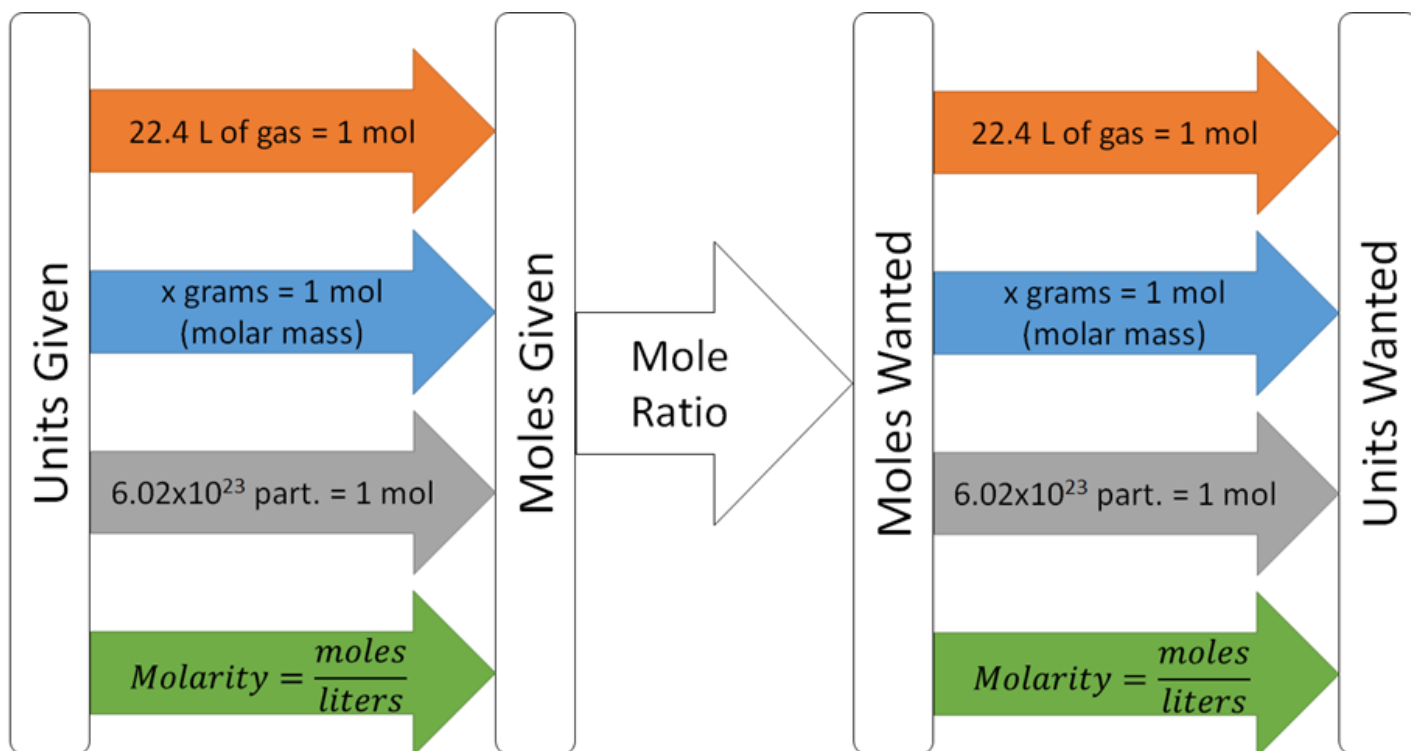


AP CHEMISTRY MASTER EQUATION SHEET

Reaction Stoichiometry

$M = \frac{\text{moles solvent}}{L \text{ of solution}}$	0.10 Molarity, 0.10 M, or [0.10] all mean the same thing.
$M_1V_1 = M_2V_2$	This can be used as a shortcut during titrations when the acid to base mole ratio is 1:1. If the mole ratio is not 1:1, you must take this into account during your calculations.

Stoichiometry Conversions



Matter and Intermolecular Forces

$PV = nRT$	Ideal gas law. Gases behave most ideally at low pressure and high temperature.
$\left(P + \frac{n^2a}{V^2}\right)(V - nb) = nRT$	Non-ideal gas law, accounts for volume taken up by gas particles and changes in pressure due to attraction between gas particles. Check the references section for a and b values.
$R = 8.31 \frac{L \cdot atm}{mol \cdot K}$ $R = 0.0821 \frac{L \cdot atm}{mol \cdot K}$ $R = 62.364 \frac{L \cdot atm}{mol \cdot K}$	Gas law constants. Make sure units match when you choose which one to use.
$1 \text{ atm} = 760 \text{ mmHg (or torr)} = 101.3 \text{ kPa}$	Some useful unit conversions.
$P_{total} = P_{gas A} + P_{gas B} + P_{gas C}$	Partial pressures
$P_{gas A} = P_{total} \times \frac{\text{moles gas A}}{\text{total moles of gas}}$	The partial pressure of a gas is directly proportional to the mole fraction of the gas in the mixture.
$K = ^\circ C + 273$	Conversion from Celsius to Kelvin. Most gas law equations require the Kelvin temperature.
$(KE)_{avg} = \frac{3}{2}RT$	The average kinetic energy of a collection of gas particles is related to its temperature.
$KE_{per \ molecule} = \frac{1}{2}mv^2$	The kinetic energy of a single gas molecule is related to its temperature and mass.
$KE_{gas \ particle} = \frac{1}{2}mu_{rms}^2$	This is a restatement of the previous equation. The u_{rms}^2 represent the average molecular speed.
$u_{rms} = \sqrt{\frac{3RT}{M}}$	The average molecular speed of a gas particle depends on the temperature and mass.

$\frac{\text{Effusion rate for gas 1}}{\text{Effusion rate for gas 2}} = \frac{\sqrt{M_2}}{\sqrt{M_1}}$	<p>The relative effusion rate of a gas is a ratio based on molar masses.</p>
$R_f = \frac{\text{distance travelled by molecule}}{\text{distance travelled by solvent front}}$	<p>The retention factor found during chromatography can be used to identify/ separate a mixture of compounds.</p>
$\%T = T \times 100 = \left(\frac{P}{P_0}\right) \times 100$	<p>Percent transmittance is a ratio of the light that can pass through a sample to the total amount of light.</p>
$A = -\log\left(\frac{P}{P_0}\right) = -\log(T)$	<p>Absorbance is the negative log of transmittance.</p>
$A = \epsilon bc$	<p>Beer's Law. A is absorbance, b is the path length (width of the cuvette, usually 1), and c is the concentration. The molar absorptivity coefficient (ϵ) is found from the slope of the graph of absorbance versus known concentrations.</p>

Chemical Kinetics

$\text{Rate} = \frac{\Delta[A]}{\Delta t}$	<p>Rate is the change in concentration of a species over time.</p>
$\begin{aligned} \text{instan. rate} &= \text{slope of tangent line} \\ &= \frac{\text{rise}}{\text{run}} \end{aligned}$	<p>In a graph of concentration of a species over time, the instantaneous rate is found using the slope of a line tangent to that point.</p>
$k = Ae^{-\frac{E_a}{RT}}$	<p>Arrhenius equation, used to relate the rate law constant to the activation energy.</p>
$\begin{aligned} \text{Rate} &= k[A] \\ \ln[A] &= -kt + \ln[A]_0 \\ t_{1/2} &= \frac{0.693}{k} \end{aligned}$	<p>First order rate laws Linear slope of $\ln[A]$ vs. t gives k</p>
$\begin{aligned} \text{Rate} &= k[A]^2 \\ \frac{1}{[A]} &= kt + \frac{1}{[A]_0} \\ t_{1/2} &= \frac{1}{k[A]_0} \end{aligned}$	<p>Second order rate laws Linear slope of $1/[A]$ vs. t gives k</p>
$\begin{aligned} \text{rate} &= k \\ [A] &= -kt + [A]_0 \\ t_{1/2} &= \frac{[A]_0}{2k} \end{aligned}$	<p>Zeroth order rate laws Linear slope of $[A]$ vs. t gives k</p>
<p>Rate Laws with Many Reactants</p> $\text{Rate} = k \cdot [A]^m \cdot [B]^n \cdot [C]^p$ $\text{Rate} = k' \cdot [A]^m \text{ where } k' = k \cdot [B]^n \cdot [C]^p$	<p>To determine a pseudo rate constant, conduct experiments where each reactant is limited by a large excess of the others, then monitor concentration. Create your rate law plots. The slope of the linear relationship indicates the order with respect to the limiting reactant and the pseudo rate constant. To get the entire raw law, you would repeat this until you knew the order with respect to each reactant.</p>

Equilibrium

$K_c = \frac{[C]^c [D]^d}{[A]^a [B]^b}$ $K_p = \frac{(P_{gas C})^c (P_{gas D})^d}{(P_{gas A})^a (P_{gas B})^b}$ $Q = \frac{[C]^c [D]^d}{[A]^a [B]^b}$	<p>For a reaction: $aA + bB \rightleftharpoons cC + dD$</p> <p>When $Q = K$, reaction is at equilibrium.</p>
$K_{sp} = [A]^a [B]^b$	<p>For the dissolving of an ionic compound: $A_a B_b(s) \rightleftharpoons aA^+(aq) + bB^-(aq)$</p>
$K_a = \frac{[H_3O^+][A^-]}{[HA]}$ $K_b = \frac{[BH^+][OH^-]}{[B]}$	<p>Where: $HA(aq) + H_2O(l) \rightleftharpoons H_3O^+(aq) + A^-(aq)$ $B(aq) + H_2O \rightleftharpoons BH^+(aq) + OH^-(aq)$</p>
$K_w = 1 \times 10^{-14} = K_a \times K_b$	<p>You can use K_w (constant) and K_a or K_b to find the other.</p>
$pK_w = pK_a + pK_b$	<p>Leads to $14 = pH + pOH$</p>
$pK_a = -\log(K_a)$ $pK_b = -\log(K_b)$	<p>The symbol p usually means "log of."</p>
$pH = -\log[H^+]$ $pOH = -\log[OH^-]$	<p>Use these to find pH and pOH. Remember that for strong acids and bases, $[H^+]$ or $[OH^-]$ is the same as the concentration of the acid or base. For weak acids and bases, you must use a RICE chart to find the concentration of $[H^+]$ or $[OH^-]$ at equilibrium.</p>
$pH = pK_a + \log\left(\frac{[base]}{[acid]}\right)$	<p>In buffer systems (such as those that occur when a weak acid/base is titrated by a strong one) at the $\frac{1}{2}$ equivalence point, $pH = pK_a$.</p>

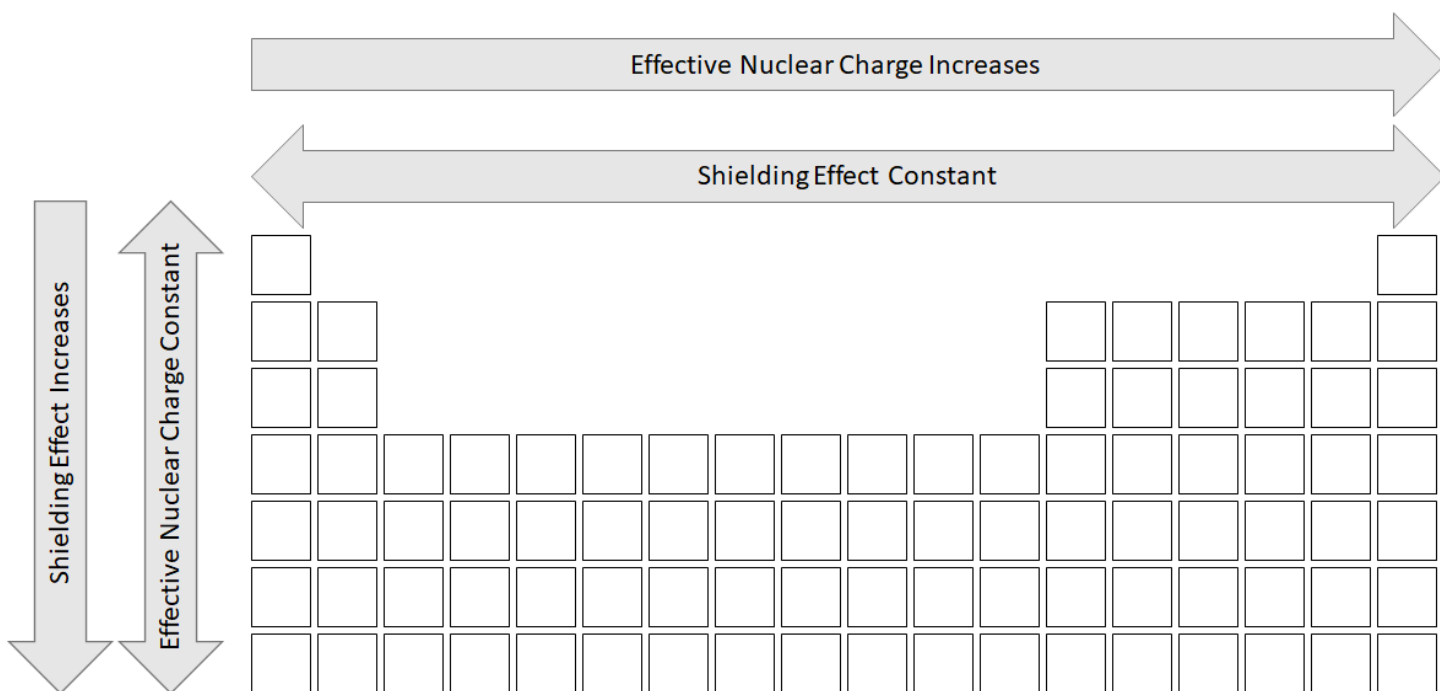
Thermochemistry and Thermodynamics, Electrochemistry

$q = mc\Delta T, \text{ for water, } c = 4.18 \text{ J/g}\cdot^{\circ}\text{C}$	<p>If the system loses heat, this will have a negative value. If it gains heat, this will be positive. Think about the direction of the heat flow.</p>
$w = -P\Delta V$	<p>Work is the energy needed to move something against a force, for example lifting books from the ground to a shelf does work against gravity. The energy of a system changes due to work and energy transfer (such as heat). We focus on the work done by expanding (negative) or compression (positive) of gases.</p>
$KE = \frac{1}{2}mv^2$	<p>The kinetic energy of a particle is related to its mass and velocity. If a collection of small and larger particles are at the same temperature, they have the same average KE. Particles may move faster or slower (velocity) depending on mass.</p>
$PE = \frac{kq_1q_2}{d}$	<p>The potential energy between two charges is related to the size of the charge and the distance between them. An important idea for bonding.</p>
$\Delta H_{rxn}^{\circ} = \sum n_p \Delta H_f^{\circ}(\text{products}) - \sum n_r \Delta H_f^{\circ}(\text{reactants})$	<p>A “pen and paper” way to calculate the change in enthalpy using a standard enthalpy of formation values from Thermodynamic Data.</p>
$\Delta H_{rxn}^{\circ} = (H_{\text{bonds broken}}) - (H_{\text{bonds formed}})$	<p>Another “pen and paper” way to calculate the change in enthalpy, using the Average Bond Enthalpies table. Bond breaking is endothermic, bond formation is exothermic, apply the correct sign.</p>
$\Delta S_{\text{universe}} = \Delta S_{\text{system}} + \Delta S_{\text{surroundings}}$	<p>The total disorder of the universe is the sum of the disorder of the system and surroundings. Disorder drives many chemical processes and can have positive or negative signs.</p>
$\Delta S_{rxn}^{\circ} = \sum n_p S^{\circ}(\text{products}) - \sum n_r S^{\circ}(\text{reactants})$	<p>A “pen and paper” way to calculate the change in disorder using absolute</p>

	disorder values from Thermodynamic Data .
$\Delta G = \Delta H - T\Delta S$	The change in Gibbs free energy can be found from the change in enthalpy and entropy. A negative value for a process indicates a release of energy, and is considered to be spontaneous (doesn't require any energy input to occur).
$\Delta G_{rxn}^{\circ} = \sum n_p \Delta G_f^{\circ} (products) - \sum n_r \Delta G_f^{\circ} (reactants)$	A "pen and paper" way to calculate the change in Gibbs free energy using standard free energy of formation values from Thermodynamic Data .
$\Delta G = \Delta G^{\circ} + RT \ln(Q)$	Used when not at equilibrium or not at standard conditions.
$\Delta G^{\circ} = -RT \ln(K)$	Used when at equilibrium ($\Delta G = 0$).
$\Delta G^{\circ} = -nFE^{\circ}$	The maximum cell potential (E°) is related to the free energy difference between reactants and products in an electrochemical cell. When cell potential is positive, ΔG is negative and spontaneous (galvanic cell). If cell potential is negative, ΔG is positive and non-spontaneous (electrolytic cell, energy must be added for electrons to flow).
$I = \frac{q}{t}$	Current is charge per unit time, typically coulombs/ second.

Periodicity and Bonding

<p><i>formal charge</i> $= (\# ve^-) - (\text{assigned } e^-)$</p>	<p>Nonbonding electrons are assigned to their atom. Bonding electrons are divided equally between their atoms.</p>
<p>$E = h\nu$</p>	<p>Used to calculate the energy of a photon. Energy is equal to Plank's constant (6.626×10^{-34} J·s) times velocity.</p>
<p>$c = \lambda\nu$</p>	<p>The speed of light (3.00×10^8 m·s) is equal to the frequency of the photon times the velocity of the photon.</p>
<p>Ionization Energy</p>	<p>Energy needed to ionize an atom by removing an electron.</p>
<p>Electronegativity</p>	<p>How well an atom attracts electrons in a shared bond. "Sexiness"</p>



Reference Materials

Polyatomic Ions

Ion	Name	Ion	Name
H_3O^+	hydronium	CrO_4^{2-}	chromate
Hg_2^{2+}	dimercury (I)	$\text{Cr}_2\text{O}_7^{2-}$	dichromate
NH_4^+	ammonium	MnO_4^-	permanganate
$\text{C}_2\text{H}_3\text{O}_2^-$	acetate	NO_2^-	nitrite
CH_3COO^-		NO_3^-	nitrate
$\text{C}_2\text{O}_4^{2-}$	oxalate	O_2^{2-}	peroxide
CO_3^{2-}	carbonate	OH^-	hydroxide
HCO_3^-	hydrogen (bi)carbonate	CN^-	cyanide
PO_4^{3-}	phosphate	SCN^-	thiocyanate
ClO^-	hypochlorite	SO_3^{2-}	sulfite
ClO_2^-	chlorite	SO_4^{2-}	sulfate
ClO_3^-	chlorate	HSO_4^-	hydrogen sulfate
ClO_4^-	perchlorate	$\text{S}_2\text{O}_3^{2-}$	thiosulfate

Simple Solubility Rules

- Most nitrate (NO_3^-) salts are soluble.
- Most salts containing the alkali metal (group 1) ions (Li^+ , Na^+ , K^+ , Cs^+ , Rb^+) and the ammonium (NH_4^+) ion are soluble.
- Most chloride, bromide and iodide salts are soluble. Exceptions are salts containing the ions Ag^+ , Pb^{2+} , and Hg^{2+} .
- Most sulfate salts are soluble. Notable exceptions are BaSO_4 , PbSO_4 , HgSO_4 and CaSO_4 .
- Most hydroxide salts are only slightly soluble. The important soluble hydroxides are NaOH and KOH . The hydroxides of barium, strontium and calcium are marginally soluble.
- Most sulfide (S^{2-}), carbonate (CO_3^{2-}), chromate (CrO_4^{2-}), and phosphate (PO_4^{3-}) salts are only slightly soluble.

Solubility Product Constants (K_{sp}) at 25°C

Halides	K_{sp}	Sulfates	K_{sp}	Hydroxides	K_{sp}
AgCl	1.8×10^{-10}	PbSO_4	6.3×10^{-7}	$\text{Al}(\text{OH})_3$	3.0×10^{-34}
AgBr	5.0×10^{-13}	BaSO_4	1.1×10^{-10}	$\text{Zn}(\text{OH})_2$	3.0×10^{-16}
AgI	8.3×10^{-17}	CaSO_4	2.4×10^{-5}	$\text{Ca}(\text{OH})_2$	6.5×10^{-6}
PbCl_2	1.7×10^{-5}	Sulfides	K_{sp}	$\text{Mg}(\text{OH})_2$	7.1×10^{-12}
PbBr_2	2.1×10^{-6}	NiS	4.0×10^{-20}	$\text{Fe}(\text{OH})_2$	7.9×10^{-16}
PbI_2	7.9×10^{-9}	CuS	8.0×10^{-37}	Carbonates	K_{sp}
PbF_2	3.6×10^{-8}	Ag_2S	8.0×10^{-51}	CaCO_3	4.5×10^{-9}
CaF_2	3.9×10^{-11}	ZnS	3.0×10^{-23}	SrCO_3	9.3×10^{-10}
Chromates	K_{sp}	FeS	8.0×10^{-19}	ZnCO_3	1.0×10^{-10}
PbCrO_4	1.8×10^{-14}	CdS	1.0×10^{-27}	Ag_2CO_3	8.15×10^{-12}

Ag_2CrO_4	1.2×10^{-12}	PbS	3.0×10^{-28}	BaCO_3	5.0×10^{-9}
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Solubility Guidelines for Aqueous Solutions

Ions that form Soluble Compounds	Exceptions	Ions that form Insoluble Compounds	Exceptions
Group 1 ions (Li^+ , Na^+ , etc.)		Carbonate (CO_3^{2-})	When combined with group 1 ions or ammonium
Ammonium (NH_4^+)		Chromate (CrO_4^{2-})	When combined with group 1 ions, Ca^{2+} , Mg^{2+} , or ammonium
Nitrate (NO_3^-)		Phosphate (PO_4^{3-})	When combined with group 1 ions or ammonium
Acetate ($\text{C}_2\text{H}_3\text{O}_2^-$ or CH_3COO^-)		Sulfide (S^{2-})	When combined with group 1 ions or ammonium
Hydrogen carbonate (HCO_3^-)		Hydroxide (OH^-)	When combined with group 1 ions, Ca^{2+} , Ba^{2+} , Sr^{2+} , or ammonium
Chlorate (ClO_3^-)			
Perchlorate (ClO_4^-)			
Halides (F^- , Cl^- , etc.)	When combined with Ag^+ , Pb^{2+} , or Hg_2^{2+}		
Sulfates (SO_4^{2-})	When combined with Ag^+ , Ca^{2+} , Sr^{2+} , Ba^{2+} , and Pb^{2+}		

Strong Acids and Bases

Acids: HCl, HBr, HI, HNO₃, H₂SO₄ (first H⁺ only), HClO₃, HClO₄

Bases: Group 1 hydroxides (NaOH, KOH, etc.), Ca(OH)₂^{*}, Ba(OH)₂^{*}, Sr(OH)₂^{*}

(*considered strong at concentrations below 0.10 M)

K_a and K_b Values for Some Common Weak Acids and Weak Bases

Acid	Formula	K _a	Base	Formula	K _b
Hydrogen sulfate	HSO ₄ ⁻	1.2x10 ⁻²	Ammonia	NH ₃	1.8x10 ⁻⁵
Chlorous acid	HClO ₂	1.2x10 ⁻²	Methylamine	CH ₃ NH ₂	4.4x10 ⁻⁴
Hydrofluoric acid	HF	7.2x10 ⁻⁴	Ethylamine	C ₂ H ₅ NH ₂	5.6x10 ⁻⁴
Nitrous acid	HNO ₂	4.0x10 ⁻⁴	Hydroxylamine	OHNH ₂	1.1x10 ⁻⁸
Lactic acid	HC ₃ H ₅ O ₃	1.4x10 ⁻⁴	Hydrazine	H ₂ NNH ₂	3.0x10 ⁻⁶
Acetic acid	HC ₂ H ₃ O ₂	1.8x10 ⁻⁵	Aniline	C ₆ H ₅ NH ₂	3.8x10 ⁻¹⁰
Propanoic acid	HC ₃ H ₅ O ₂	1.3x10 ⁻⁵	Pyridine	C ₅ H ₅ N	1.7x10 ⁻⁹
Hypochlorous acid	HOCl	3.5x10 ⁻⁸			
Hypobromous acid	HOBr	2x10 ⁻⁹			
Hypoiodous acid	HOI	2x10 ⁻¹¹			
Hydrocyanic acid	HCN	6.2x10 ⁻¹⁰			

Van der Waals Constants

Gas	a	b
He	0.0341	0.0237
Ne	0.211	0.0171
Ar	1.35	0.0322
Kr	2.32	0.0398
Xe	4.19	0.0511
H ₂	0.244	0.0266
N ₂	1.39	0.0391
O ₂	1.36	0.0318
Cl ₂	6.49	0.0562
CO ₂	3.59	0.0427
CH ₄	2.25	0.0428
NH ₃	4.17	0.0371
H ₂ O	5.46	0.0305

Partial Pressure of Water Vapor for Select Temperatures

Temperature (°C)	Pressure (mmHg)	Temperature (°C)	Pressure (mmHg)
15	12.8	26	25.2
16	13.6	27	26.7
17	14.5	28	28.3
18	15.5	29	30.0
19	16.5	30	31.8
20	17.5	31	33.7
21	18.7	32	35.7
22	19.8	33	37.7
23	21.1	34	39.9
24	22.4	35	42.2
25	23.8	36	44.6

Rules to Determine Oxidation States

The oxidation state of ...	Summary	Examples
An atom in elemental form is zero	Element = 0	Na(s), O ₂ (g), O ₃ (g), Hg(l)
A monatomic ion is the same as its charge	Monatomic ion = charge of ion	Na ⁺ , Cl ⁻
Fluorine is -1 in its compounds	Fluorine = -1	HF, PF ₃
Oxygen is -2 in its compounds (except peroxides, O ₂ ²⁻ , in which oxygen is -1)	Oxygen = -2	H ₂ O, CO ₂
Hydrogen is +1 in covalent compounds	Hydrogen = +1	H ₂ O, HCl, NH ₃

Remember that the sum of oxidation states must equal the charge.

Standard Reduction Potentials

Half-Reaction	\mathcal{E}° (V)	Half-Reaction	\mathcal{E}° (V)
$F_2 + 2e^- \rightarrow 2F^-$	2.87	$O_2 + 2H_2O + 4e^- \rightarrow 4OH^-$	0.40
$Ag^{2+} + e^- \rightarrow Ag^+$	1.99	$Cu^{2+} + 2e^- \rightarrow Cu$	0.34
$Co^{3+} + e^- \rightarrow Co^{2+}$	1.82	$Hg_2Cl_2 + 2e^- \rightarrow 2Hg + 2Cl^-$	0.27
$H_2O_2 + 2H^+ + 2e^- \rightarrow 2H_2O$	1.78	$AgCl + e^- \rightarrow Ag + Cl^-$	0.22
$Ce^{4+} + e^- \rightarrow Ce^{3+}$	1.70	$SO_4^{2-} + 4H^+ + 2e^- \rightarrow H_2SO_3 + H_2O$	0.20
$PbO_2 + 4H^+ + SO_4^{2-} + 2e^- \rightarrow PbSO_4 + 2H_2O$	1.69	$Cu^{2+} + e^- \rightarrow Cu^+$	0.16
$MnO_4^- + 4H^+ + 3e^- \rightarrow MnO_2 + 2H_2O$	1.68	$2H^+ + 2e^- \rightarrow H_2$	0.00
$2e^- + 2H^+ + IO_4^- \rightarrow IO_3^- + H_2O$	1.60	$Fe^{3+} + 3e^- \rightarrow Fe$	-0.036
$MnO_4^- + 8H^+ + 5e^- \rightarrow Mn^{2+} + 4H_2O$	1.51	$Pb^{2+} + 2e^- \rightarrow Pb$	-0.13
$Au^{3+} + 3e^- \rightarrow Au$	1.50	$Sn^{2+} + 2e^- \rightarrow Sn$	-0.14
$PbO_2 + 4H^+ + 2e^- \rightarrow Pb^{2+} + 2H_2O$	1.46	$Ni^{2+} + 2e^- \rightarrow Ni$	-0.23
$Cl_2 + 2e^- \rightarrow 2Cl^-$	1.36	$PbSO_4 + 2e^- \rightarrow Pb + SO_4^{2-}$	-0.35
$Cr_2O_7^{2-} + 14H^+ + 6e^- \rightarrow 2Cr^{3+} + 7H_2O$	1.33	$Cd^{2+} + 2e^- \rightarrow Cd$	-0.40
$O_2 + 4H^+ + 4e^- \rightarrow 2H_2O$	1.23	$Fe^{2+} + 2e^- \rightarrow Fe$	-0.44
$MnO_2 + 4H^+ + 2e^- \rightarrow Mn^{2+} + 2H_2O$	1.21	$Cr^{3+} + e^- \rightarrow Cr^{2+}$	-0.50
$IO_3^- + 6H^+ + 5e^- \rightarrow \frac{1}{2}I_2 + 3H_2O$	1.20	$Cr^{3+} + 3e^- \rightarrow Cr$	-0.73
$Br_2 + 2e^- \rightarrow 2Br^-$	1.09	$Zn^{2+} + 2e^- \rightarrow Zn$	-0.76
$VO_2^+ + 2H^+ + e^- \rightarrow VO^{2+} + H_2O$	1.00	$2H_2O + 2e^- \rightarrow H_2 + 2OH^-$	-0.83
$AuCl_4^- + 3e^- \rightarrow Au + 4Cl^-$	0.99	$Mn^{2+} + 2e^- \rightarrow Mn$	-1.18
$NO_3^- + 4H^+ + 3e^- \rightarrow NO + 2H_2O$	0.96	$Al^{3+} + 3e^- \rightarrow Al$	-1.66
$ClO_2 + e^- \rightarrow ClO_2^-$	0.954	$H_2 + 2e^- \rightarrow 2H^-$	-2.23
$2Hg^{2+} + 2e^- \rightarrow Hg_2^{2+}$	0.91	$Mg^{2+} + 2e^- \rightarrow Mg$	-2.37
$Ag^+ + e^- \rightarrow Ag$	0.80	$La^{3+} + 3e^- \rightarrow La$	-2.37
$Hg_2^{2+} + 2e^- \rightarrow 2Hg$	0.80	$Na^+ + e^- \rightarrow Na$	-2.71
$Fe^{3+} + e^- \rightarrow Fe^{2+}$	0.77	$Ca^{2+} + 2e^- \rightarrow Ca$	-2.76
$O_2 + 2H^+ + 2e^- \rightarrow H_2O_2$	0.68	$Ba^{2+} + 2e^- \rightarrow Ba$	-2.90
$MnO_4^- + e^- \rightarrow MnO_4^{2-}$	0.56	$K^+ + e^- \rightarrow K$	-2.92
$I_2 + 2e^- \rightarrow 2I^-$	0.54	$Li^+ + e^- \rightarrow Li$	-3.05
$Cu^+ + e^- \rightarrow Cu$	0.52		

Calculating Voltage from Standard Reduction Potential

Locate the two relevant half reactions. The one with the most positive reduction potential will run as written as a reduction. The other half reaction is “flipped” (sign and direction) and runs as oxidation. Multiply each half reaction by a whole number to ensure the number of electrons produced and consumed on both sides of the arrow are equal. Do NOT multiply the voltages by these factors, just sum them together.

Average Bond Enthalpies

Single Bonds (kJ/mol)							
C–H	413	N–H	391	O–H	463	F–F	155
C–C	348	N–N	163	O–O	146		
C–N	293	N–O	201	O–F	190	Cl–F	253
C–O	358	N–F	272	O–Cl	203	Cl–Cl	242
C–F	485	N–Cl	200	O–I	234		
C–Cl	328	N–Br	243			Br–F	237
C–Br	276					Br–Cl	218
C–I	240	H–H	436	S–H	339	Br–Br	193
C–S	259	H–F	567	S–F	327		
		H–Cl	431	S–Cl	253	I–Cl	208
Si–H	323	H–Br	366	S–Br	218	I–Br	175
Si–Si	226	H–I	299	S–S	266	I–I	151
Si–C	301			S–N	464		
Si–O	368						
Multiple Bonds (kJ/mol)							
C=C	614	N=N	418	O=O	495		
C≡C	839	N≡N	941				
C=N	615	N=O	607	S=O	523		
C≡N	891			S=S	418		
C=O	799						

Table of Thermodynamic Data

[Click here.](#)

Specific Heats

Substance	J/(g·°C)	cal/(g·°C)
Water (liquid)	4.18	1.00
Water (gas)	1.9	0.45
Water (solid)	2.1	0.50
Ethanol	2.4	0.58
Chloroform	0.96	0.23
Aluminum	0.90	0.21
Iron	0.46	0.11
Silver	0.24	0.057
Mercury	0.14	0.033

Heats of Physical Change

Substance	ΔH_{fus} (kJ/mol)	ΔH_{vap} (kJ/mol)
Ammonia (NH ₃)	5.66	23.3
Ethanol (C ₂ H ₆ O)	4.93	38.6
Hydrogen (H ₂)	0.12	0.90
Methanol (CH ₄ O)	3.22	35.2
Oxygen (O ₂)	0.44	6.82

Water (H ₂ O)	6.01	40.7
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