or offset the stress and the system reaches a new state of equilibrium
- temperature - $\downarrow$T, shift to side producing heat; $\uparrow$T shift away from heat
- concentration - $\uparrow$[ ], shift away; $\downarrow$[ ], shift to produce more
- pressure - $\uparrow$P= $\uparrow$[ ], shift away; $\downarrow$P=$\downarrow$[ ], shift to produce more
- catalyst - speeds up both fwd and rev rxns; reaches EQ faster

$2NO(g) + Cl_2(g)\uparrow \rightleftharpoons 2NOCl(g) + 76\ kJ$

[Cl$_2$] increases sharply as Cl$_2$ is injected

When equilibrium is re-established [Cl$_2$] is higher than initial since Cl$_2$ is added

EQUILIBRIUM EXPRESSIONS & EQUILIBRIUM CONSTANT
- system at EQ - [reactants] and [products] constant; ratio also constant
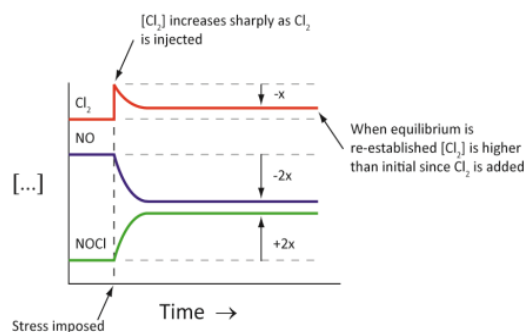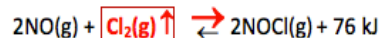- **equilibrium expression** for reaction: $\quad aA + bB \rightleftharpoons eE + fF$

$$Keq = \frac{[E]^e[F]^f}{[A]^a[B]^b} = a\ constant \qquad Keq = \frac{[products]}{[reactants]}$$

$K_{eq} > 1$ then products **GREATER THAN** reactants
$K_{eq} < 1$ then products **LESS THAN** reactants
$K_{eq} = 1$ products **EQUAL** reactants

- solids, pure liquids not included in equilibrium expression (liquid is pure if there is only one)

DEPENDENCE OF Keq ON TEMPERATURE
- T is only factor that affects Keq
- Keq increases when the T of an endothermic reaction is increased
- Keq decreases when the T of an exothermic reaction is increased

EQUILIBRIUM CONSTANT CALCULATIONS (examples)   *USE MOLARITY in ICE TABLES!*

- Suppose that 4.00 moles of HI(g) is placed into a 2.00 L flask at 425°C react to produce H_2 and I_2. At equilibrium the concentrations of H_2 and I_2 are found to each be 0.214 mol/L. Calculate the value of Keq.

$$2HI(g) \rightleftharpoons H_2(g) + I_2(g)$$

	2HI	H_2	I_2
I	2.0	0	0
C	-0.428	+0.214	+0.214
E	1.572	0.214	0.214

$$Keq = \frac{[H_2][I_2]}{[HI]^2} = \frac{(0.214)(0.214)}{(1.572)^2}$$

$$\boxed{Keq = 1.85 \times 10^{-2}}$$

- A certain amount of $NO_2(g)$ was placed into a 5.00 L bulb and reacted according to the equation: $2NO(g) + O_2(g) \rightleftharpoons 2NO_2(g)$. When equilibrium was reached, the concentration of NO(g) was 0.800 M. If Keq has a value of 24.0, how many moles of NO_2 were originally placed into the bulb?

	2NO	O_2	$2NO_2$
I	0	0	x
C	+.800	+.400	-.800
E	0.800	0.400	x-.800

$$Keq = \frac{[NO_2]^2}{[NO]^2[O_2]}$$

$$[NO_2] = \sqrt{Keq[NO]^2[O_2]}$$
$$= \sqrt{(24)(.8)^2(.4)}$$
$$= 2.48 \, M$$

$$\boxed{moles \, NO_2 = 12.4 \, mol}$$

- if asked about shift, **Trial Keq (Q)** is used
- Keq = 49 for the equilibrium:

$2NO(g) + O_2(g) \rightleftharpoons 2NO_2(g)$ If 2.0 mol of NO(g), 0.20 mol of $O_2(g)$, and 0.40 mol of $NO_2(g)$ are put into a 2.0 L bulb, which way will the reaction shift in order to reach equilibrium?

> If Q = K_{eq}, then the system is at **EQUILIBRIUM** and no shift will occur
>
> If Q < K_{eq}, then $\frac{[PRODUCTS]}{[REACTANTS]}$ is **TOO SMALL** and shift right, more **PRODUCTS**
>
> If Q > K_{eq}, then $\frac{[PRODUCTS]}{[REACTANTS]}$ is **TOO BIG** and shift left, more **REACTANTS**

> *write Keq first – points direction of shift

$$Q = \frac{[NO_2]^2}{[NO]^2[O_2]} = \frac{(0.20)^2}{(1.0)^2(0.1)} = 0.4$$

$$Keq = 49 > Q = 0.4$$

$$\Rightarrow \text{Shifts RIGHT to products}$$

UNIT 3: SOLUBILITY EQUILIBRIUM

INTRO TO SOLUBILITY

- **solution** = homogeneous mixtures, uniform throughout
- **solute** = substance in lesser [ ]
- **solvent** = substance in greater [ ]
- **electrolyte** = substance that dissolves to give electrically conducting solution containing ions
- **non-electrolyte** = substance that dissolves to give non-conducting solution of neutral molecules

> MOLECULAR SOLUTIONS contain only neutral molecules that are non-electrolytes.
> IONIC SOLUTIONS contain ions that are electrolytes.

- **saturated** = a solution that has dissolved the maximum amount at a given T
- saturation exists when:
  a) some undissolved material is present
  b) equilibrium exists between the dissolved ions & undissolved solid
- **solubility** = equilibrium concentration of the substance in solution at a given temperature
- **molar solubility** - units of mol/L (M)
- saturated solution shown with equilibrium equation of dissolving reaction & crystallization reaction

$$Ag_2SO_4(s) \rightleftharpoons 2Ag^+(aq) + SO_4^{2-}(aq)$$
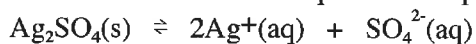
CALCULATING SOLUBILITY & ION CONCENTRATIONS
- solubility units mol/L or g/L or g/100mL of solution
- dissociation equation - shows the ions formed
- use dilution calculations if mixing solutions

# PREDICTING THE SOLUBILITY OF SALTS
- use solubility table
- **low solubility = < 0.1 M** to make a saturated solution
- when two ions form a compound having "low solubility" they will form a **precipitate** (ppt)
- compounds with alkali ions (Li^+, Na^+, K^+, Rb^+, Cs^+, Fr^+), H^+, NH_4^+, or NO_3^- are soluble in water
- when adding ions, anions added as sodium salts; cations added as nitrate salts
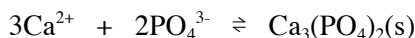
WRITING FORMULA, COMPLETE & NET IONIC EQUATIONS
- all ppt reacts used will be double replacement reactions
- formula reaction = balanced chemical reaction with reactants & products given by chemical formula

$$3Ca(NO_3)_2(aq) \ + \ 2Na_3PO_4(aq) \ \rightleftharpoons \ 6NaNO_3(aq) \ + \ Ca_3(PO_4)_2(s)$$
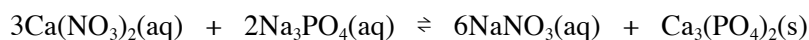
- complete ionic equation = all soluble compounds shown as ions

$$3Ca^{2+} \ + \ 6NO_3^- \ + \ 6Na^+ \ + \ 2PO_4^{3-} \ \rightleftharpoons \ 6Na^+ \ + \ 6NO_3^- \ + \ Ca_3(PO_4)_2(s)$$
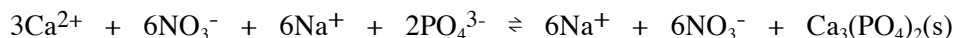
- net ionic equation = spectator ions removed

$$3Ca^{2+} \ + \ 2PO_4^{3-} \ \rightleftharpoons \ Ca_3(PO_4)_2(s)$$

SEPARATING MIXTURES OF IONS BY PRECIPITATION
- **qualitative analysis** = set of experiments to detect the presence of certain ions in a solution
- using differences in precipitation to separate mixtures of ions

THE SOLUBILITY PRODUCT & CALCULATIONS
- low solubility substances - very little to make a saturated solution $\Rightarrow$ reaches a dynamic equilibrium

$$ex. \quad PbCl_2(s) \rightleftharpoons Pb^{2+}(aq) + 2Cl^-(aq)$$
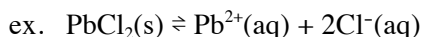
- solubility product expression $\Rightarrow$ solubility product constant $\Rightarrow$ $Ksp = [Pb^{2+}][Cl^-]^2$
- don't include solids
- see table for Ksp values
- larger Ksp $\Rightarrow$ products favoured $\Rightarrow$ greater solubility
- calculating Ksp, molar solubility and concentrations of ions

PREDICTING WHETHER A PRECIPITATE WILL FORM
- use Trial Ion Product (TIP) and compare to Ksp

> TIP > K_{sp} then **PRECIPITATE**
> TIP = K_{sp} then solution is saturated, **NO PRECIPITATE**
> TIP < K_{sp} then solution is not saturated, **NO PRECIPITATE**

> When the ion concentrations are **less than 0.1 M,** the **TRIAL ION PRODUCT (T.I.P.)** must be calculated to determine whether or not a **PRECIPITATE** will form.

SOLUBILITY EQUILIBRIA & LE CHATELIER'S PRINCIPLE
- **Le Chatelier's Principle** = when a stress is applied to a system at equilibrium, the system readjusts to relieve or offset the stress and the system reaches a new state of equilibrium
- **common ion effect** = decreasing the solubility of a salt by adding another salt with similar ions

UNIT 4: ACIDS & BASES

CHARACTERISTICS OF ACIDS & BASES

- acids:
  - begin with 'H'
  - taste sour
  - conduct electricity
  - turn litmus paper red
- bases:
  - end in 'OH'
  - taste bitter
  - conduct electricity
  - turn litmus paper blue
  - feel slippery
- salts $\Rightarrow$ all other ionic compounds
- organic acids $\Rightarrow$ end in 'COOH'
- neutralization reaction: acid + base $\longrightarrow$ salt + water

> **ARRHENIUS THEORY**
>
> an **ACID** is any compound that produces hydrogen ions, $H^+(aq)$, in water
>
> a **BASE** is any compound that produces hydroxide ions, $OH^-(aq)$, in water

> H^+ is called a **PROTON**
>
> H_3O^+ is called the **HYDRONIUM ION**

BRØNSTED–LOWRY THEORY OF ACIDS & BASES

> **BRØNSTED–LOWRY THEORY**
>
> **ACID** is any substance that can **DONATE** a proton to another substance
>
> **BASE** is any substance that can **ACCEPT** a proton from another substance

> **CONJUGATE ACID–BASE PAIRS** = pairs of chemical that differ by only ONE PROTON
>
> **CONJUGATE ACID** has an extra proton
>
> **CONJUGATE BASE** lacks a proton

- **amphiprotic** = substances that can act as either an acid or a base depending on the kind of substances they react with $\Rightarrow$ negative ion with an extra 'H'

STRENGTHS OF ACIDS & BASES

- depends on how well it ionizes in water
- **strong** acids and bases = 100% ionized
  - acids: top 6 acids on table
  - bases: O^{2-} and NH_2^- and soluble metal hydroxides
- **weak** acids and bases = <100% ionized
  - acids: on left side of table
  - bases: on right side of table
- higher ionization = higher electrical conductivity

> **LEVELLING EFFECT** = All strong acids have **IDENTICAL STRENGTHS** in water because they are all **100% ionized in aqueous solutions.**

> strongest acid in aqueous solution = H_3O^+
>
> strongest base in aqueous solution = OH^-

EQUILIBRIUM CONSTANT FOR THE IONIZATION OF WATER

ACIDIC	$[H_3O^+] > [OH^-]$
NEUTRAL	$[H_3O^+] = [OH^-]$
BASIC	$[H_3O^+] < [OH^-]$

> $K_w = [H_3O^+][OH^-] = 1.00 \times 10^{-14}$ (at 25 °C)

pH & pOH

pH < 7.00	acidic solution
pH > 7.00	basic solution
pH = 7.00	neutral solution

at 25°C

> $pH = -\log[H_3O^+]$
>
> $pOH = -\log[OH^-]$

> $[H_3O^+] = \text{antilog}(-pH) = 10^{-pH}$
>
> $[OH^-] = \text{antilog}(-pOH) = 10^{-pOH}$

$[H_3O^+][OH^-] = Kw$

$[H_2O^+]$ $[OH^-]$

$pH = -\log[H_3O^+]$ $pOH = -\log[OH^-]$

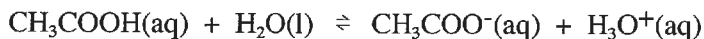pH   pOH

$pH + pOH = 14.00$
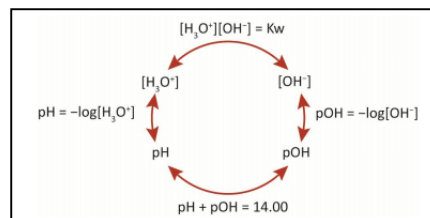
MIXING STRONG ACIDS & BASES

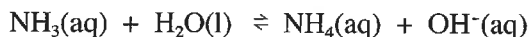- use concentrations of strong acids and bases to find $[H_3O^+]$ and $[OH^-]$

ACID & BASE EQUILIBRIUM CONSTANTS

- acid ionization reaction of weak acid CH_3COOH with water:

$$CH_3COOH(aq) + H_2O(l) \rightleftharpoons CH_3COO^-(aq) + H_3O^+(aq)$$

$$K_a = \frac{[CH_3COO^-][H_3O^+]}{[CH_3COOH]} = 1.76 \times 10^{-5}$$

- base ionization of weak base NH_3:

$$NH_3(aq) + H_2O(l) \rightleftharpoons NH_4(aq) + OH^-(aq)$$

$$K_b = \frac{[NH_4^+][OH^-]}{[NH_3]} = 1.79 \times 10^{-5}$$

THE RELATIVE STRENGTHS OF ACIDS & BASES

$$K_a \times K_b = K_w \qquad \text{(at 25 °C, } K_w = 1.00 \times 10^{-14})$$

- position of equilibrium depends on strength of acids
- stronger acids and bases will react to produce their weaker conjugates
- weaker acids and bases are favoured in an equilibrium

SALT HYDROLYSIS

= reaction between water and the cation or anion (or both) contained in the salt to produce an acidic or basic solution

> ANIONS (–) hydrolyze to give basic solutions
>
> CATIONS (+) hydrolyze to give acidic solutions

- **spectator ions** = ions that do not participate in the reaction being considered
    - spectator cations = alkali metals (group 1) and alkaline earth metals (group 2)
    - spectator ions = conjugate bases of strong acids: ClO_4^-, I^-, Br^-, Cl^- and NO_3^-
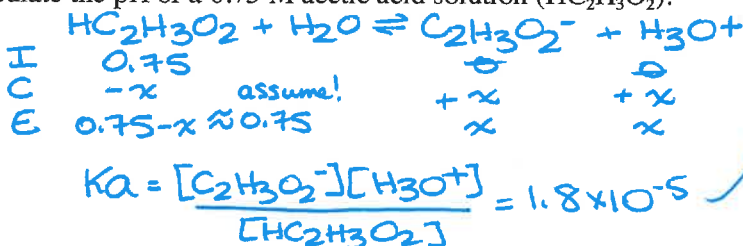- metal ions (Fe^{3+}, Cr^{3+}, Al^{3+}) react with water to form acidic solutions
- metal oxides - produce basic solutions in water
- nonmetal oxides - produce acidic solutions in water (like SO_2 producing H_2SO_4 in acid rain)
- metalloid oxides are amphiprotic

CALCULATIONS INVOLVING KA

- use ICE tables
- Calculate the pH of a 0.75 M acetic acid solution ($HC_2H_3O_2$).

$$HC_2H_3O_2 + H_2O \rightleftharpoons C_2H_3O_2^- + H_3O^+$$

	$HC_2H_3O_2$		$C_2H_3O_2^-$	H_3O^+
I	0.75		0	0
C	$-x$	assume!	$+x$	$+x$
E	$0.75-x \approx 0.75$		x	x

$$Ka = \frac{[C_2H_3O_2^-][H_3O^+]}{[HC_2H_3O_2]} = 1.8 \times 10^{-5}$$

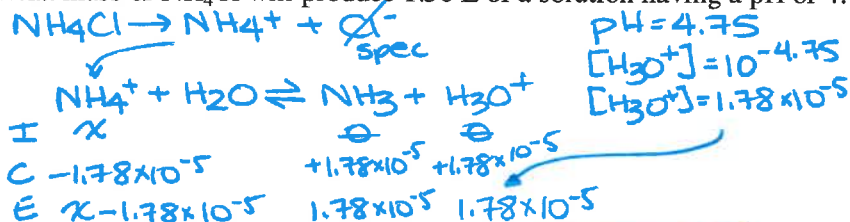$$\frac{x^2}{0.75} = 1.8 \times 10^{-5}$$

$$x = \sqrt{(1.8 \times 10^{-5})(.75)}$$

$$x = 3.67 \times 10^{-3} M$$

$$pH = -\log(3.67 \times 10^{-3})$$

$$\boxed{pH = 2.43}$$

- What mass of NH_4Cl will produce 1.50 L of a solution having a pH of 4.75?

$$NH_4Cl \rightarrow NH_4^+ + \underset{spec}{Cl^-}$$

$$NH_4^+ + H_2O \rightleftharpoons NH_3 + H_3O^+$$

	NH_4^+	NH_3	H_3O^+
I	x	0	0
C	-1.78×10^{-5}	$+1.78 \times 10^{-5}$	$+1.78 \times 10^{-5}$
E	$x - 1.78 \times 10^{-5}$	1.78×10^{-5}	1.78×10^{-5}

$$pH = 4.75$$
$$[H_3O^+] = 10^{-4.75}$$
$$[H_3O^+] = 1.78 \times 10^{-5}$$

$$Ka = \frac{[NH_3][H_3O^+]}{[NH_4^+]}$$

$$5.6 \times 10^{-10} = \frac{(1.78 \times 10^{-5})^2}{x - 1.78 \times 10^{-5}}$$
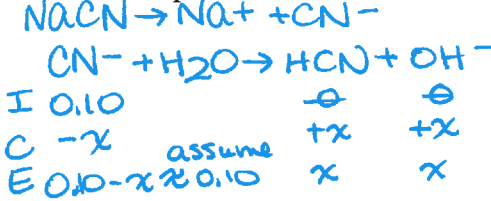
$$x = 0.565 M$$

$$g = \left(\frac{0.565 \, mol}{L}\right)(1.5 L)\left(53.5 \frac{g}{mol}\right)$$

$$= \boxed{45.3 g}$$

CALCULATIONS INVOLVING KB

- Calculate the pH of a 0.10 M NaCN.

$$NaCN \rightarrow Na^+ + CN^-$$

$$CN^- + H_2O \rightarrow HCN + OH^-$$

	CN^-	HCN	OH^-
I	0.10	0	0
C	$-x$ assume	$+x$	$+x$
E	$0.10 - x \approx 0.10$	x	x

$$Kb = \frac{[HCN][OH^-]}{[CN^-]}$$

$$2.04 \times 10^{-5} = \frac{x^2}{0.10}$$

$$x = 1.43 \times 10^{-3} = [OH^-]$$

$$pOH = 2.85$$

$$\boxed{pH = 11.15}$$

$$Kb = \frac{Kw}{Ka(HCN)} = \frac{1 \times 10^{-14}}{4.9 \times 10^{-10}} = 2.04 \times 10^{-5}$$

ACID-BASE TITRATIONS
- **titration** = quantitative chemical analysis used to determine the unknown concentration of a known reactant
- **equivalence point** = mole ratio in the reaction exactly equals the mole ratio required by the stoichiometry of the reaction

INDICATORS
- **indicator** = weak organic acid or base that has different colours for conjugate acid and base forms
- An indicator will be in its conjugate acid (HIn) form in highly acidic solutions.
- An indicator will be in its conjugate base (In–) form in highly basic solutions.
- **transition point** = indicator is mid–way through its colour change and [HIn] = [In–]
- mixtures of indicators blend their colours

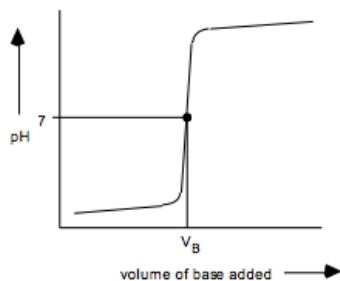> At the transition point of any indicator, the following relationships exist:
>
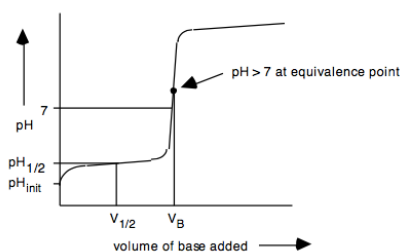> $[HIn] = [In^-]$ $K_{In} = [H_3O^+]$ $pK_{In} = pH$

TITRATION CURVES
- **standard solution** = solution with precisely known concentration
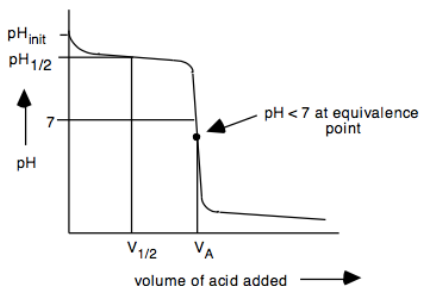- STRONG ACID with a STRONG BASE:

- pH rise almost vertical
- produces neutral solution
- pH=7 at equivalence point

- WEAK ACID with a STRONG BASE:

- initial ↑ in pH
- produces basic solution
- pH>7 at equivalence point

- WEAK BASE with a STRONG ACID:

- initial ↓ in pH
- produces acidic solution
- pH<7

BUFFER SOLUTIONS
- **buffer** = resists large changes in pH resulting from addition of acids or bases
- made of **weak acid and salt** of its conjugate base or **weak base and salt** of its conjugate acid
- buffers do not maintain a constant pH, the simply resist large changes

UNIT 5: ELECTROCHEMISTRY

CHARACTERISTICS OF REDOX

- **electrochemistry** = branch of chemistry involving conversion of chemical energy to electrical energy
- involve the transfer of electrons from one substance to another

> **OXIDATION = loss of electrons**
>
> **REDUCTION = gain of electrons**

> The **OXIDIZING AGENT** is the reactant **reduced (gains e⁻)** during a reaction.
>
> The **REDUCING AGENT** is the reactant **oxidized (loses e⁻)** during a reaction.

$$2Al + 3Cu^{2+} \rightarrow 2Al^{3+} + 3Cu$$
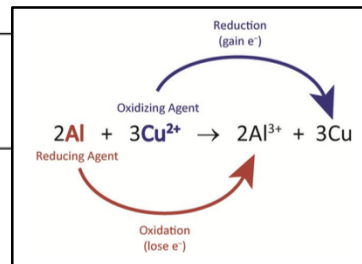
- **oxidation number** = real or apparent charge an atom or ion has when all of the bonds are assumed to be ionic
- **rules:**

Elemental form	**zero (0)**. Only one kind of atom present, no charge
Atomic ions	**= the charge on the atom** (monatomic ion)
Group 1A Li,Na,K,Rb,Cs	**+1** unless in elemental form
Group 2A Be,Mg,Ca,Sr,Ba	**+2** unless in elemental form
Hydrogen (H)	**+1** when bonded to a nonmetal, **−1** when bonded to a metal
Oxygen (O)	**−1** in peroxides O_2^-, **−2** in all other compounds (most common)
Fluorine (F)	**−1**, always
Neutral compounds	The sum of all oxidation numbers of atoms or ions in a neutral compound is **zero**.
Ionic compounds	The sum of all oxidation numbers of atoms in an ionic compound is the **charge** on the polyatomic ion.

OXIDATION = loss of electrons = increase in oxidation number

REDUCTION = gain of electrons = decrease in oxidation number

OXIDIZING AGENTS gain electrons, tend to be **CATIONS (+) OR NONMETALS**

REDUCING AGENTS lose electrons, tend to be **ANIONS (−) OR METALS**

PREDICTING SPONTANEOUS REACTIONS

- spontaneous reactions will occur when there is an **oxidizing agent** and a **reducing agent** and the oxidizing agent must be **above** the reducing agent in the table

BALANCING HALF REACTIONS

- must be balanced for mass and charge
  - balance major atoms other than oxygen and hydrogen
  - balance oxygen by adding water (H_2O)
  - balance hydrogens by adding H^+
  - if basic add OH^- to both sides to neutralize H^+
 - balance charge by adding electrons

BALANCING REDOX REACTIONS

- balance both half-reactions then add together (make sure electrons cancel out)
- balanced redox equations **do not show electrons** and the number of atoms and the total charge are balanced on both sides of the equation
- **disproportionation** = same chemical undergoes oxidation and reduction
- can balance increase and decrease in oxidation numbers; do not add electrons in this case

REDOX TITRATIONS
- for determining the concentration of unknown oxidizing or reducing agent
- acidic $KMnO_4$ is common oxidizing agent for redox titrations
- NaI or KI common reducing agent for redox titrations; two step process that uses starch as an indicator

ELECTROCHEMICAL CELLS
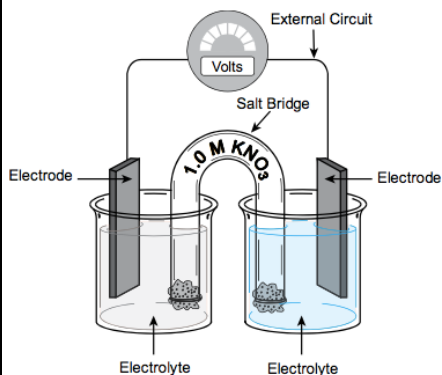
- **electrochemical cell** = spontaneous redox reaction that produces useful electrical work
- also known as **galvanic** or **voltaic** cell
- must be in two separate containers (cells)
- **anode** = electrode where **oxidation** occurs
- **cathode** = electrode where **reduction** occurs

> **Electrons flow through the wire from the anode to the cathode.**
> The salt bridge keeps the two compartments **electrically neutral by allowing ions to migrate between the cells.**

- positive cations are attracted to the cathode via salt bridge
- negative anions are attracted to the anode via salt bridge

- **voltage** = work done per electron transferred = electric potential
- $E°$ = **standard reduction potential** = voltage at standard state
- oxidation potential = negative of reduction potential
- if $E°$ is **positive**, the reaction is **spontaneous**
- if $E°$ is **negative**, the reaction is **non-spontaneous**
- surface area of an electrode does not affect its voltage
- when an electrochemical cell reaches **equilibrium**, the voltage of the cell is **0.00 V**

> The electrochemical cell potential is determined by **adding the half–cell voltages for the reduction potential and oxidation potential.**
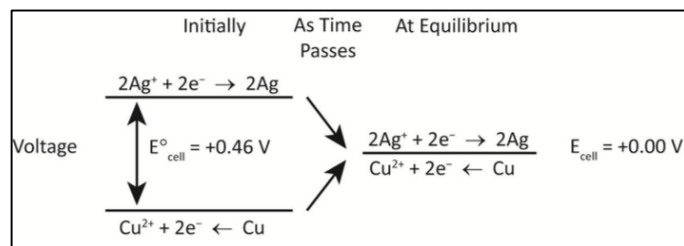
SELECTING PREFERRED REACTIONS
- when a cell consists of a mixture of substances
  - strongest oxidizing agent will become reduced
  - strongest reducing agent will become oxidized

CORROSION OF METALS
- **corrosion** = oxidation of a metal
- rust is just $Fe_2O_3 \cdot xH_2O$
- preventing corrosion
  - isolate from environment
    - protective layer – paint, grease, plastic
    - apply corrosion resistant metal (ex. tin – oxidizes first and protects Fe)
  - electrochemical methods
    - **cathodic protection** = protecting a substance from unwanted oxidation by connecting it to a stronger reducing agent
      - ex. Mg and Zn are stronger reducing agents and become oxidized first = **sacrificial anode**
        - Fe becomes the cathode
  - change the conditions
   - remove oxygen or increase $[OH^-]$ ions
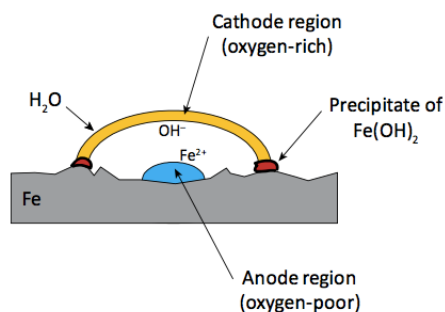
ELECTROLYSIS

- **electrolysis** = process of supplying electrical energy to a molten ionic compound or a solution containing ions so as to produce a chemical change
- involve **non-spontaneous** reactions
- require energy to occur = **endothermic**

# ELECTROLYSIS OF MOLTEN BINARY SALTS

- contain only two ions
- no salt bridge
- anode connected to + of power source
- cathode connected to – of power source
- E° is negative $\Rightarrow$ must add voltage via power supply
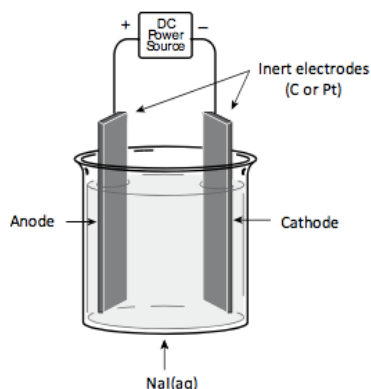
ELECTROLYSIS OF AQUEOUS SOLUTIONS

- ions from salt plus water
- water may behave as an oxidizing or reducing agent
- strongest oxidizing agent becomes reduced (lowest)
- strongest reducing agent becomes oxidized (highest)
- half-cell **overpotential** = difference between actual potential required for electrolysis and calculated potential
- variance is very big for oxidation and reduction of water
- use arrow-head to account for overpotential effect of water

ELECTROPLATING

- **electroplating** = electrolytic process in which a metal is reduced or "plated out" onto an object connected to the cathode
- **cathode** = material receiving the plating
- **electroplating solution** = contains ions of metal which is to be plated
- **anode** = same metal which is to be "plated out"
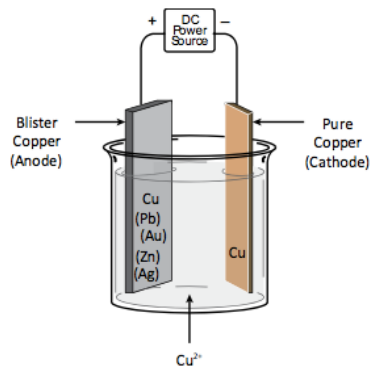
ELECTROREFINING

- **electrorefining** = process of purifying a metal by electrolysis
- "plating out" a metal from an impure anode onto a pure cathode