

Harold's High School Chemistry
2nd Semester
Cheat Sheet
 3 October 2026

Chapter 9: Chemists Have Solutions

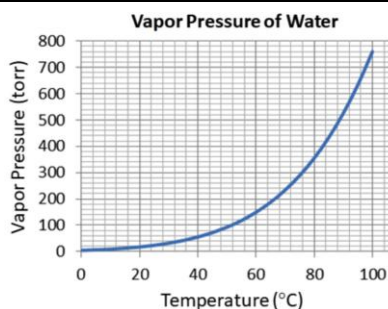
Term	Description	Equation												
Dissolving Compounds	- When <u>ionic</u> compounds dissolve, they split up into their individual ions – charged <u>atoms</u> or groups of atoms. - When polar <u>covalent</u> compounds dissolve, they split up into their individual <u>molecules</u>.													
Solubility of Solutes	<table border="1"> <thead> <tr> <th>State</th><th>Temperature</th><th>Pressure</th></tr> </thead> <tbody> <tr> <td>Solid (s)</td><td>↑</td><td>-</td></tr> <tr> <td>Liquid (l)</td><td>-</td><td>-</td></tr> <tr> <td>Gas (g)</td><td>↓</td><td>↑</td></tr> </tbody> </table>	State	Temperature	Pressure	Solid (s)	↑	-	Liquid (l)	-	-	Gas (g)	↓	↑	
State	Temperature	Pressure												
Solid (s)	↑	-												
Liquid (l)	-	-												
Gas (g)	↓	↑												
Concentration	A chemical's behavior often depends on concentration.													
Molarity	$\text{Molarity} = \frac{\text{moles of solute}}{\text{liters of solution}} M$	Liters												
Molality	$\text{Molality} = \frac{\text{moles of solute}}{\text{kilograms of solvent}} m$	Kilograms												
Freezing Point Depression	$\Delta T = -i \cdot K_f \cdot m$	$K_{f_H_2O} = 1.86\text{ }^{\circ}\text{C}/m$												
Boiling Point Elevation	$\Delta T = i \cdot K_b \cdot m$	$K_{b_H_2O} = 0.512\text{ }^{\circ}\text{C}/m$												
van't Hoff Factor (<i>i</i>)	<i>i</i> is the number of moles of particles formed in solution per mole of solute.													

Chapter 10: It's a Gas!

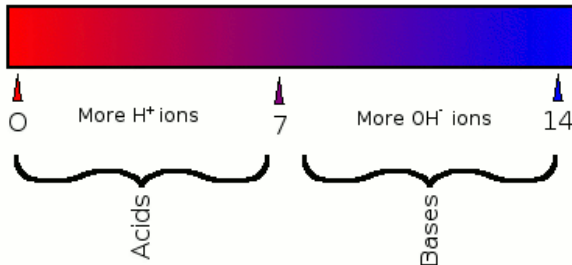
Term	Equation	Note
Ideal Gas Law	$PV = nRT$	$P = \text{Pressure (atm)}$ $V = \text{Volume (L)}$ $n = \# \text{ Moles of gas}$ Ideal Gas Constant (R): $R = 0.0821 \frac{\text{L atm}}{\text{mole K}}$ $T = \text{Temperature (K)}$
Boyle's Law	$P_1 V_1 = P_2 V_2$	Assumes $T = \text{constant}$
Charles's Law	$\frac{V_1}{T_1} = \frac{V_2}{T_2}$	Assumes $P = \text{constant}$
Combined Gas Law	$\frac{P_1 V_1}{T_1} = \frac{P_2 V_2}{T_2}$	Since $nR = \text{constant}$
Avogadro's Law	$n = \frac{PV}{RT} \text{ and } \frac{V_1}{n_1} = \frac{V_2}{n_2}$	Same number of molecules or atoms
Volumes of Gases	$V = \frac{nRT}{P}$	Relationship between volumes of those gases
Kelvin (T)	$K = ^\circ\text{C} + 273.15$ $0 \text{ K} = \text{Absolute Zero}$	$^\circ\text{F} = \left(\frac{9}{5}\right)^\circ\text{C} + 32 \quad [32^\circ - 212^\circ]$ $^\circ\text{C} = \left(\frac{5}{9}\right)(^\circ\text{F} - 32) \quad [0^\circ - 100^\circ]$
Pressure (P)	$P = \frac{F}{A}$	$\text{Pressure} = \frac{\text{Force}}{\text{Area}}$
	$1 \text{ atm} = 760 \text{ torr}$ $1 \text{ torr} = 1 \text{ mmHg}$ $1 \text{ mmHg} = 133.322 \text{ Pa}$ $1 \text{ atm} = 101,325 \text{ Pascals (Pa)}$	Units can be torr, atm, or Pa
Ideal Gas	- The molecules or atoms that make up the gas occupy <u>no volume</u>. - The molecules or atoms that make up the gas are <u>not attracted</u> to each other. - 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 .	Standard Temperature and Pressure
Dalton's Law	$P_{\text{total}} = P_a + P_b + P_c + \dots$	Partial pressures
Mole Fraction	$X_a = \frac{P_a}{P_{\text{total}}}$	$X_a = \frac{\text{moles of compound } a}{\text{total moles in the mixture}}$
Vapor Pressure	Boiling point = the temperature at which a liquid's vapor pressure equals the external air pressure.	
Extrapolation	Extending a trend in data to situations for which no measurements have been made. Usually linear approximations.	

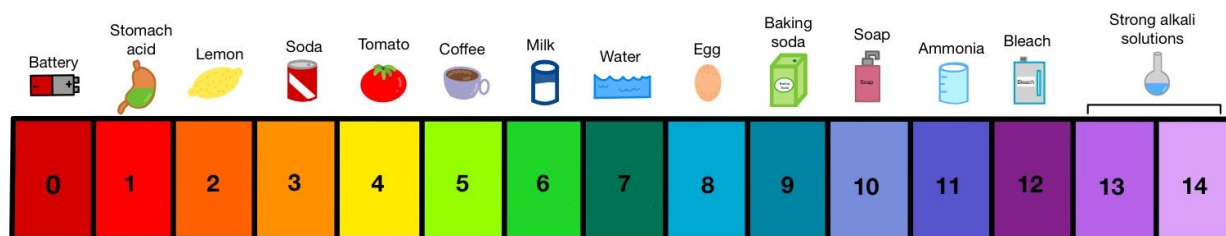
Vapor Pressure of Water

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
38	49.7		100	760.0



Chapter 11: Some Pretty Basic (and Acidic) Chemicals

Term	Equation / Definition	Note
Acids (Acidic)	A chemical that <u>donates</u> an H^+ . Have more H^+ ions.	1. Tend to taste <u>sour</u> 2. Are covalent electrolytes 3. Turn <u>blue</u> litmus paper <u>red</u>
Bases (Alkaline)	A chemical that <u>accepts</u> an H^+ . Have more OH^- ions.	1. Tend to taste <u>bitter</u> 2. Tend to feel slippery when mixed with water 3. Turn <u>red</u> litmus paper <u>blue</u>
Litmus	An <u>acid</u> / <u>base</u> indicator that is usually on a strip of paper	
Amphoteric	Capable of reacting as either an <u>acid</u> or a <u>base</u> (H_2O , metal oxides)	
Covalent Electrolytes	hydrofluoric acid + water $\rightleftharpoons$ fluoride ion + hydronium ion $HF + H_2O \rightleftharpoons F^- + H_3O^+$ <u>acid</u> <u>base</u>	
	ammonia + water $\rightleftharpoons$ ammonium ion + hydroxide ion $NH_3 + H_2O \rightleftharpoons NH_4^+ + OH^-$ <u>base</u> <u>acid</u>	
Ionic Electrolytes	hydrochloric acid + sodium hydroxide (lye) $\rightarrow$ water + salt $HCl(aq) + NaOH(aq) \rightarrow H_2O(l) + NaCl(aq)$ <u>acid</u> <u>base</u> (Na^+OH^-) (Na^+Cl^-)	
Acid/Base Identification Rules	- Ammonia (NH_3) is a covalent <u>base</u>. - If a covalent compound starts with an H, it can usually act like an <u>acid</u>. - Ionic compounds that contain the hydroxide ion (OH^-) can act as <u>bases</u>. - An <u>acid</u> reacts with an ionic <u>base</u> to make water (H_2O) and salt ($NaCl$).	
pH Scale	Potential of hydrogen (pH). Amount of hydronium ion (H_3O^+) in the solution.	
	pH of 0: highly <u>acidic</u> pH of 7: neutral pH of 14: highly <u>alkaline</u>	
pH	$pH \approx -\log[H_3O^+]$, $pOH \approx -\log[OH^-]$, $pH + pOH = 14.00$ at $25^\circ C$	
Polyprotic Acid	An <u>acid</u> that can donate two or more H^+ ions. Examples: sulfuric acid (H_2SO_4), carbonic acid (H_2CO_3) $CO_2(g) + H_2O(l) \rightarrow H_2CO_3(aq)$	
Titration	The process by which an <u>acid</u> of known concentration is added to a <u>base</u> of unknown concentration (or vice versa) until the stoichiometric equivalence point is reached to determine the unknown concentration.	



pH of Acids					pH of Bases				
Acid	Name	1 mM	10 mM	100 mM	Base	Name	1 mM	10 mM	100 mM
H ₂ SeO ₄	selenic acid	2.74	1.83	1	Ba(OH) ₂	barium hydroxide	11.27	12.22	13.08
H ₂ SO ₄	sulfuric acid (oil of vitriol)	2.75	1.87	1	Sr(OH) ₂	strontium hydroxide (caustic alkali)	11.27	12.22	13.09
HI	hydroiodic acid	3.01	2.04	1.1	NaOH	sodium hydroxide (lye)	10.98	11.95	12.88
HBr	hydrobromic acid	3.01	2.04	1.1	KOH	potassium hydroxide (caustic potash)	10.98	11.95	12.88
HCl	hydrochloric acid (muriatic acid)	3.01	2.04	1.1	Na ₂ SiO ₃	sodium metasilicate	11	11.91	12.62
HNO ₃	nitric acid	3.01	2.04	1.1	Ca(OH) ₂ (CaO:H ₂ O)	calcium hydroxide (lime)	11.27	12.2	12.46
H ₃ PO ₄	orthophosphoric acid	3.06	2.26	1.6	Na ₃ PO ₄	trisodium phosphate (food additive)	10.95	11.71	12.12
H ₃ AsO ₄	arsenic acid	3.08	2.31	1.7	K ₂ CO ₃	potassium carbonate (potash or pearl ash)	10.52	11	11.36
H ₂ SeO ₃	selenous acid	3.15	2.47	1.9	Na ₂ CO ₃	sodium carbonate (soda ash)	10.52	10.97	11.26
H ₂ CrO ₄	chromic acid	3.03	2.33	2.1	NH ₄ OH (NH ₃ :H ₂ O)	ammonium hydroxide (Windex)	10.09	10.61	11.12
C ₆ H ₈ O ₇	citric acid (lemon juice)	3.24	2.62	2.1	Mg(OH) ₂ (MgO:H ₂ O)	magnesium hydroxide	10.4	10.4	10.4
HF	hydrofluoric acid	3.27	2.65	2.1	CaCO ₃	calcium carbonate (limestone or calcite)	9.91	9.91	9.91
HNO ₂	nitrous acid	3.28	2.67	2.1	Fe(OH) ₂	iron(II) hydroxide (ferrous hydroxide)	9.45	9.45	9.45
HNCO	isocyanic acid	3.35	2.76	2.2	Cd(OH) ₂	cadmium hydroxide	9.36	9.36	9.36
CH ₂ O ₂	formic acid (formic or methanoic acid)	3.47	2.91	2.4	Na ₂ B ₄ O ₇	sodium borate (Borax)	9.21	9.17	9.05
H ₂ Se	hydrogen selenide	3.49	2.93	2.4	Co(OH) ₂	cobalt(II) hydroxide	9.15	9.15	9.15
H ₂ MoO ₄	molybdic acid	3.46	2.94	2.4	Zn(OH) ₂	zinc hydroxide	8.88	8.88	8.88
C ₃ H ₆ O ₃	lactic acid (milk acid)	3.51	2.96	2.4	Ni(OH) ₂	nickel(II) hydroxide	8.37	8.37	8.37
C ₂ H ₄ O ₂	acetic acid (vinegar)	3.91	3.39	2.9	CH ₃ COOK	potassium acetate (diuretic salt)	7.87	8.33	8.75
H ₂ CO ₃	carbonic acid	4.68	4.18	3.7	CH ₃ COONa	sodium acetate (salt of acetic acid)	7.87	8.33	8.75
H ₂ S	hydrogen sulfide	4.97	4.47	4	KHCO ₃	potassium hydrogen carbonate	8.27	8.25	8.13
H ₃ AsO ₃	arsenious acid	6.07	5.58	5.1	NaHCO ₃	sodium hydrogen carbonate (baking soda)	8.27	8.22	8.02
HCN	hydrocyanic acid	6.11	5.62	5.1	Be(OH) ₂	beryllium hydroxide	7.9	7.9	7.9
H ₃ BO ₃	boric acid	6.12	5.62	5.1	Cu(OH) ₂	copper(II) hydroxide	7.69	7.69	7.69
H ₄ SiO ₄	silicic acid	6.4	5.91	5.4	Pb(OH) ₂	lead(II) hydroxide	7.54	7.54	7.54
H ₄ SiO ₄	silicic acid (SiO ₂ precipitation)	6.4	6.26	6.3	Cr(OH) ₃	chromium(III) hydroxide	7.04	7.04	7.04
H ₂ O	pure water	7.0	7.0	7.0	Hg(OH) ₂	mercury(II) hydroxide	7.03	7.03	7.03

¹ Calculated at 25°C; limited solubility and precipitation are included where applicable.

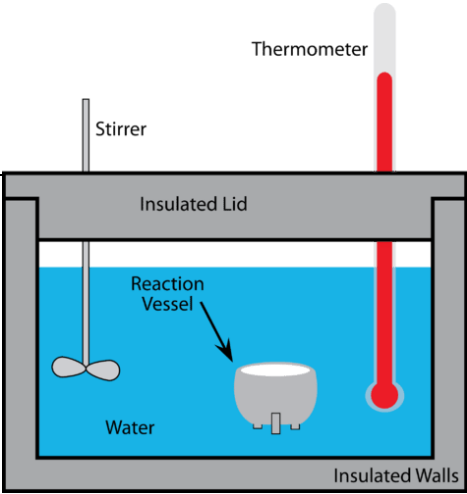
Source: Aqion (27 Oct 2024), [pH of Common Acids and Bases, https://www.aqion.de/](https://www.aqion.de/)

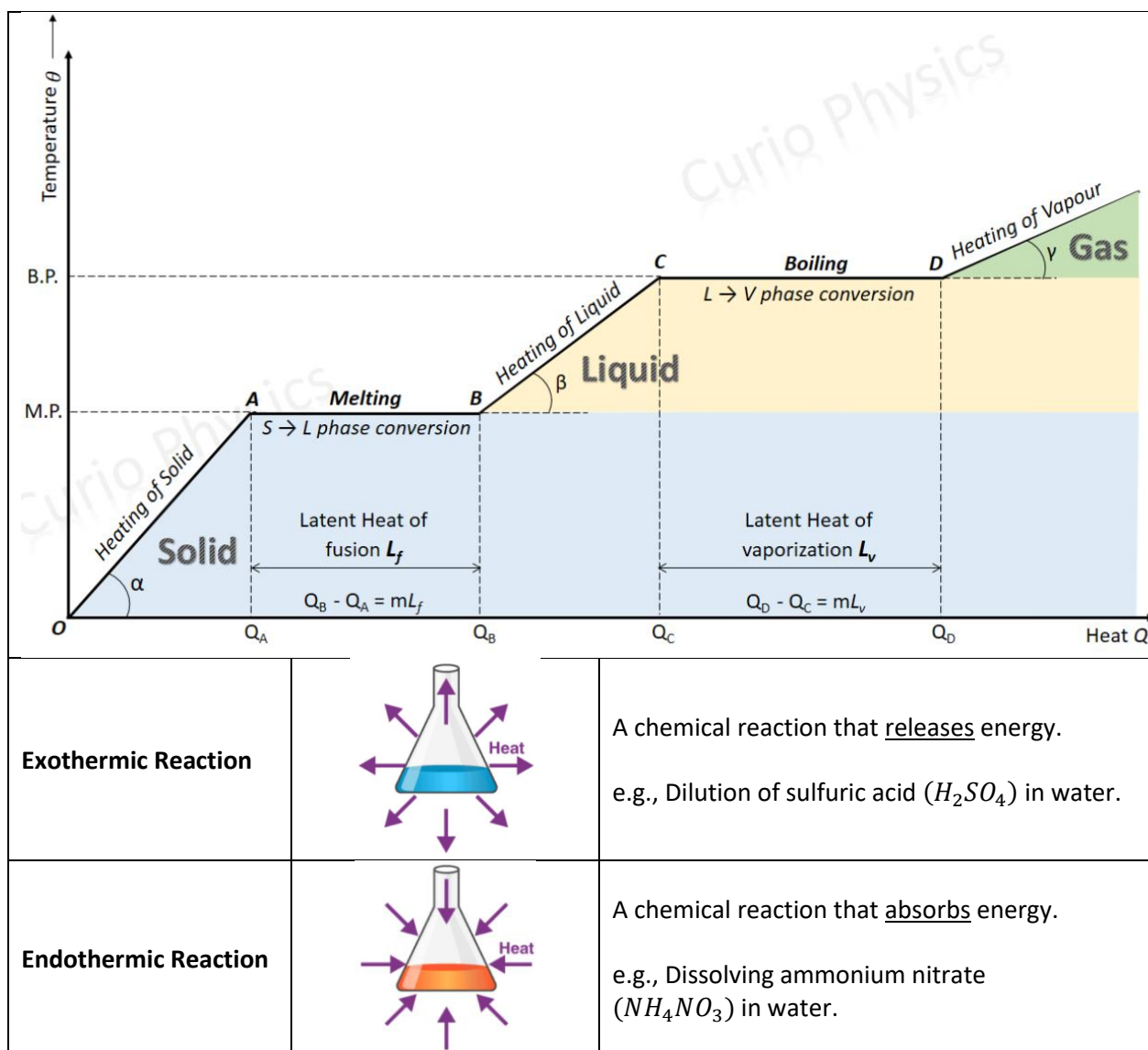
Chapter 12: Reduction and Oxidation

Term	Description	Equation																																								
Chemical Reaction Types	1. Synthesis (Formation) 2. Decomposition 3. Single Displacement 4. Double Displacement 5. Combustion 6. Acid-Base 7. Reduction-Oxidation	1. $A + B \rightarrow AB$ 2. $AB \rightarrow A + B$ 3. $A + BC \rightarrow AC + B$ 4. $AB + CD \rightarrow AD + CB$ 5. $C_3H_8 + 5O_2 \rightarrow 3CO_2 + 4H_2O + \text{heat (propane)}$ 6. $HA + BOH \rightarrow BA + H_2O$ 7. $4Fe^0 + 3O_2^0 \rightarrow 2Fe_2O_3$																																								
Oxidation State	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="10">Covalent</th><th colspan="10">Ionic</th></tr><tr><td colspan="5">$CF_4 \rightarrow C^{4+} + 4F^-$</td><td colspan="5"></td><td colspan="5">$Al_2O_3 \rightarrow 2Al^{3+} + 3O^{2-}$</td><td colspan="5"></td></tr></table>		Covalent										Ionic										$CF_4 \rightarrow C^{4+} + 4F^-$										$Al_2O_3 \rightarrow 2Al^{3+} + 3O^{2-}$									
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Oxidation State Rules	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+ (except for hydrides). 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-.																																									

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

Chapter 13: The Heat Is On

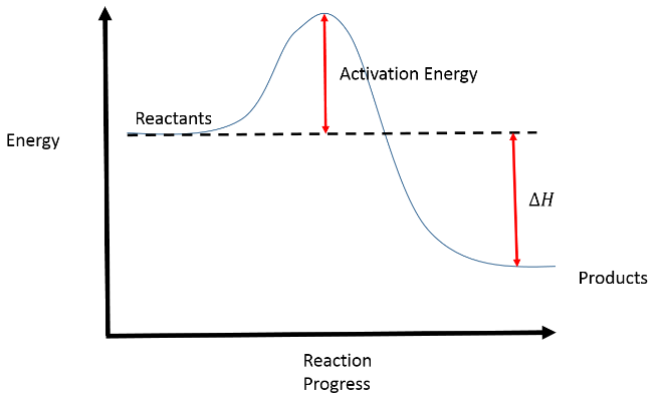
Term	Equation	Note
Energy	$E = W + q + K + \Delta U$	The ability to do work. The SI unit of energy is the joule (J). Conservation of energy ($E_i = E_f$).
Work	$W = Fd$	The application of a force to move an object over a distance.
Heat	q	Energy exchanged because of a temperature difference or a phase 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	$1 \text{ calorie} = 4.184 \text{ J}$	The amount of heat required to raise 1 gram of water 1 degree Celsius.
Calorie	$1 \text{ Food Calorie} = 1,000 \text{ chemist calories}$	Big 'C' vs. little 'c'.
Specific Heat Capacity (c)	$c_{H_2O} = 4.184 \frac{\text{J}}{\text{g}^\circ\text{C}}$	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	$q = m \cdot c \cdot \Delta T$ where: $m = \text{mass}$ $c = \text{specific heat capacity}$ $\Delta T = T_{\text{final}} - T_{\text{initial}}$	
Calorimeter	$-q_{\text{object}}$ $= q_{\text{liquid}} + q_{\text{calorimeter}}$	
Latent Heat	$q = m \cdot L$ where: L_f $= \text{Latent heat of fusion}$ L_v $= \text{Latent heat of vaporization}$	The amount of heat absorbed or released by a substance during a <u>phase change</u> . —zation



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	Mercury (l)	0.140
Calcium (s)	0.650	Lead(s)	0.130
Iron (s)	0.444	Gold (s)	0.126

Chapter 14: Thermodynamics

Term	Description	Equation
Change in Enthalpy (ΔH)	The <u>energy change</u> that accompanies a chemical or physical change. $\Delta H = (\text{Energy for breaking bonds}) - (\text{Energy from making bonds})$	
Enthalpy Observations	- For exothermic reactions, ΔH is negative (-). - For endothermic reactions, ΔH is positive (+).	
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 standard reference state is zero.	
Hess's Law	Chemical reaction: $aA + bB \rightarrow cC + dD$ $\Delta H = \Delta H_{\text{Products}} - \Delta H_{\text{Reactants}}$ $\Delta H = (c \cdot \Delta H_f^\circ(C) + d \cdot \Delta H_f^\circ(D)) - (a \cdot \Delta H_f^\circ(A) + b \cdot \Delta H_f^\circ(B))$	
Activation Energy	The energy required to initiate a chemical reaction. 	
Thermodynamics	The study of the relationships and conversions between different forms of energy. <u>1st Law</u>: Energy cannot be created or destroyed. It can only change form. <u>2nd Law</u>: The entropy of the universe can never decrease. It must always stay the same or increase.	

Entropy (S)	A measure of energy dispersal and the number of accessible microscopic arrangements. Because work is obtained from ordered molecular motion, entropy is also a measure of the molecular <u>disorder</u> , or randomness, of a system.	
Change in Entropy (ΔS)	$\Delta S_{\text{system}} + \Delta S_{\text{surroundings}} \geq 0$	
Entropy Observations	- The <u>solid</u> phase is the <u>lowest</u>-entropy phase. - The <u>gas</u> phase is the <u>highest</u>-entropy phase. - The larger the number of molecules, the higher the entropy.	
Entropy Reactions	Chemical reaction: $aA + bB \rightarrow cC + dD$ $\Delta S = \Delta S_{\text{Products}} - \Delta S_{\text{Reactants}}$ $\Delta S = (c \cdot S^\circ(C) + d \cdot S^\circ(D)) - (a \cdot S^\circ(A) + b \cdot S^\circ(B))$	

Change in Gibbs Free Energy (ΔG)	$\Delta G = \Delta H - T \cdot \Delta S$
Gibbs Free Energy (G)	a) endothermic reaction b) exothermic reaction
Efficiency	ΔG can predict chemical reaction (ideal fuel-cell) efficiency. $Efficiency = \frac{\Delta G}{\Delta H} \times 100\%$
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 standard reference state is zero.
Gibbs Reactions	Chemical reaction: $aA + bB \rightarrow cC + dD$ $\Delta G = \Delta G_{Products} - \Delta G_{Reactants}$ $\Delta G = (c \cdot \Delta G_f^\circ(C) + d \cdot \Delta G_f^\circ(D)) - (a \cdot \Delta G_f^\circ(A) + b \cdot \Delta G_f^\circ(B))$

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 ≡ N	945
N = O	607		O = O	498		C ≡ O	1072

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

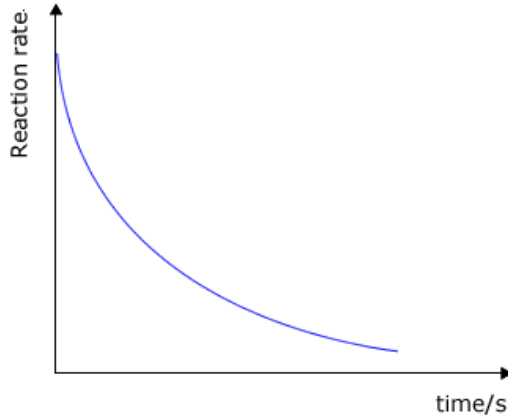
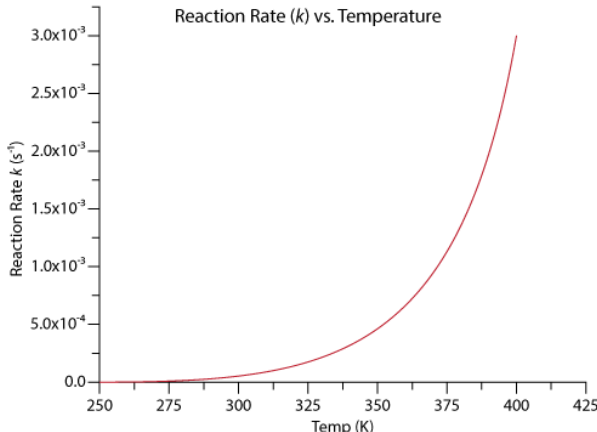
Absolute Entropies

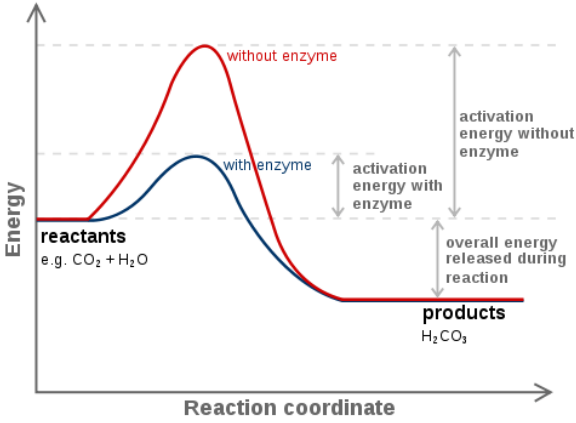
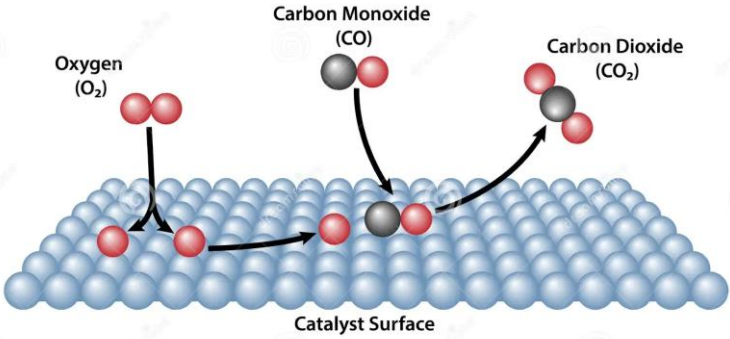
Substance	S° (J/mole·K)		Substance	S° (J/mole·K)		Substance	S° (J/mole·K)
$CH_4(g)$	186.1		$C_2H_6O(l)$	160.7		$H_2O_2(aq)$	143.9
$C_2H_6(g)$	229.5		$CO_2(g)$	213.7		$HCl(g)$	186.8
$C_6H_6(g)$	269.2		$CO(g)$	197.5		$HCl(aq)$	56.5
$C_6H_6(l)$	173.4		$CS_2(g)$	237.9		$NH_3(g)$	192.7
$CH_4O(g)$	239.7		$CS_2(l)$	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
$C(\text{graphite})$	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

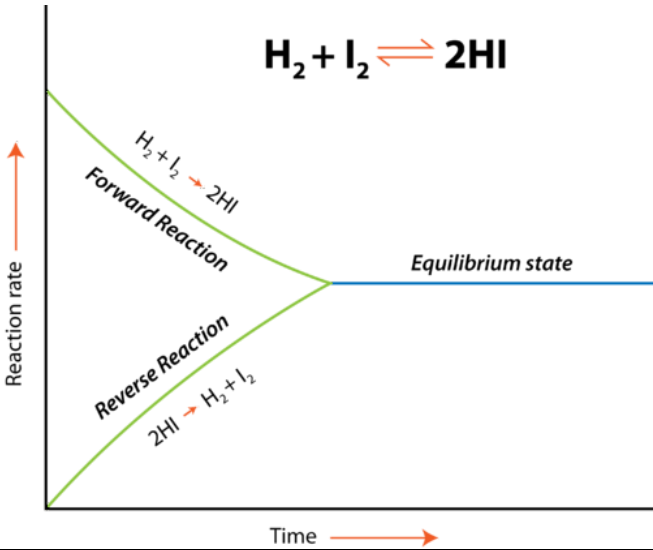
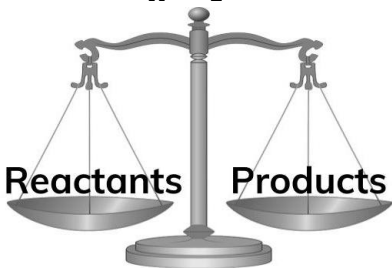
Compound	ΔG_f° (kJ/mole)		Compound	ΔG_f° (kJ/mole)		Compound	ΔG_f° (kJ/mole)
$CH_4(g)$	-50.8		$C_2H_6O(l)$	-174.9		$H_2O_2(aq)$	-134.1
$C_2H_6(g)$	-32.9		$CO_2(g)$	-394.4		$HCl(g)$	-95.3
$C_6H_6(g)$	129.7		$CO(g)$	-137.2		$HCl(aq)$	-131.3
$C_6H_6(l)$	124.4		$CS_2(g)$	66.9		$NH_3(g)$	-16.4
$CH_4O(g)$	-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

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: $A + B \rightarrow C + D$ <i>reactants</i> $\rightarrow$ <i>products</i> [] = "The concentration of"
Reaction Rate Observations	The reaction rate depends upon: • Reactant's concentration • Surface area over which the reaction can occur • Temperature NOTE: A catalyst provides an alternative reaction pathway.	
Rate Equation	$Rate = k[A]^x[B]^y$	Chemical reaction: $aA + bB \rightarrow cC + dD$ $rate = k \times [A]^x [B]^y [C]^z$ rate constant orders with respect to each reactant concentrations of reactants
Rate Constant	• The units for the rate constant depend on the overall order of the reaction. • The rate constant generally increases with temperature according to $k = Ae^{-E_a/(RT)}$. • The rate constant decreases with increasing activation energy.	
 Reaction rate time/s  Reaction Rate (k) vs. Temperature Reaction Rate <i>k</i> (s⁻¹) Temp (K)		

Catalyst	A chemical that increases the rate of a chemical reaction <u>without being used up</u> in the process. (i.e., provides an alternative reaction pathway with lower activation energy.) 
Catalase	A common enzyme found in nearly all living organisms exposed to oxygen that 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: $2CO(g) + O_2(g) \rightarrow 2CO_2(g)$ The catalyst is the platinum-coated (solid) mesh that lowers the activation energy.  1 Adsorption 2 Reaction 3 Desorption
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: $NO(g) + O_3(g) \rightarrow NO_2(g) + O_2(g)$ Step 2: $O_3(g) \rightarrow O_2(g) + O(g)$ due to UV light Step 3: $NO_2(g) + O(g) \rightarrow NO(g) + O_2(g)$

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	 $\text{H}_2 + \text{I}_2 \rightleftharpoons 2\text{HI}$	
Equilibrium Constant	$K = \frac{[\text{C}]_{eq}^c [\text{D}]_{eq}^d}{[\text{A}]_{eq}^a [\text{B}]_{eq}^b}$	Chemical reaction: $a\text{A} + b\text{B} \rightleftharpoons c\text{C} + d\text{D}$ <i>reactants $\rightleftharpoons$ products</i>
	The equilibrium constant for a given reaction changes with temperature.	
Interpreting the Constant	$K < 1$: - 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. $K = 1$: - When the equilibrium constant is 1, the reaction is balanced (not necessarily equal) between reactants and products. $K > 1$: - 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 pure <u>solids</u> or pure <u>liquids</u> as reactants or products in the equation for the equilibrium constant. Aqueous solutes are fine.
Comparing Q with K	$Q = \frac{[C]^c [D]^d}{[A]^a [B]^b}$
	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 shifts to relieve the stress and reestablish equilibrium.
Equilibrium Shifting	A system in equilibrium will:
Concentration	• $\uparrow$: <u>shift away</u> from the side that experiences an increase in concentration. • $\downarrow$: <u>shift towards</u> the side that experiences a decrease in concentration.
Temperature	• $\uparrow$: When temperature is <u>raised</u>, an equilibrium will shift away from the side that contains energy. • $\downarrow$: When temperature is <u>lowered</u>, it will shift towards the side that contains energy.
Pressure (assumes a volume change)	• $\uparrow$: When pressure is <u>raised</u>, an equilibrium will shift away from the side that has the most gas molecules. • $\downarrow$: When pressure is <u>lowered</u>, it will shift towards the side that has the most gas molecules. • $\leftrightarrow$: If there are <u>no gas</u> molecules, the equilibrium doesn't shift when pressure is changed. • 0: No shift when both sides have equal numbers of moles of gas.
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.

Sources

Where possible, these chapters and content are taken verbatim from the High School textbook:

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

Acknowledgements

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