Harold’s High School Chemistry
Cheat Sheet

25 August 2025

Chapter 1: Measuring Up

<table>
<tr><th>Term</th><th>Description</th><th>Equation</th></tr>
<tr><td>Rulers</td><td colspan="2">When using a ruler that is marked off in 16ths of an inch, report your answers to a hundredth of an inch.</td></tr>
<tr><td>Units</td><td colspan="2">The unit of measurement is just as important as the number.
You must always list the units, followed by the compound!
Example: 6.28 mL H_2O</td></tr>
<tr><td>Significant Figures</td><td colspan="2">1. All non-zero figures (1, 2, 3, 4, 5, 6, 7, 8, and 9) are significant.
2. A zero (0) is significant if it is between two significant digits.
3. A zero (0) is also significant if it’s at the end of the number <i>and</i> to the right of the decimal point.</td></tr>
<tr><td>Using SigFigs</td><td colspan="2">1. When adding and subtracting measurements, you must report your answer to the same precision as the least precise number in the problem.
2. When multiplying and dividing measurements, you must report your answer with the same number of significant figures as the measurement that has the fewest significant figures.
3. There is always some error in the last significant figure of a measurement.</td></tr>
<tr><td>Precision vs. Accuracy</td><td colspan="2">- Precision: The consistency and reproducibility of measurements (e.g., 10 decimal places).
- Accuracy: How close a measurement is to the true or accepted value.</td></tr>
<tr><td>Prefixes</td><td colspan="2">
<table>
<tr><th>Prefix</th><th>Abbreviation</th><th>Meaning</th><th>Scientific</th></tr>
<tr><td>giga</td><td>G</td><td>1,000,000,000</td><td>10^9</td></tr>
<tr><td>mega</td><td>M</td><td>1,000,000</td><td>10^6</td></tr>
<tr><td>kilo</td><td>k</td><td>1,000</td><td>$\mathbf{10^3}$</td></tr>
<tr><td>hector</td><td>H</td><td>100</td><td>10^2</td></tr>
<tr><td>deca</td><td>Da</td><td>10</td><td>10^1</td></tr>
<tr><td>centi</td><td>c</td><td>0.01</td><td>$\mathbf{10^{-2}}$</td></tr>
<tr><td>milli</td><td>m</td><td>0.001</td><td>$\mathbf{10^{-3}}$</td></tr>
<tr><td>micro</td><td>µ</td><td>0.000001</td><td>10^{-6}</td></tr>
<tr><td>nano</td><td>n</td><td>0.000000001</td><td>10^{-9}</td></tr>
</table>
</td></tr>
<tr><td>Scientific Notation</td><td>$14{,}000{,}000 = 1.4 \times 10^7 = 1.4\text{E}7$</td><td>$0.00000014 = 1.4 \times 10^{-7} = 1.4\text{E}{-7}$</td></tr>
<tr><td>Measuring</td><td>Volume (L)</td><td>1 cm^3 = 1 mL</td></tr>
<tr><td></td><td>Mass (g)</td><td>weight = mass × gravity</td></tr>
<tr><td></td><td>Density (ρ)</td><td>$density = \frac{mass}{volume} \quad or\ \rho = \frac{m}{V}$</td></tr>
<tr><td>Unit Conversion
(Train Track Method)</td><td colspan="2">0.1436 mL = _?_ m^3
$$\begin{array}{c|c|c|c|c} 143.6\ \text{L} & 1000\ \cancel{\text{mL}} & 1\ \text{cm}^3 & (1\ \text{m})^3 & \mathbf{0.1436\ m^3} \\ \hline & 1\ \text{L} & 1\ \cancel{\text{mL}} & (100\ \text{cm})^3 & \end{array}$$</td></tr>
</table>

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Chapter 2: What’s The Matter

Term	Description	Equation

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Chapter 3: Making Sense of Atoms and Elements

Term	Description	Equation

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Chapter 4: The Modern View of Atoms and Their Chemistry

Term	Description	Equation

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Chapter 5: Covalent Compounds and Their Molecular Geometry

Term	Description	Equation

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Chapter 6: Physical and Chemical Changes

Term	Description	Equation

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Chapter 7: Stoichiometry

Term	Description	Equation

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Chapter 8: Still More on Stoichiometry

Common Polyatomic Ions (Memorize)

#	Name	Ion
1	Ammonium	NH_4^+
2	Hydronium	H_3O^+
3	Acetate	$C_2H_3O_2^-$
4	Cyanide	CN^-
5	Bicarbonate	HCO_3^-
6	Carbonate	CO_3^{2-}
7	Hydroxide	OH^-
8	Nitrite	NO_2^-
9	Nitrate	NO_3^-
10	Sulfite	SO_3^{2-}
11	Sulfate	SO_4^{2-}
12	Phosphate	PO_4^{3-}
13	Chlorite	ClO_2^-
14	Manganate	MnO_4^{2-}

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Chapter 9: Chemists Have Solutions

<table>
<tr><th>Term</th><th>Description</th><th>Equation</th></tr>
<tr><td>Dissolving Compounds</td><td colspan="2">When ionic compounds dissolve, they split up into their individual atoms.
When polar covalent compounds dissolve, they split up into their individual molecules.</td></tr>
<tr><td>Solubility of Solutes</td><td colspan="2"><table><tr><th>State</th><th>Temperature</th><th>Pressure</th></tr><tr><td>Solid</td><td>↑</td><td>-</td></tr><tr><td>Liquid</td><td>-</td><td>-</td></tr><tr><td>Gas</td><td>↓</td><td>↑</td></tr></table></td></tr>
<tr><td>Concentration</td><td colspan="2">The behavior of a chemical often depends on concentration.</td></tr>
<tr><td>Molarity</td><td>$Molarity = \frac{moles\ of\ solute}{liters\ of\ solution}\ M$</td><td>Liters</td></tr>
<tr><td>Molality</td><td>$Molality = \frac{moles\ of\ solute}{kilograms\ of\ solvent}\ m$</td><td>Kilograms</td></tr>
<tr><td>Freezing Point Depression</td><td>$\Delta T = -i \cdot K_f \cdot m$</td><td>$K_{f\_H_2O} = 1.86\ ^{\circ}C/m$</td></tr>
<tr><td>Boiling Point Elevation</td><td>$\Delta T = i \cdot K_b \cdot m$</td><td>$K_{b\_H_2O} = 0.512\ ^{\circ}C/m$</td></tr>
</table>

Chapter 10: It's a Gas!

Term	Equation	Note
Ideal Gas Law	$\boldsymbol{PV = nRT}$	$P = Pressure\ (atm)$ $V = Volume\ (L)$ $n = \#\ Moles\ of\ gas$ Ideal Gas Constant (R): $R = 0.0821 \frac{L\ atm}{mole\ K}$ $T = Temperature\ (K)$
Boyle's Law	$P_1V_1 = P_2V_2$	Assumes $T = constant$
Charle's Law	$\frac{V_1}{T_1} = \frac{V_2}{T_2}$	Assumes $P = constant$
Combined Gas Law	$\frac{P_1\,V_1}{T_1} = \frac{P_2\,V_2}{T_2}$	Since $nR = constant$
Avagadro's Law	$n = \frac{PV}{RT}$	Same number of molecules or atoms
Volumes of Gases	$V = \frac{nRT}{P}$	Relationship between volumes of those gases
Kelvin (T)	$K = °C + 273.15$ $0\ K = Absolute\ Zero$	$°F = \left(\frac{9}{5}\right)°C + 32 \qquad [32° - 212°]$ $°C = \left(\frac{5}{9}\right)(°F - 32) \qquad [0° - 100°]$
Pressure (P)	$P = \frac{F}{A}$	$Pressure = \frac{Force}{Area}$
	$1\ atm = 760\ torr$ $1\ torr = 1\ mmHg$ $1\ mmHg = 133.322\ Pa$ $1\ atm = 101{,}325\ Pascals\ (Pa)$	Units can be torr, atm, or Pa
Ideal Gas	a. The molecules or atoms that make up the gas occupy <u>no volume</u>. b. The molecules or atoms that make up the gas are <u>not attracted</u> to each other. c. The <u>collisions</u> that occur between the molecules or atoms that make up the gas are <u>elastic</u>, which means no energy is lost in such a collision. This is also true for any collisions between the molecules or atoms that make up the gas and the walls of the container in which the gas is held.	
STP	A gas is at STP if its pressure is 1 atm and its temperature is $0°C$.	**S**tandard **T**emperature and **P**ressure
Dalton's Law	$P_{total} = P_a + P_b + P_c + \cdots$	Partial pressures
Mole Fraction	$X_a = \frac{P_a}{P_{total}}$	$X_a = \frac{moles\ of\ compound\ a}{total\ moles\ in\ the\ mixture}$
Vapor Pressure	Boiling point = The temperature at which a liquid's vapor pressure is equal to the external air pressure.	
Extrapolation	Extending a trend in data to situations for which no measurements have been made. Usually linear approximations.	

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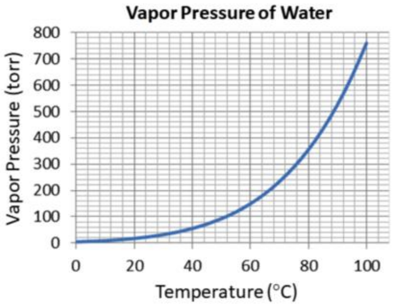
Vapor Pressure of Water

Temperature (°C)	Vapor Pressure (torr)		Temperature (°C)	Vapor Pressure (torr)
0	4.6		39	52.4
2	5.3		40	55.3
4	6.1		42	61.5
6	7.0		44	68.3
8	8.0		46	75.5
10	9.2		48	83.7
12	10.5		50	92.5
14	12.0		52	102.1
15	12.8		54	112.5
16	13.6		56	123.8
17	14.5		58	136.1
18	15.5		60	149.4
19	16.5		62	163.8
20	17.5		64	179.3
21	18.7		66	196.1
22	19.8		68	214.2
23	21.1		70	233.7
24	22.4		72	254.6
25	23.8		74	277.2
26	25.2		76	301.4
27	26.7		78	327.3
28	28.3		80	355.1
29	30.0		82	384.9
30	31.8		84	416.8
31	33.7		86	450.9
32	35.7		88	487.1
33	37.7		90	525.8
34	39.9		92	567.0
35	42.2		94	610.9
36	44.6		96	657.6
37	47.1		98	707.3
	49.7		100	760.0



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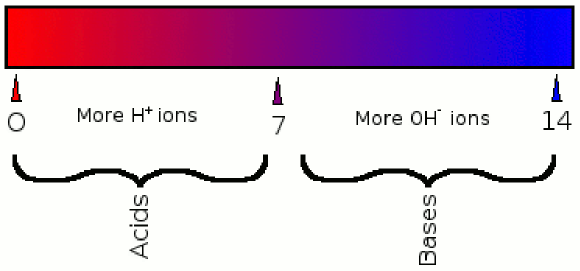
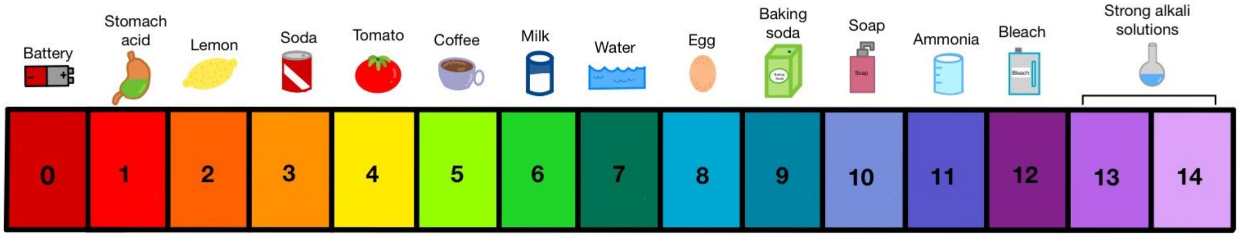
Chapter 11: Some Pretty Basic (and Acidic) Chemicals

Term	Equation / Definition	Note
Acids **(Acidic)**	A chemical that donates an H^+. Have more H^+ ions.	1. Tend to taste sour 2. Are covalent electrolytes 3. Turn blue litmus paper red
Bases **(Alkaline)**	A chemical that accepts an H^+. Have more OH^- ions.	1. Tend to taste bitter 2. Tend to feel slippery when mixed with water 3. Turn red litmus paper blue
Litmus	An acid/base indicator that is usually on a strip of paper	
Amphoteric (Amphiprotic)	Capable of reacting as either an acid or a base (H_2O, metal oxides)	
Covalent Electrolytes	hydrofluoric acid + water ⟶ fluoride ion + hydronium ion $HF + H_2O \longrightarrow F^- + H_3O^-$ acid base	
	ammonium hydroxide (ammonia) + water → ammonium ion + hydroxide ion $NH_3 + H_2O \longrightarrow NH_4^+ + OH^-$ base acid	
Ionic Electrolytes	hydrochloric acid + sodium hydroxide (lye) → water + salt $HCl\ (aq) + NaOH\ (aq) \longrightarrow H_2O(l) + NaCl\ (aq)$ acid base (Na^+OH^-) (Na^+Cl^-)	
Acid/Base Identification Rules	- Ammonia (NH_3) is a covalent base. - If a covalent compound starts with an H, it can usually act like an acid. - Ionic compounds that contain the hydroxide ion (OH^-) can act as bases. - An acid reacts with an ionic base to make water (H_2O) and salt ($NaCl$).	
pH Scale	Potential hydrogen (pH). Amount of hydronium ion (H_3O^-) in the solution. pH of 0: highly acidic pH of 7: neutral pH of 14: highly alkaline	
Polyprotic Acid	An acid that can donate two or more H^+ ions. Examples: hydrogen sulfate (H_2SO_4), carbonic acid (H_2CO_3) $CO_2\ (g) + H_2O\ (l) \longrightarrow H_2CO_3\ (aq)\quad acid\ rain$	
Titration	The process by which an acid of known concentration is added to a base of unknown concentration (or vice versa) until a neutral pH is reached to determine the unknown concentration.	



pH of Acids

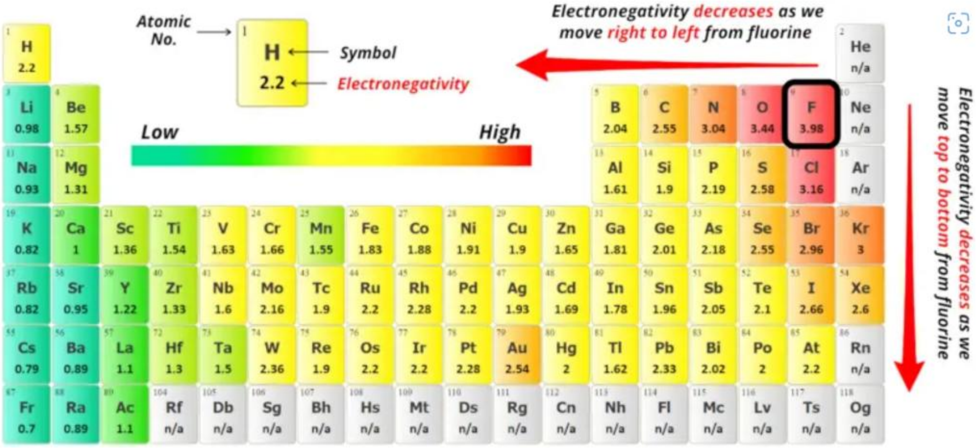
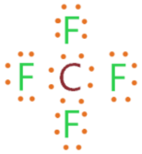
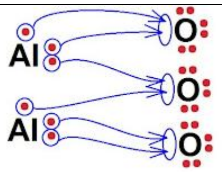
Acid	Name	1 mM	10 mM	100 mM
H_2SeO_4	selenic acid	2.74	1.83	1
H_2SO_4	sulfuric acid (oil of vitriol)	2.75	1.87	1
HI	hydroiodic acid (muriatic acid)	3.01	2.04	1.1
HBr	hydrobromic acid	3.01	2.04	1.1
HCl	hydrochloric acid (gastric acid)	3.01	2.04	1.1
HNO_3	nitric acid	3.01	2.04	1.1
H_3PO_4	orthophosphoric acid	3.06	2.26	1.6
H_3AsO_4	arsenic acid	3.08	2.31	1.7
H_2SeO_3	selenous acid	3.15	2.47	1.9
H_2CrO_4	chromic acid	3.03	2.33	2.1
$C_6H_8O_7$	citric acid (lemon juice)	3.24	2.62	2.1
HF	hydrofluoric acid	3.27	2.65	2.1
HNO_2	nitrous acid	3.28	2.67	2.1
HOCN	isocyanic acid	3.35	2.76	2.2
CH_2O_2	formic acid, (formic or methanoic acid)	3.47	2.91	2.4
H_2Se	hydrogen selenide	3.49	2.93	2.4
H_2MoO_4	molybdic acid	3.46	2.94	2.4
$C_3H_6O_3$	lactic acid (milk acid)	3.51	2.96	2.4
$C_2H_4O_2$	acetic acid (vinegar)	3.91	3.39	2.9
H_2CO_3	carbonic acid	4.68	4.18	3.7
H_2S	hydrogen sulfide	4.97	4.47	4
H_3AsO_3	arsenious acid	6.07	5.58	5.1
HCN	hydrocyanic acid	6.11	5.62	5.1
H_3BO_3	boric acid	6.12	5.62	5.1
H_4SiO_4	silicic acid	6.4	5.91	5.4
H_4SiO_4	silicic acid	6.4	6.26	6.3
H_2O	pure water	7.0	7.0	7.0

pH of Bases

Base	Name	1 mM	10 mM	100 mM
$Ba(OH)_2$	barium hydroxide	11.27	12.22	13.08
$Sr(OH)_2$	strontium hydroxide (caustic alkali)	11.27	12.22	13.09
NaOH	sodium hydroxide (lye)	10.98	11.95	12.88
KOH	potassium hydroxide (caustic potash)	10.98	11.95	12.88
Na_2SiO_3	sodium metasilicate	11	11.91	12.62
$Ca(OH)_2$ ($CaO{:}H_2O$)	calcium hydroxide (lime)	11.27	12.2	12.46
Na_3PO_4	trisodium phosphate (food additive)	10.95	11.71	12.12
K_2CO_3	potassium carbonate (potash or pearl ash)	10.52	11	11.36
Na_2CO_3	sodium carbonate (soda ash)	10.52	10.97	11.26
NH_4OH ($NH_3{:}H_2O$)	ammonium hydroxide (Windex)	10.09	10.61	11.12
$Mg(OH)_2$ ($MgO{:}H_2O$)	magnesium hydroxide	10.4	10.4	10.4
$CaCO_3$	calcium carbonate (limestone or calcite)	9.91	9.91	9.91
$Fe(OH)_2$	iron(II) hydroxide (ferrous hydroxide)	9.45	9.45	9.45
$Cd(OH)_2$	cadmium hydroxide	9.36	9.36	9.36
$Na_2B_4O_7$	sodium borate (Borax)	9.21	9.17	9.05
$Co(OH)_2$	cobalt(II) hydroxide	9.15	9.15	9.15
$Zn(OH)_2$	zinc hydroxide	8.88	8.88	8.88
$Ni(OH)_2$	nickel(II) hydroxide	8.37	8.37	8.37
CH_3COOK	potassium acetate (diuretic salt)	7.87	8.33	8.75
CH_3COONa	sodium acetate (acetic acid)	7.87	8.33	8.75
$KHCO_3$	potassium hydrogen carbonate	8.27	8.25	8.13
$NaHCO_3$	sodium hydrogen carbonate (baking soda)	8.27	8.22	8.02
$Be(OH)_2$	beryllium hydroxide	7.9	7.9	7.9
$Cu(OH)_2$	copper(II) hydroxide	7.69	7.69	7.69
$Pb(OH)_2$	lead(II) hydroxide	7.54	7.54	7.54
$Cr(OH)_3$	chromium(III) hydroxide	7.04	7.04	7.04
$Hg(OH)_2$	mercury(II) hydroxide	7.03	7.03	7.03

Source: aqion (27 Oct 2024), pH of Common Acids and Bases, https://www.aqion.de/

Chapter 12: Reduction and Oxidation

<table>
<tr><th>Term</th><th>Description</th><th>Equation</th></tr>
<tr><td>**Chemical Reaction Types**</td><td>1. Synthesis (Formation)
2. Decomposition
3. Single Displacement
4. Double Displacement
5. Combustion
6. Acid-Base
7. Reduction-Oxidation</td><td>1. $A + B \longrightarrow AB$
2. $AB \longrightarrow A + B$
3. $A + BC \longrightarrow AC + B$
4. $AB + CD \longrightarrow AD + CB$
5. $C_3H_8 + 5O_2 \longrightarrow 3CO_2 + 4H_2O$ + heat (propane)
6. $HA + BOH \longrightarrow BA + H_2O$
7. $Fe^0 + O_2^0 \longrightarrow Fe_2^{3+} + O_3^{2-}$</td></tr>
<tr><td>**Oxidation State**</td><td colspan="2">The charge on an ion, or for a molecule, the charge that an atom would have, if the shared electrons in a bond were always given to the more **electronegative** atom.

<table>
<tr><th colspan="2">Covalent</th><th colspan="2">Ionic</th></tr>
<tr><td>$CF_4 \longrightarrow C^{4+} + 4F^-$</td><td> </td><td>$Al_2O_3 \longrightarrow 2Al^{3+} + 3O^{2-}$</td><td> </td></tr>
</table></td></tr>
<tr><td>**Oxidation State Rules**</td><td colspan="2">These rules are **always** true:
1. For ions composed of only one atom, the oxidation state is equal to the charge on the ion.
2. For all elements and homonuclear molecules, the oxidation state of each atom is 0.
3. For molecules and polyatomic ions, the sum of the oxidation states must equal the total charge.
4. The oxidation state of F in any compound is 1-.
5. In covalent compounds and polyatomic ions, H has an oxidation state of 1+.

These rules are **usually** true:
6. In most compounds, O has an oxidation state of 2-.
7. Group 7A elements (especially Cl) usually have an oxidation state of 1-.</td></tr>
</table>

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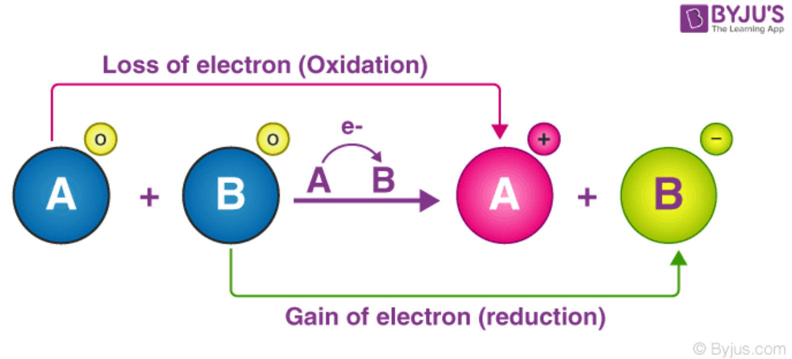
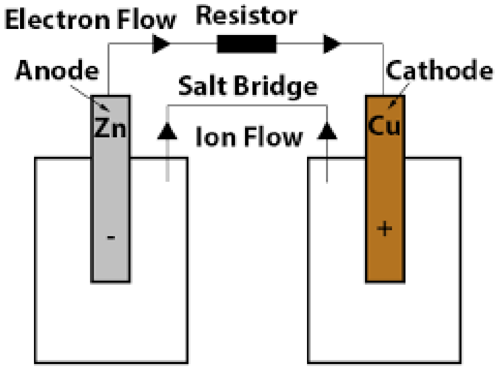
<table>
<tr><td>Reduction-Oxidation (Redox)</td><td colspan="2">

Reactions that transfer electrons from one set of atoms to another.
LEO says GER.</td></tr>
<tr><td>Oxidized</td><td>Atom loses an electron</td><td>“Lose Electrons Oxidation” (LEO)</td></tr>
<tr><td>Reduced</td><td>Atom gains an electron</td><td>“Gain Electrons Reduction” (GER)</td></tr>
<tr><td></td><td colspan="2"><table>
<tr><th>Oxidation</th><th>Reduction</th></tr>
<tr><td>Complete loss of electrons (ionic reactions)</td><td>Complete gain of electrons (ionic reactions)</td></tr>
<tr><td>Shift of electrons away from an atom in a covalent bond</td><td>Shift of electrons towards an atom in a covalent bond</td></tr>
<tr><td>Gain of oxygen</td><td>Loss of oxygen</td></tr>
<tr><td>Loss of hydrogen by a covalent compound</td><td>Gain of hydrogen by a covalent compound</td></tr>
<tr><td>An increase in oxidation number</td><td>A decrease in oxidation number</td></tr>
</table></td></tr>
<tr><td>Oxidizing Agent</td><td colspan="2">If a reactant is reduced, it is called the oxidizing agent, because it oxidizes another reactant.</td></tr>
<tr><td>Reducing Agent</td><td colspan="2">If a reactant is oxidized, it is called the reducing agent, because it reduces another reactant.</td></tr>
<tr><td rowspan="4">Batteries</td><td>Galvanic Cell (battery)</td><td>A system that uses a redox reaction to produce electrical energy.</td></tr>
<tr><td>Anode</td><td>The negative electrode in a Galvanic cell.</td></tr>
<tr><td>Cathode</td><td>The positive electrode in a Galvanic cell.</td></tr>
<tr><td colspan="2">
</td></tr>
<tr><td>Balancing Charge</td><td colspan="2">For a chemical equation to be balanced, the total charge on each side of the equation must be the same.</td></tr>
<tr><td>Electroplating</td><td colspan="2">Using reduction (usually powered by electricity) to coat an object with a metal.</td></tr>
<tr><td>Corrosion</td><td colspan="2">The chemical deterioration of a substance, usually by oxidation. (rust)</td></tr>
</table>

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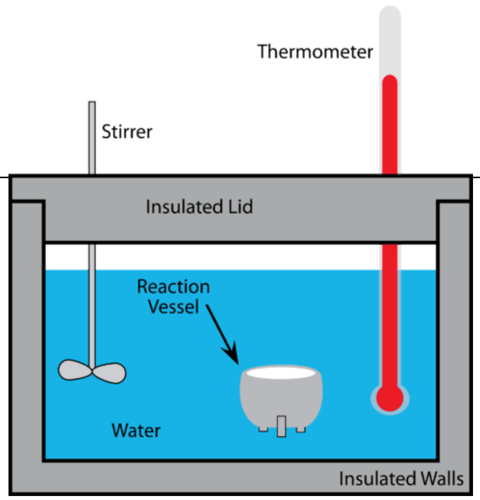
Chapter 13: The Heat Is On

Term	Equation	Note
Energy	$E = W + q + K + \Delta U$	The ability to do work. Units are energy are always in Joules (J). Conservation of energy $\left(E_i = E_f\right)$.
Work	$W = Fd$	The application of a force to move an object over a distance.
Heat	q	Energy that is exchanged because of a difference in temperature or a phase change.
Kinetic Energy	$K = \frac{1}{2}mv^2$	Energy that is in <u>motion</u>.
Potential Energy	$\Delta U_g = mgh$	Energy that is <u>stored</u>.
calorie	$1\ calorie = 4.184\ J$	The amount of heat required to raise 1 gram of water 1 degree Celsius.
Calorie	1 <u>Food</u> **C**alorie = 1,000 chemist **c**alories	Big 'C' vs. little 'c'.
Specific Heat Capacity (c)	$c_{H_2O} = 4.184\ \frac{J}{g\ °C}$	The amount of heat it takes to raise a <u>specific mass</u> of a substance 1 °C.
Heat Capacity (C)	$q = C \cdot \Delta T$	The amount of heat it takes to raise an <u>entire object</u> 1 °C. Takes the mass of the object into account.
Measuring Heat	$\boldsymbol{q = m \cdot c \cdot \Delta T}$ *where:* $m = mass$ $c = specific\ heat\ capacity$ $\Delta T = T_{final} - T_{initial}$	
Calorimeter		
	$-q_{object}$ $= q_{liquid} + q_{calorimeter}$	In a calorimeter experiment, the temperature of the liquid and the calorimeter are always the same. In addition, the *final* temperature of the liquid, calorimeter, and object are all the same.
Latent Heat	$\boldsymbol{q = m \cdot L}$ *where:* L_f $= Latent\ heat\ of\ fusion$ L_v $= Latent\ heat\ of\ vapori$	The amount of heat absorbed or released by a substance during a <u>phase change</u>. $-zation$

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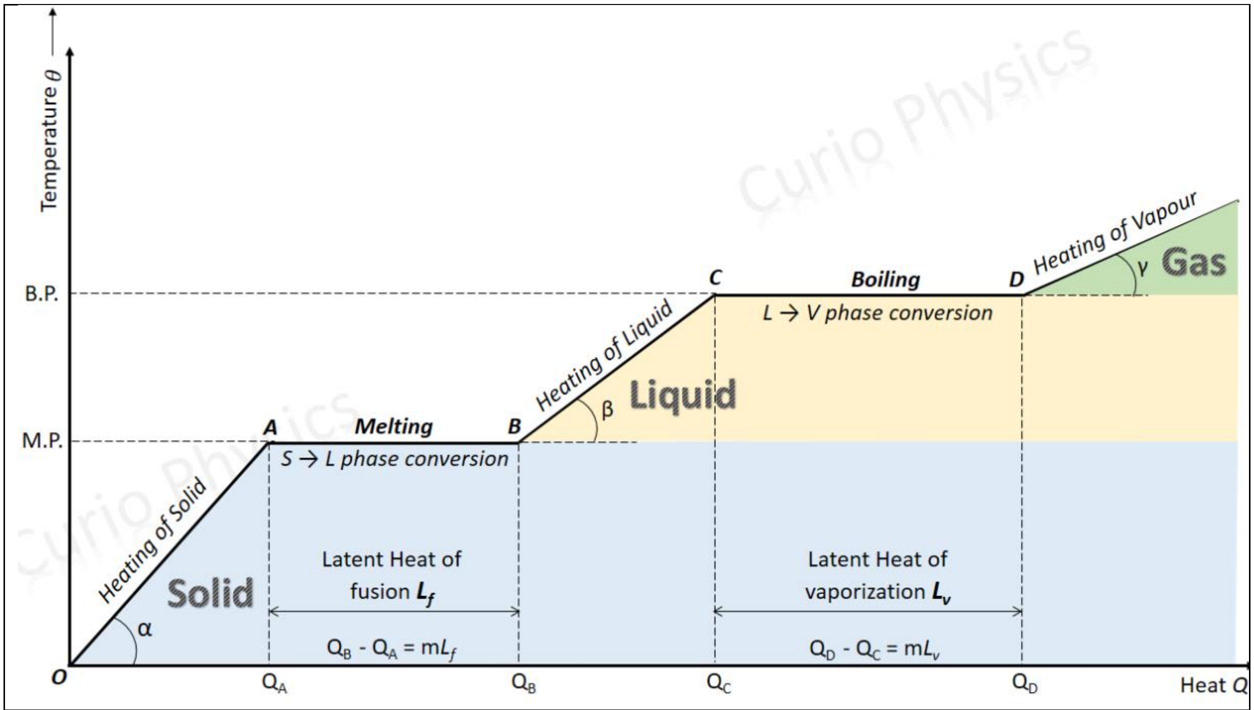
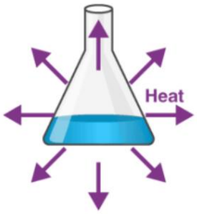
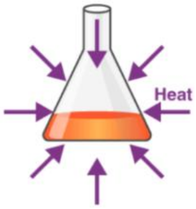


Exothermic Reaction		A chemical reaction that releases energy. e.g., Dilution of sulfuric acid (H_2SO_4) in water.
Endothermic Reaction		A chemical reaction that absorbs energy. e.g., Dissolving ammonium nitrate (NH_4NO_3) in water.

Specific Heat Capacities of Different Substances

Substance	Specific Heat J/g°C		Substance	Specific Heat J/g°C
Water (l)	**4.184**		Nickel (s)	0.440
Water (s)	2.093		Zinc (s)	0.387
Vegetable Oil	2.000		Copper (s)	0.386
Air	1.020		Brass (s)	0.380
Magnesium (s)	1.020		Sand	0.290
Aluminum (s)	0.900		Silver (s)	0.240
Glass	0.840		Tin (s)	0.210
Potassium (s)	0.757		Lead(s)	0.160
Calcium (s)	0.650		Mercury (l)	0.140
Iron (s)	0.444		Gold (s)	0.126

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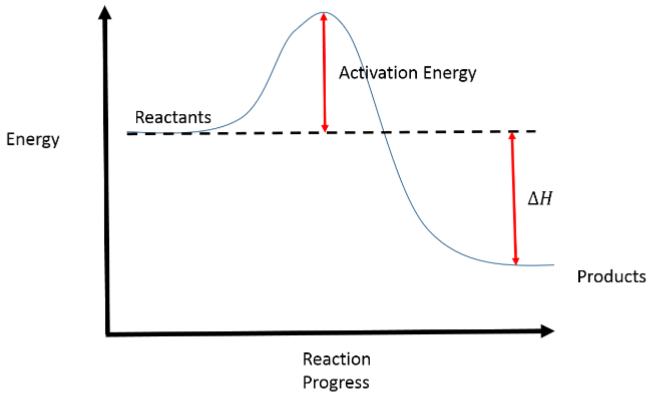
Chapter 14: Thermodynamics

Term	Description	Equation
Enthalpy (ΔH)	The energy change that accompanies a chemical or physical change.	
Enthalpy Observations	• For exothermic reactions, ΔH is negative (–). • For endothermic reactions, ΔH is positive (+).	
Enthalpy	*ΔH = (Energy for breaking bonds) – (Energy from making bonds)*	
Bond Energy	The energy required to break a mole of a given type of bond.	
State Function	A property that is independent of the path.	
Default Enthalpy	The ΔH_f° of any element in its natural phase is zero.	
Hess's Law	Chemical reaction: $aA + bB \rightarrow cC + dD$ $\Delta H = \Delta \mathrm{H}_{Products} - \Delta \mathrm{H}_{Reactants}$ $\Delta H = \left(c \cdot \Delta H_f^\circ(C) + d \cdot \Delta H_f^\circ(D)\right) - \left(a \cdot \Delta H_f^\circ(A) + b \cdot \Delta H_f^\circ(B)\right)$	
Activation Energy	The energy required to initiate a chemical reaction.  	
Thermodynamics	The study of the relationships and conversions between different forms of energy. • 1st Law: Energy cannot be created or destroyed. It can only change form. • 2nd Law: The entropy of the universe can never decrease. It must always stay the same or increase.	

Entropy (ΔS)	A measure of the amount of thermal energy in a system that is not available to do useful work. Because work is obtained from ordered molecular motion, entropy is also a measure of the molecular disorder, or randomness, of a system.
Change in Entropy	$\Delta S_{system} + \Delta S_{surroundings} \geq 0$
Entropy Observations	• The solid phase is the lowest-entropy phase. • The gas phase is the highest-entropy phase. • The larger the number of molecules, the higher the entropy.
Entropy Reactions	Chemical reaction: $aA + bB \rightarrow cC + dD$ $\Delta S = \Delta S_{Products} - \Delta S_{Reactants}$ $\Delta S = \left(c \cdot S^\circ(C) + d \cdot S^\circ(D)\right) - \left(a \cdot S^\circ(A) + b \cdot S^\circ(B)\right)$

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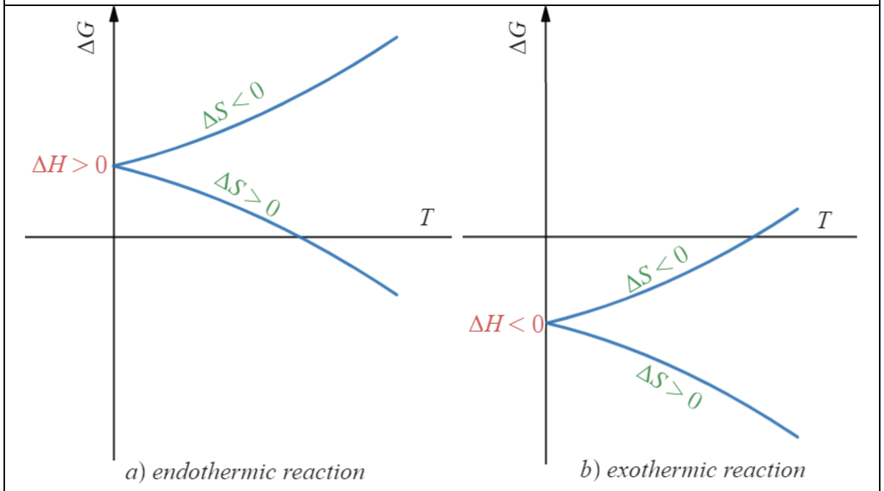
Gibbs Free Energy (ΔG)	$\boldsymbol{\Delta G = \Delta H - T \cdot \Delta S}$  a) endothermic reaction b) exothermic reaction
Efficiency	ΔG can predict chemical reaction efficiency. $Efficienncy = \frac{\Delta G}{\Delta H} \times 100\%$
Gibbs Observations	• When ΔG is negative (–), the reaction is <u>spontaneous</u>. • When ΔG is positive (+), the reaction is <u>not spontaneous</u>. • The ΔG_f° of any element in its natural phase is zero.
Gibbs Reactions	Chemical reaction: $aA + bB \rightarrow cC + dD$ $\Delta G = \Delta G_{Products} - \Delta G_{Reactants}$ $\Delta G = \left(c \cdot \Delta G_f^{\circ}(C) + d \cdot \Delta G_f^{\circ}(D)\right) - \left(a \cdot \Delta G_f^{\circ}(A) + b \cdot \Delta G_f^{\circ}(B)\right)$

Bond Energies

Bond	Energy (kJ/mole)		Bond	Energy (kJ/mole)		Bond	Energy (kJ/mole)
$H-H$	436		$C-C$	350		$F-F$	159
$H-C$	410		$C-F$	450		$Cl-Cl$	243
$H-O$	460		$C-Cl$	330		$S-H$	340
$H-Cl$	432		$C-O$	350		$S-F$	310
$H-N$	390		$C-N$	300		$O-Cl$	200
$C=C$	611		$C=O$	732		$N\equiv N$	945
$N=O$	607		$O=O$	498		$C\equiv O$	1072

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Standard Enthalpies of Formation

Compound	ΔH_f° (kJ/mole)		Compound	ΔH_f° (kJ/mole)		Compound	ΔH_f° (kJ/mole)
$CH_4\ (g)$	-74.9		$C_2H_6O\ (l)$	-277.7		$H_2O_2\ (aq)$	-191.2
$C_2H_6\ (g)$	-84.7		$CO_2\ (g)$	-393.5		$HCl\ (g)$	-92.3
$C_6H_6\ (g)$	82.6		$CO\ (g)$	-110.5		$HCl\ (aq)$	-167.2
$C_6H_6\ (l)$	49.0		$CS_2\ (g)$	116.9		$NH_3\ (g)$	-45.9
$CH_4O\ (g)$	-200.7		$CS_2\ (l)$	89.7		$NH_3\ (aq)$	-80.3
$CH_4O\ (l)$	-238.7		$H_2O\ (g)$	-241.8		$NH_4Cl\ (s)$	-314.6
$C_2H_6O\ (g)$	-235.1		$H_2O\ (l)$	-285.8		$NaOH\ (s)$	-425.9

Absolute Entropies

Substance	S° (J/mole·K)		Substance	S° (J/mole·K)		Substance	S° (J/mole·K)
$CH_4\ (g)$	186.1		$C_2H_6O\ (l)$	160.7		$H_2O_2\ (aq)$	143.9
$C_2H_6\ (g)$	229.5		$CO_2\ (g)$	213.7		$HCl\ (g)$	186.8
$C_6H_6\ (g)$	269.2		$CO\ (g)$	197.5		$HCl\ (aq)$	56.5
$C_6H_6\ (l)$	173.4		$CS_2\ (g)$	237.9		$NH_3\ (g)$	192.7
$CH_4O\ (g)$	239.7		$CS_2\ (l)$	153.1		$NH_3\ (aq)$	111.3
$CH_4O\ (l)$	126.8		$H_2O\ (g)$	188.7		$NH_4Cl\ (s)$	94.9
$C_2H_6O\ (g)$	282.6		$H_2O\ (l)$	70.0		$NaOH\ (s)$	64.4
$C\ (graphite)$	5.7		$H_2\ (g)$	130.6		$Cl_2\ (g)$	223.0
$C\ (g)$	158.0		$O_2\ (g)$	205.0		$N_2\ (g)$	191.6

Standard Gibbs Free Energy of Formation

Compound	ΔG_f° (kJ/mole)		Compound	ΔG_f° (kJ/mole)		Compound	ΔG_f° (kJ/mole)
$CH_4\ (g)$	-50.8		$C_2H_6O\ (l)$	-174.9		$H_2O_2\ (aq)$	-134.1
$C_2H_6\ (g)$	-32.9		$CO_2\ (g)$	-394.4		$HCl\ (g)$	-95.3
$C_6H_6\ (g)$	129.7		$CO\ (g)$	-137.2		$HCl\ (aq)$	-131.3
$C_6H_6\ (l)$	124.4		$CS_2\ (g)$	66.9		$NH_3\ (g)$	-16.4
$CH_4O\ (g)$	-162.0		$CS_2\ (l)$	65.3		$NH_3\ (aq)$	-26.6
$CH_4O\ (l)$	-166.4		$H_2O\ (g)$	-228.6		$NH_4Cl\ (s)$	-203.1
$C_2H_6O\ (g)$	-168.8		$H_2O\ (l)$	-237.1		$NaOH\ (s)$	-379.7

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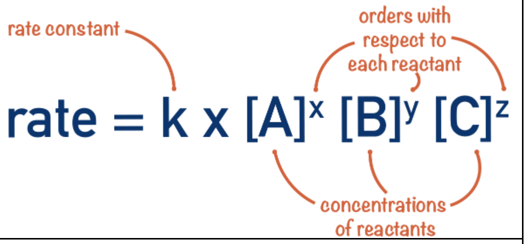
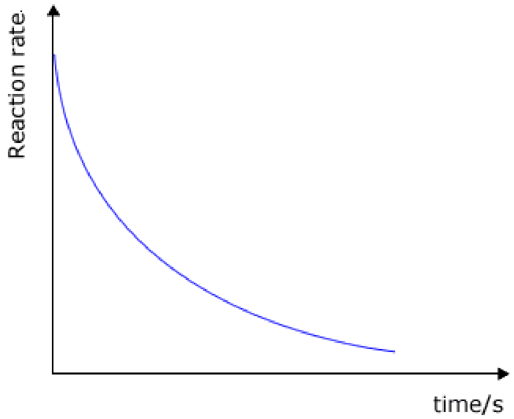
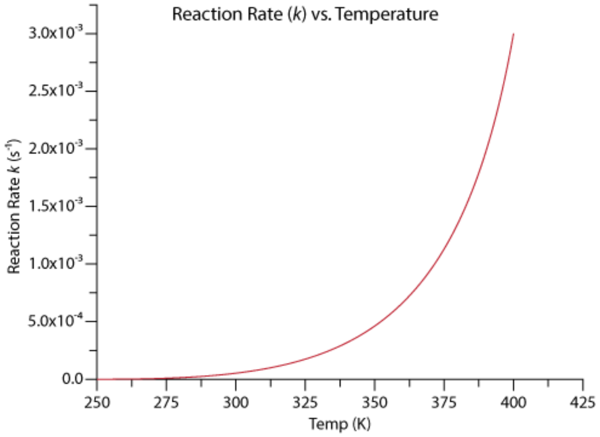
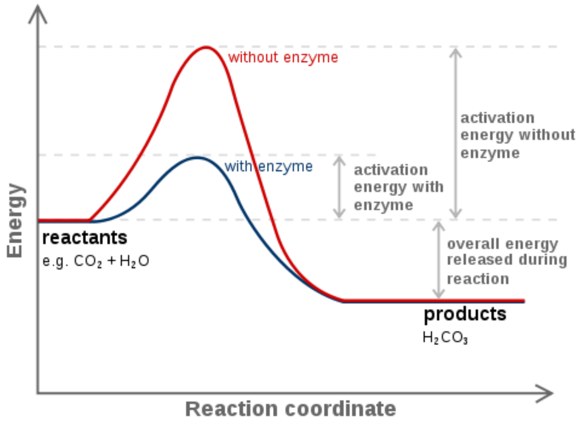
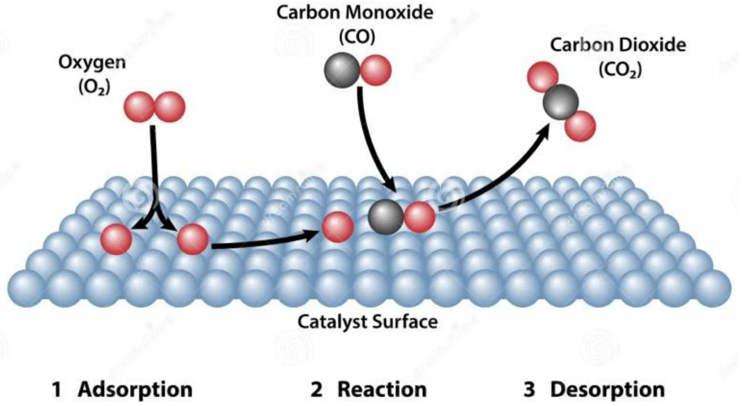
Chapter 15: Kinetics

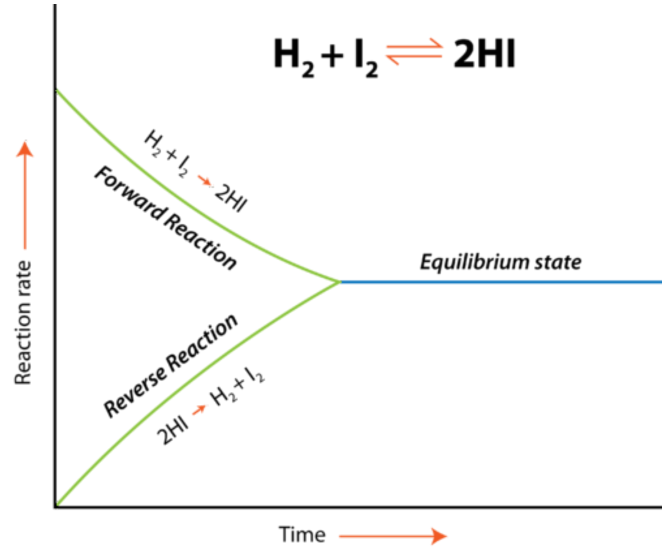
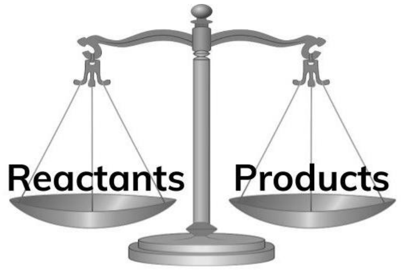
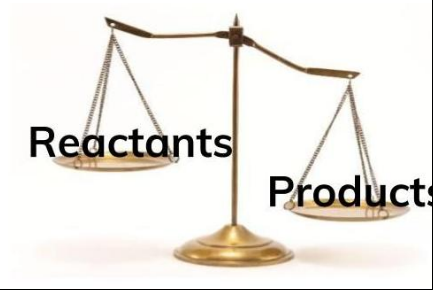
Term	Equation	Note
Kinetics	The study of chemical reaction rates.	
Rate	$Rate = \frac{\Delta[C]}{\Delta t}$ $Rate = \frac{-\Delta[A]}{\Delta t}$	Chemical reaction: $A + B \rightarrow C + D$ $reactants \rightarrow products$ [ ] = "The concentration of"
Reaction Rate Observations	The reaction rate is usually proportional to the: • Concentration of reactants • Surface area over which the reaction can occur • Temperature	
Rate Equation	$Rate = k[A]^x[B]^y$	Chemical reaction: $aA + bB \rightarrow cC + dD$ 
Rate Constant	• The units for the rate constant depend on the overall order of the reaction. • The rate constant increases exponentially with increasing temperature. • The rate constant decreases with increasing activation energy.	

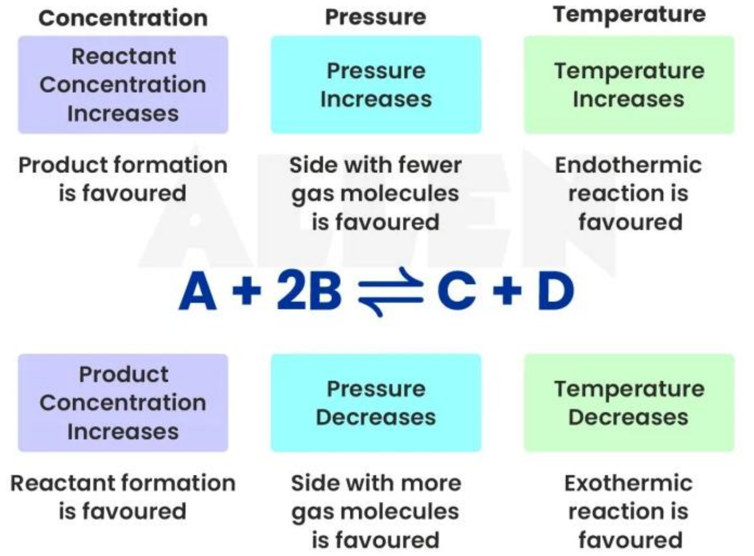




Catalyst	A chemical that increases the rate of a chemical reaction <u>without being used up</u> in the process. i.e., It lowers the activation energy of the reaction by pulling the molecules closer to one another than they would normally be. 
Catalase	A common enzyme found in nearly all living organisms exposed to oxygen which catalyzes the decomposition of H_2O_2 to H_2O and O_2.
Heterogeneous Catalyst	A catalyst in a phase that is <u>different</u> from that of the reactants. e.g., Catalytic converter (g) + (s).
Catalytic Converter 	A device attached to the exhaust of a car that speeds up the following reaction: $2CO\ (g) + O_2\ (g) \rightarrow 2CO_2\ (\boldsymbol{g})$ The catalyst is the platinum-coated (**solid**) mesh that lowers the activation energy. 
Homogeneous Catalyst	A catalyst in a phase that is the <u>same</u> as that of the reactants. $2O_3(g) \rightarrow 3O_2(g)$
Reaction Mechanism	A detailed, step-by-step process that tells you exactly how a reaction occurs. $2O_3(g) \rightarrow 3O_2(g)$ Step 1: $\boldsymbol{NO}\ (g) + O_3\ (g) \rightarrow \boldsymbol{NO_2}\ (g) + O_2\ (g)$ Step 2: $O_3\ (g) \rightarrow O_2\ (g) + O\ (g)$ due to UV light Step 3: $\boldsymbol{NO_2}\ (g) + O\ (g) \rightarrow \boldsymbol{NO}\ (g) + O_2\ (g)$

Chapter 16: Chemical Equilibrium

Term	Description	Equation
Chemical Equilibrium	The state that occurs in a chemical reaction when the rate of the forward reaction equals the rate of the reverse reaction.	
Example		
Equilibrium Constant	$K = \frac{[C]_{eq}^{c}\,[D]_{eq}^{d}}{[A]_{eq}^{a}\,[B]_{eq}^{b}}$	Chemical reaction: $aA + bB \rightarrow cC + dD$ $reactants \rightarrow products$
	The value of the equilibrium constant for a given reaction changes with temperature.	
Interpreting the Constant	$\boldsymbol{K < 1}$: - When the equilibrium constant is small, the reaction makes fewer products and has lots of reactants. - The smaller it is, the more reactants there are and the fewer products there are at equilibrium. $\boldsymbol{K = 1}$: - When the equilibrium constant is 1, the reaction is balanced between reactants and products. $\boldsymbol{K > 1}$: - When the equilibrium constant is large, the reaction makes lots of products and has few reactants. - The larger it is, the more products there are and the fewer reactants there are at equilibrium.	
	$K = 1$ 	$K > 1$ 

Focus on Gas	Do not include <u>solid</u> or <u>liquid</u> reactants or products in the equation for the equilibrium constant.
Interpreting K Results	$K = \frac{[C]_{eq}^{c}\ [D]_{eq}^{d}}{[A]_{eq}^{a}\ [B]_{eq}^{b}}$
	For the above equation: • >: If the result is <u>greater than</u> the equilibrium constant, the reaction will shift towards the reactants. • =: If the result is <u>equal</u> to the equilibrium constant, the reaction is at equilibrium. • <: If the result is <u>less than</u> the equilibrium constant, the reaction will shift towards the products.
Le Chatelier's Principle	When a system at equilibrium is stressed, it will shift in a way that relieves the stress and reestablishes equilibrium.
	
K Shifting: Concentration	A system in equilibrium will: • ↑: <u>shift away</u> from the side that experiences an increase in concentration. • ↓: <u>shift towards</u> the side that experiences a decrease in concentration.
K Shifting: Temperature	• ↑: When temperature is <u>raised</u>, an equilibrium will shift away from the side that contains energy. • ↓: When temperature is <u>lowered</u>, it will shift towards the side that contains energy.
K Shifting: Pressure	• ↑: When pressure is <u>raised</u>, an equilibrium will shift away from the side that has the most gas molecules. • ↓: When pressure is <u>lowered</u>, it will shift towards the side that has the most gas molecules • ↔: If there are <u>no gas</u> molecules, the equilibrium doesn't shift when pressure is changed.
Acid Ionization Constant (K_a)	The equilibrium constant for the reaction between an <u>acid</u> and water.
Base Ionization Constant (K_b)	The equilibrium constant for the reaction between a <u>base</u> and water.

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Sources

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- Dr. Jay L. Wile (2015). Discovering Design with Chemistry, 1st Edition.

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