

QUICK REVISION



12th Chemistry
All Chapterwise Formulas

For Physics and Mathematics

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1. Common Equations Used in Chemistry

- (1) Equation for density : $d = \frac{m}{V}$
- (2) Converting °F to °C : $^{\circ}\text{C} = (^{\circ}\text{F} - 32) \times \frac{5}{9}$
- (3) Converting °C to °F: $^{\circ}\text{F} = ^{\circ}\text{C} \times \frac{9}{5} + 32$
- (4) Converting °C to K: $\text{K} = (^{\circ}\text{C} + 273.15)$
- (5) Percent composition of an element = $\frac{n \times \text{molar mass of element}}{\text{molar mass of compound}} \times 100\%$
where n = the number of moles of the element in one mole of the compound.
- (6) % yield = $\frac{\text{Actual yield}}{\text{Theoretical yield}} \times 100$
- (7) Molarity (M) = $\frac{\text{Moles of solute}}{\text{Liters of solution}}$
- (8) Dilution of Solution : $M_i V_i = M_f V_f$
- (9) Boyle's law - Constant T and n : $PV = k$
- (10) Boyle's law - for calculating changes in pressure of volume :
 $P_1 V_1 = P_2 V_2$
- (11) Charles' law - Constant P and n : $\frac{V}{T} = k$
- (12) Charles' law - for calculating temperature or volume changes :
 $\frac{V_1}{T_1} = \frac{V_2}{T_2}$
- (13) Avogadro's law - Constant P and T: $V = kn$
- (14) Ideal Gas equation : $PV = nRT$
- (15) Calculation of changes in pressure, temperature, or volume of gas when n is constant: $\frac{P_1 V_1}{T_1} = \frac{P_2 V_2}{T_2}$
- (16) Calculation of density or molar mass of gas: $d = \frac{PM}{RT}$
- (17) Dalton's law of partial pressures - for calculating partial pressures:
 $P_i = P_{\text{total}} x_i$
where, x_i is the mole fraction of the i^{th} component in the total mixture of n components.

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- (18) Root-mean-square speed of gas molecules: $C_{\text{rms}} = \sqrt{\frac{3RT}{M}}$
- (19) van der Waal's equation; for calculating the pressure of a non-ideal gas
 $(P + \frac{an^2}{V^2})(V-nb) = nRT$
- (20) Definition of heat capacity, where s is specific heat: $C = ms$
- (21) Calculation of heat change in terms of specific heat : $q = ms\Delta t$
- (22) Calculation of heat change in terms of heat capacity : $q = C\Delta t$
- (23) Electrical force: $F_{\text{el}} = k \frac{q_1 q_2}{r^2}$
- (24) Potential energy: $V = k \frac{q_1 q_2}{r}$
- (25) Calculation of standard enthalpy of reaction:
 $\Delta H_{\text{rxn}}^{\circ} = \sum nH_f^{\circ}(\text{products}) - \sum m\Delta H_f^{\circ}(\text{reactants})$
[where n and m are coefficients in equation].
- (26) Mathematical statement of the first law of thermodynamics: $\Delta E = q + W$
- (27) Work done in gas expansion or compression : $W = -P\Delta V$.
- (28) Definition of enthalpy: $H = E + PV$.
- (29) Enthalpy (or energy) change for a constant - pressure process :
 $\Delta H = \Delta E + P\Delta V$.
- (30) Enthalpy (or energy) change for a constant pressure process:
 $\Delta E = \Delta H - RT\Delta n$.
- (31) Relationship of wavelength and frequency: $c = \lambda\nu$.
where $c = 3 \times 10^8$ m/s, is the speed of light.
- (32) Energy of a photon: $E = h\nu$.
- (33) Energy of an electron in the n^{th} sate in a hydrogen atom:
 $E_n = -R_H \left(\frac{1}{n^2} \right)$ where R_H = Rydberg constant = 2.18×10^{-18} J
- (34) Energy of a photon emitted as the electron undergoes a transition from the n_i level to the n_f level :
 $\Delta E = h\nu = R_H \left(\frac{1}{n_i^2} - \frac{1}{n_f^2} \right)$ where R_H = Rydberg constant = 2.18×10^{-18} J.
- (35) DeBroglie relationship of wavelength of a particle to its mass m and velocity v : $\lambda = \frac{h}{mv}$.

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- (36) Uncertainty in the position (x) or in the momentum (p) of a particle:

$$\Delta x \Delta p \geq \frac{h}{4\pi}$$

- (37) Formal charge on an atom in a Lewis structure = Total number of valence electrons in the free atom – Total number of non-bonding electrons – $\frac{1}{2}$ (total number of bonding electrons).

- (38) Enthalpy change of a reaction for bond energies:

$$\Delta H^\circ = \Sigma \text{B.E. (reactants)} - \Sigma \text{B.E. (products)}$$

- (39) Dipole moment in terms of charge (Q) and distance of separation (r) between charges : $\mu = Q \times r$.

- (40) Bragg equation for calculating the distance between planes of atoms in a crystal lattice: $2d \sin \theta = n\lambda$.

- (41) Clausius - Clapeyron equation for determining ΔH_{vap} of a liquid:

$$\ln P = - \frac{\Delta H_{\text{vap}}}{RT} + C$$

- (42) Calculation of ΔH_{vap} , vapour pressure, or boiling point of a liquid:

$$\ln \frac{P_1}{P_2} = \frac{H_{\text{vap}}}{R} \left(\frac{1}{T_1} - \frac{1}{T_2} \right)$$

- (43) Entropy change of heat flow at constant temperature: $\Delta S = \frac{q}{t}$.

- (44) Calculating the molality of a solution : Molality (m) = $\frac{\text{moles of solute}}{1000 \text{ g solvent}}$.

- (45) Henry's law for calculating solubility of gases: $C = kP$

- (46) Raoult's law relating the vapour pressure of a liquid to its vapour pressure in a solution: $P_{\text{solution}} = x_{\text{solvent}} P_{\text{solvent}}^\circ$

- (47) Vapour pressure lowering in terms of the concentration of solution :

$$\Delta P = x_2 P_1^\circ$$

- (48) Boiling point elevation: $\Delta T_b = K_b m$.

- (49) Freezing point depression: $\Delta T_f = K_f m$.

- (50) Osmotic pressure of a solution: $\pi = i MRT$, where i = van't Hoff factor

- (51) The Van't Hoff factor for an electrolyte solution:

$$i = \frac{\text{Actual number of particles in solution after dissociation}}{\text{Number of formula units initially dissolved in solution}}$$

- (52) Rate law expression; the sum ($x + y$) gives the overall of the reaction:

$$\text{rate} = k [A]^x [B]^y$$

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- (53) Relationship between concentration and time for a first order reaction :

$$\ln \frac{[A]_0}{[A]} = kt$$

- (54) Equation for the graphical determination of k for a first - order reaction:

$$\ln [A] = - kt + \ln [A]_0$$

- (55) Half - life for a first - order reaction : $t_{1/2} = \frac{\ln 2}{k} = \frac{0.693}{k}$.

- (56) Relationship between concentration and time for a second-order reaction: $\frac{1}{[A]} = \frac{1}{[A]_0} + kt$.

- (57) The Arrhenius equation expressing the dependence of the rate constant on activation energy and temperature : $k = Ae^{-E_a/RT}$

- (58) Equation for the graphical determination of activation energy :

$$\ln k = \left(-\frac{E_a}{k} \right) \frac{1}{T} + \ln A$$

- (59) Relationship of rate constants at two different temperatures :

$$\ln \frac{k_1}{k_2} = \frac{E_a}{R} \left(\frac{1}{T_1} - \frac{1}{T_2} \right)$$

- (60) Law of Mass Action - General expression of equilibrium constant:

$$K = \frac{[C]^c [D]^d}{[A]^a [B]^b}$$

- (61) Relationship between K_p and K_c : $K_p = K_c (RT)^{\Delta n}$

where, Δn = (moles of gaseous product) – (moles of gaseous reactant)

- (62) The equilibrium constant for the overall reaction is given by the product of the equilibrium constants for the individual reactions :

$$K_c = K'_c K''_c$$

- (63) Ion-product constant of water : $K_w = [H^+][OH^-]$

- (64) Definition of pH of a solution : $pH = -\log [H^+]$

- (65) Definition of pOH of a solution : $pOH = -\log [OH^-]$.

- (66) Another form of ion-product constant of water: $pH + pOH = 14$

- (67) Percent ionization = $\frac{\text{Ionized acid concentration at equilibrium}}{\text{Initial concentration of acid}} \times 100\%$

- (68) Relationship between the acid and base ionization constants of a conjugate acid-base pair : $K_a K_b = K_w$

- (69) Henderson - Hasselbach equation: $pH = pK_a + \log \frac{[\text{conjugate base}]}{[\text{acid}]}$

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(70) The second law of thermodynamics (spontaneous process) :

$$\Delta S_{\text{univ}} = \Delta S_{\text{sys}} + \Delta S_{\text{surr}} > 0$$

(71) The second law of thermodynamics (equilibrium process):

$$\Delta S_{\text{univ}} = \Delta S_{\text{sys}} + \Delta S_{\text{surr}} = 0$$

(72) Standard entropy change of a reaction:

$\Delta S_{\text{rxn}}^{\circ} = \sum nS^{\circ}(\text{products}) - \sum mS^{\circ}(\text{reactants})$, where n and m are coefficients in the equation.

(73) Free-energy change at constant temperature : $\Delta G = \Delta H - T\Delta S$

(74) Standard free-energy change of a reaction:

$$\Delta G_{\text{rxn}}^{\circ} = \sum nG_f^{\circ}(\text{products}) - \sum mG_f^{\circ}(\text{reactants}),$$

where n and m are coefficients in the equation.

(75) Relationship between free-energy change and standard free-energy change and reaction quotient:

$$\Delta G = \Delta G^{\circ} + RT \ln Q.$$

(76) Relationship between standard free-energy change and the equilibrium constant: $\Delta G^{\circ} = RT \ln K.$

(77) Standard emf of an electrochemical cell :

$$E^{\circ}_{\text{cell}} = E^{\circ}_{\text{oxd}} - E^{\circ}_{\text{red}} = E^{\circ}_{\text{cathode}} - E^{\circ}_{\text{anode}}.$$

(78) Standard free energy change:

$$\Delta G^{\circ} = -nFE^{\circ}_{\text{cell}}, \text{ where } F \text{ is the Faraday constant.}$$

(79) Relationship of the standard emf of the cell to the equilibrium constant:

$$E^{\circ}_{\text{cell}} = \frac{RT}{nF} \ln K.$$

(80) The Nernst equation - for calculating the emf of a cell under non-standard conditions: $E = E^{\circ} - \frac{RT}{nF} \ln Q.$

(81) Relationship between mass defect and energy released:

$$\Delta E = (\Delta m) c^2.$$

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2.

Colours of Common Aqueous Ions

Ionic Species	Solution Concentration	
	1.0 mol /L	0.01 mol/L
Chromate	Yellow	Pale yellow
Chromium (III)	blue-green	green
Chromium (II)	dark blue	pale blue
Cobalt (II)	red	pink
Copper (I)	blue-green	pale blue green
Copper (II)	blue	pale blue
Dichromate	orange	pale orange
Iron (II)	lime green	colourless
Iron (III)	orange-yellow	pale yellow
Manganeses (II)	pale pink	colourless
Nickel (II)	blue-g	pale blue-green
Permanganate	deep purple	purple-pink

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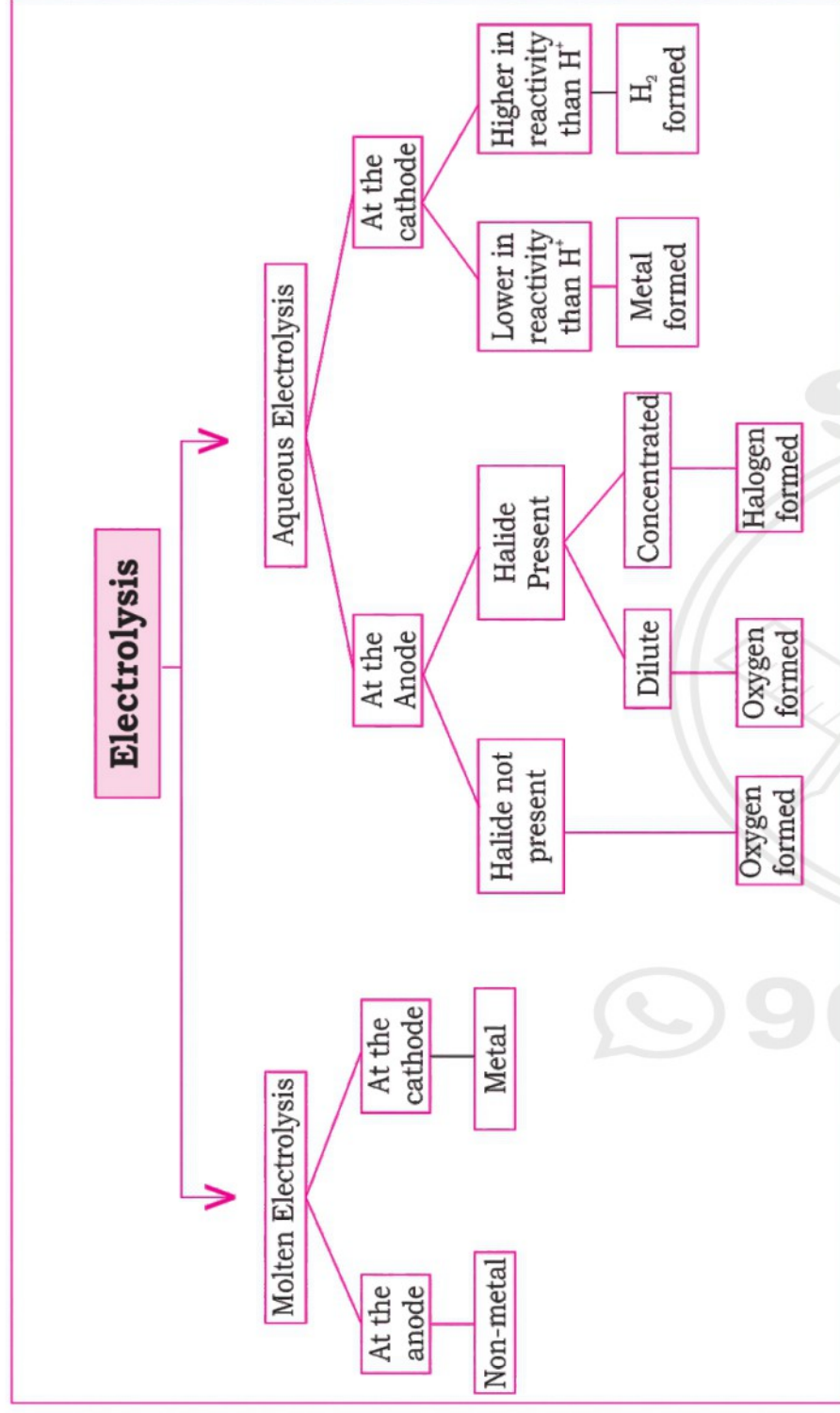
Acid-Base Indicators at 298.15K

Indicator	pH Range	Colour Change as pH Increases	K_a
Methyl violet	0.0 - 1.6	yellow to blue	$\sim 2 \times 10^{-1}$
Cresol red	0.0-1.0	red to yellow	$\sim 3 \times 10^{-1}$
	7.0 - 8.8	yellow to red	3.5×10^{-9}
Thymol blue	1.2-2.8	red to yellow	2.2×10^{-2}
	8.0 - 9.6	yellow to blue	6.3×10^{10}
Orange IV	1.4-2.8	red to yellow	$\sim 1 \times 10^{-4}$
Methyl orange	2.2-4.4	red to yellow	3.5×10^{-4}
Bromocresol green	3.8-5.4	yellow to blue	1.3×10^{-5}
Methyl red	4.8 - 6.0	red to yellow	1.0×10^{-5}
Chlorophenol red	5.2 - 6.8	yellow to red	5.6×10^{-7}
Bromothymol blue	6.0 - 7.6	yellow to blue	5.0×10^{-8}
Phenol red	6.6 - 8.0	yellow to red	1.0×10^{-8}
Phenolphthalein	8.2 -10.0	colourless to pink	3.2×10^{-10}
Thymolphthalein	9.4 - 10.6	colourless to blue	1.0×10^{-10}
Alizarin yellow	10.1-12.0	yellow to red	6.9×10^{-12}
Indigo carmine	11.4 13.0	blue to yellow	$\sim 6 \times 10^{-12}$
1,3,5-trinitrobenzene	12.0-14.0	colourless to orange	$\sim 1 \times 10^{-13}$

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Flowchart for Electrolysis



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5.

Relative Strength of Acid and Bases at 298.15K

Sr. No.	Common Name IUPAC / Systeatic Name	Acid Formula	Conjugate Base Formula	K_a
1.	Perchloric acid (aqueous hydrogen perchlorate)	$\text{HClO}_4(\text{aq})$	$\text{ClO}_4^-(\text{aq})$	Very large
2.	Hydroiodic acid (aqueous hydrogen iodide)	$\text{HI}(\text{aq})$	$\text{I}^-(\text{aq})$	very large
3.	Hydrobromic acid (aqueous hydrogen bromide)	$\text{HBr}(\text{aq})$	$\text{Br}^-(\text{aq})$	very large
4.	Hydrochloric acid (aqueous hydrogen chloride)	$\text{HCl}(\text{aq})$	$\text{Cl}^-(\text{aq})$	very large
5.	Sulfuric acid (aqueous hydrogen sulfate)	$\text{H}_2\text{SO}_4(\text{aq})$	$\text{HSO}_4^-(\text{aq})$	very large
6.	Nitric acid (aqueous hydrogen nitrate)	$\text{HNO}_3(\text{aq})$	$\text{NO}_3^-(\text{aq})$	very large
7.	Hydronium ion	$\text{H}_3\text{O}^+(\text{aq})$	$\text{H}_2\text{O}(\text{l})$	1
8.	Oxalic acid	$\text{HOOC-COOH}(\text{aq})$	$\text{HOOC-COO}^-(\text{aq})$	5.6×10^{-2}
9.	Sulfurous acid (aqueous hydrogen sulfite)	$\text{H}_2\text{SO}_3(\text{aq})$	$\text{HSO}_3^-(\text{aq})$	1.4×10^{-2}
10.	Hydrogen sulfate ion	$\text{HSO}_4^-(\text{aq})$	$\text{SO}_4^{2-}(\text{aq})$	1.0×10^{-2}
11.	Phosphoric acid (aqueous hydrogen phosphate)	$\text{H}_3\text{PO}_4(\text{aq})$	$\text{H}_2\text{PO}_4^-(\text{aq})$	6.9×10^{-3}
12.	Citric acid (2-hydroxy -1,2,3-Propanetricarboxylic acid)	$\text{C}_5\text{H}_5\text{O}(\text{COOH})_3(\text{aq})$	$\text{C}_3\text{H}_5\text{O}(\text{COOH})_2\text{COO}^-(\text{aq})$	7.4×10^{-4}
13.	Hydrofluoric acid (aqueous hydrogen fluoride)	$\text{HF}(\text{aq})$	$\text{F}^-(\text{aq})$	6.3×10^{-4}

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14.	Nitrous acid (aqueous hydrogen nitrite)	HNO ₂ (aq)	NO ₂ ⁻ (aq)	5.6 × 10 ⁻⁴
15.	Formic acid (methanoic acid)	HCOOH(aq)	HCOO ⁻ (aq)	1.8 × 10 ⁻⁴
16.	Hydrogen oxalate ion	HOOCOO ⁻ (aq)	OOCOO ²⁻ (aq)	1.5 × 10 ⁻⁴
17.	Lactic acid (2-hydroxypropanoic acid)	C ₂ H ₅ OCOOH(aq)	C ₂ H ₅ OCOO ⁻ (aq)	1.4 × 10 ⁻⁴
18.	Ascorbic acid 2(1,2-dihydroxyethyl)-4,5-dihydroxy-furan-3-one	H ₂ C ₆ H ₆ O ₆ (aq)	HC ₆ H ₆ O ₆ ⁻ (aq)	9.1 × 10 ⁻⁵
19.	Benzoic acid (benzenecarboxylic acid)	C ₆ H ₅ COOH(aq)	C ₆ H ₅ COO ⁻ (aq)	6.3 × 10 ⁻⁵
20.	Acetic acid (ethanoic acid)	CH ₃ COOH(aq)	CH ₃ COO ⁻ (aq)	1.8 × 10 ⁻⁵
21.	Dihydrogen citrate ion	C ₃ H ₅ (COOH) ₂ COO ⁻ (aq)	C ₃ H ₅ OCOOH(COO) ₂ ²⁻ (aq)	1.7 × 10 ⁻⁵
22.	Butanoic acid	C ₃ H ₇ COOH(aq)	C ₃ H ₇ COO ⁻ (aq)	1.5 × 10 ⁻⁵
23.	Propanoic acid	C ₂ H ₅ COOH(aq)	C ₂ H ₅ COO ⁻ (aq)	1.3 × 10 ⁻⁵
24.	Carbonic acid (CO ₂ + H ₂ O) (aqueous hydrogen sulfide)	H ₂ CO ₃ (aq)	HCO ₃ ⁻ (aq)	4.5 × 10 ⁻⁷
25.	Hydrogen citrate ion	C ₃ H ₅ OCOOH(COO) ₂ ²⁻ (aq)	C ₃ H ₅ O(COO) ₃ ³⁻ (aq)	4.0 × 10 ⁻⁷
26.	Hydrosulfuric acid (aqueous hydrogen sulfide)	H ₂ S(aq)	HS ⁻ (aq)	8.9 × 10 ⁻⁸
27.	Hydrogen sulfite ion	HSO ₃ ⁻ (aq)	SO ₃ ²⁻ (aq)	6.3 × 10 ⁻⁸
28.	Dihydrogen phosphate ion	H ₂ PO ₄ ⁻ (aq)	HPO ₄ ²⁻ (aq)	6.2 × 10 ⁻⁸
29.	Hypochlorous acid (aqueous hydrogen hypochlorite)	HOCl(aq)	OCl ⁻ (aq)	4.0 × 10 ⁻⁸

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30.	Hydrocyanic acid (aqueous hydrogen cyanide)	HCN(aq)	CN ⁻ (aq)	6.2 × 10 ⁻¹⁰
31.	Ammonium ion	NH ₄ ⁺ (aq)	CO ₃ ²⁻ (aq)	5.6 × 10 ⁻¹⁰
32.	Hydrogen carbonate ion	HCO ₃ ⁻ (aq)	CO ₃ ²⁻ (aq)	4.7 × 10 ⁻¹¹
33.	Hydrogen ascorbate ion	HC ₆ H ₆ O ₆ ⁻ (aq)	C ₆ H ₆ O ₆ ²⁻ (aq)	2.0 × 10 ⁻¹²
34.	Hydrogen phosphate ion	HPO ₄ ²⁻ (aq)	PO ₄ ³⁻ (aq)	4.8 × 10 ⁻¹³
35.	Water	H ₂ O(l)	OH ⁻ (aq)	1.0 × 10 ⁻¹⁴

(Note: (An approximation may be used instead of the quadratic formula when the concentration of H₃O⁺ produced is less than 50% of the original acid concentration. (or the concentration of the acid is 1000 times greater than the K_a) An approximation can be used for weak bases.)

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6. Chemistry Notations

Symbol	Term	Unit (s)
c	Specific heat capacity	J/(g.°C) or J/(g.K)
E°	Standard electrical potential	V or J/C
E_k	Kinetic energy	kJ
E_p	Potential energy	kJ
ΔH	Enthalpy (heat)	kJ
$\Delta_f H^\circ$	Standard molar enthalpy of formation	kJ/mol
I	Current	A or C/s
K_c	Equilibrium constant	—
K_a	Acid ionization (dissociation) constant	—
K_b	Base ionization (dissociation) constant	—
M	Molar mass	g/mol
m	Mass	g
n	Amount of substance	mol
P	Pressure	kPa
Q	Charge	C
T	Temperature (absolute)	K
t	Temperature (celsius)	°C
t	Time	s
V	Volume	L
C	Amount concentration	mol/L

Symbol	Term
Δ	delta (change in)
$^\circ$	Standard
[]	amount concentration

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7. Miscellaneous :

25.00 °C is equivalent to 298.15 K

Specific Heat capacities at 298.15 K and 100.000 kPa

$$\begin{aligned}
 C_{\text{air}} &= 1.01 \text{ J/(g.}^\circ\text{C)} \\
 C_{\text{polystyrene foam cup}} &= 1.01 \text{ J/(g.}^\circ\text{C)} \\
 C_{\text{copper}} &= 0.385 \text{ J/(g.}^\circ\text{C)} \\
 C_{\text{aluminum}} &= 0.897 \text{ J/(g.}^\circ\text{C)} \\
 C_{\text{iron}} &= 0.449 \text{ J/(g.}^\circ\text{C)} \\
 C_{\text{tin}} &= 0.227 \text{ J/(g.}^\circ\text{C)} \\
 C_{\text{water}} &= 4.19 \text{ J/(g.}^\circ\text{C)}
 \end{aligned}$$

Water Autoionization Constant (Dissociation Constant)

$$K_w = 1.0 \times 10^{-14} \text{ at } 298.15 \text{ K (for ion concentrations in mol/L)}$$

Faraday Constant

$$F = 9.665 \times 10^4 \text{ C/mol } e^-$$

Quadratic Formula

$$x = \frac{-b \pm \sqrt{b^2 - 4ac}}{2a}$$

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8. Colours of Common Aqueous Ions :

Ionic species	Solution Concentration	
	1.0 mol/L	0.010 mol/L
Chromate	yellow	pale yellow
Chromium (III)	blue-green	green
Chromium (II)	dark blue	pale blue
Cobalt (II)	red	pink
Copper (I)	blue-green	pale blue-green
Copper (II)	blue	pale blue
Dichromate	orange	pale orange
Iron (II)	lime green	colourless
Iron (III)	orange-yellow	pale yellow
Manganese (II)	pale pink	colourless
Nickel (II)	blue-green	pale blue-green
Permanganate	deep purple	purple-pink

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