HONORS CHEM 2nd Semester Reference Sheets

Please be understanding if any of these reference sheets change, I am attempting to predict what we will cover and use for the entire school year...this is my best guess!

Contains Reference Sheets for the following chapters.

- Unit 8 Advanced Chemical Ratios
- Unit 9 Gas Laws
- Unit 10 Thermochemistry
- Unit 11 Solutions
- Unit 12 Kinetics
- Unit 13 Equilibrium
- Unit 14 Acids and Bases

Units 1-7 were in the previous Reference Sheet photocopy packet.

Reference Sheets for Unit #8 – Chemical Compositions

Limiting Reagent Walk-Through – Example #1

$N_2 + 3H_2 \rightarrow 2NH_3$ You start with 115.0g of N_2 and 15.00g of H_2

How many grams of product can you make, and how many grams of the excess (XS) reagent do you have left when done?

STEP #1 - Grams to Moles

Use molar masses to convert from grams to moles:

STEP #2 – Have versus Need

Make a little chart showing how many moles of each chemical you have versus how many moles of each chemical you would need. Pick one of your starting values (doesn't matter which one – I like to just pick the first one so I'm consistent), do dimensional analysis to figure out how many moles of the other chemical you would need to have in order to complete the reaction.



STEP #3 – Identify Limiting

Compare the amount you have with the amount you need to see which chemical you don't have enough of, and which chemical you will have extra left over of.

N_2 +	3H₂ →	2NH₃
HAVE:	4.106 mol	7.426 mol
NEED:		12.318 mol

You can see here that you only have 7.426 mol of H_2 , but you would need 12.318 mol to use up all the N_2 you have. So you do not have enough. That means the H_2 is the Limiting Reagent, and N_2 is the excess reagent. If you had more than you needed that would mean it was the excess

reagent and the other was the limiting. That means that you only need to do one calculation to determine which is limiting and which is excess because you can use the answer to one calculation to conceptually explain either result!

STEP #4 – Do Stoich with Limiting

Convert from moles of limiting reactant to desired unit of unknown substance asked for in the problem – use mole highway to determine where to start and end. It is now just a normal stoichiometry problem once you know which number to use! Example pathway: moles of $A \rightarrow$ moles of $B \rightarrow$ grams of B



STEP #5 – Find XS left

Use moles of Limiting Reagent and mole ratio to calculate how many moles of Excess Reagent are needed to use up all the limiting reagent during the reaction. Add this value to the little chart that you already made. Then, just subtract to find how many moles of XS are left over.



Convert your answer into whatever unit is asked for – if it doesn't specify then it is ok to leave it in moles.

$$\begin{array}{c|cccc} 1.631 \text{ mol } N_2 \text{ left over} & 28.01 \text{ g } N_2 \\ \hline 1 \text{ mol } N_2 \end{array} = 45.68 \text{ g of } N_2 \text{ left over} \end{array}$$

$AI(OH)_3 + 3NaCI \rightarrow AICI_3 + 3NaOH$ You start with 28.50 g of $AI(OH)_3$ and 45.00g of NaCl

How many grams of NaOH can you make, and how many moles of the excess (XS) reagent do you have left when done?

STEP #1 - Grams to Moles

Use molar masses to convert from grams to moles:

28.50 g	1 mol	= 0.365 moles	75.00 g	1 mol	= 1.283 moles
AI(OH) ₂	AI(OH) ₂	AI(OH)₂	NaCl	NaCl	Al(OH)₃
/ ((011)3	78.00 g Al(OH)₃			58.44 g NaCl	-

STEP #2 – Have versus Need

Make a little chart showing how many moles of each chemical you have versus how many moles of each chemical you would need. Pick one of your starting values (doesn't matter which one – I like to just pick the first one so I'm consistent), do dimensional analysis to figure out how many moles of the other chemical you would need to have in order to complete the reaction.

	Al(OH)₃	+ 3NaCl →	AlCl₃+3NaOH			
HAVE:	0.365 mol	1.283 mol		0.365 mol Al(OH) ₃	3 mol NaCl	= 1.095 moles H ₂ would be needed
NEED:		1.095 mol			1 mol Al(OH)₃	to use up all the N ₂

STEP #3 – Identify Limiting

Compare the amount you have with the amount you need to see which chemical you don't have enough of, and which chemical you will have extra left over of.

	Al(OH)₃	+ 3NaCl →	AlCl ₃ +3NaOH
HAVE:	0.365 mol	1.283 mol	
NEED:		1.095 mol	

You can see here that you have 1.283 mol of NaCl, but you would only need 1.095 mol to use up all the Al(OH)₃ you have. So you have more than enough NaCl, you will have extra left over. That means NaCl is the excess reagent, and that Al(OH)₃ is the limiting reagent.

STEP #4 – Do Stoich with Limiting

Convert from moles of limiting reactant to desired unit of unknown substance asked for in the problem – use mole highway to determine where to start and end. It is now just a normal stoichiometry problem once you know which number to use! Example pathway: moles of $A \rightarrow$ moles of $B \rightarrow$ grams of B

	LR Al(OH)₃	xs + 3NaCl - 2	AICl ₃ + 3NaC)H
HAVE:	0.365 mo	1.283 mol	-	
NEED:		1.095 mol	-	
0.36 AI((5 mol ⊃H)₃	3 mol NaOH 1 mol Al(OH)₃	40.00 g NaOH 1 mol NaOH	= 43.8 g NaOH made during the rxn

STEP #5 – Find XS left

Use moles of Limiting Reagent and mole ratio to calculate how many moles of Excess Reagent are needed to use up all the limiting reagent during the reaction. Add this value to the little chart that you already made. Then, just subtract to find how many moles of XS are left over.



This time you already know the moles of XS you have and the moles of XS needed to use up all the limiting reagent! So you can skip straight to subtracting.

Convert your answer into whatever unit is asked for – if it doesn't specify then it is ok to leave it in moles.

It only asked for moles, so there is no need to convert our answer to grams! You are done!

Problem:

A compound is 75.46% carbon, 4.43% hydrogen, and 20.10% oxygen by mass. It has a molecular weight of 318.31 g/mol. What is the molecular formula for this compound?

Strategy:

- 1. Find the empirical formula
 - Get the mass of each element by assuming a certain overall mass for the sample (100 g is a good mass to assume when working with percentages).

(.7546) (100 g) = 75.46 g C (.0443) (100 g) = 4.43 g H (.2010) (100 g) = 20.10 g O

- Convert the mass of each element to moles using molar mass of the element. (75.46 g C) (1 mol/ 12.00 g C) = 6.289 mol C
 - (4.43 g H) (1 mol/ 12.00 g C) = 0.289 mol C(4.43 g H) (1 mol/ 1.008 g H) = 4.39 mol H(20.10 g O) (1 mol/ 16.00 g O) = 1.256 mol O
- Find the ratio of the moles of each element. Divide by the smallest mole value found in the previous step. (1.2561 O)/(1.256) = 1 mol O

(6.289 mol C)/(1.256) = 5.007 mol C(4.39 mol H)/(1.256) = 3.50 mol H

- Use the mole values found in the previous step as subscripts to write the empirical formula. $C_5H_{3.5}O_5$
- Multiplying the mole ratios by two to get whole number, the empirical formula becomes: $C_{10}H_7O_2$
- 2. Find the molar mass of the empirical formula. 10(12.00) + 7(1.008) + 2(16.00) = 159.06 g/mol
- **3. Figure out how many empirical units are in a molecular unit.** (318.31 g/mol) / (159.06 g/mol) = 2.001 empirical units per molecular unit
- 4. Write the molecular formula.

Since there are two empirical units in a molecular unit, the molecular formula is: $2 x (C_{10}H_7O_2) = C_{20}H_{14}O_4$

This technique requires that you burn a sample of the unknown substance in a large excess of oxygen gas. The combustion products will be trapped separately from each other and the weight of each combustion product will be determined. From this, you will be able to calculate the empirical formula of the substance. This technique has been most often applied to organic compounds.

Some points to make about combustion analysis:

- 1. The elements making up the unknown substance almost always include carbon and hydrogen. Oxygen is often involved and nitrogen is involved sometimes. Other elements can be involved, but problems with C and H tend to predominate followed by C, H and O and then by C, H, O and N.
- 2. We must know the mass of the unknown substance before burning it.
- 3. All the carbon in the sample winds up as CO_2 and all the hydrogen in the sample winds up as H_2O .
- 4. If oxygen is part of the unknown compound, then its oxygen winds up incorporated into the oxides. The mass of oxygen in the sample will almost always be determined by subtraction.
- 5. Often the N is determined via a second experiment and this introduces a bit of complexity to the problem. Nitrogen dioxide is the usual product when nitrogen is involved. Sometimes the nitrogen product is N₂, sometimes NH₃.
- 6. Sometimes the problem asks you for the empirical formula and sometimes for the molecular formula (or both). Two points:
 - a. You have to know the molar mass to get to the molecular formula
 - b. You have to calculate the empirical formula first, even if the question doesn't ask for it.

Here is a brief overview of the solution steps before doing the example problems:

- 1. Determine the grams of each element present in the original compound. Carbon is always in CO₂ in the ratio (12.011 g / 44.0098 g), hydrogen is always in H₂O in the ratio (2.0158 g / 18.0152 g), etc.
- 2. Convert grams of each element to the number of moles. You do this by dividing the grams by the atomic weight of the element. Many times students will want to use 2.016 for hydrogen, thinking that it is H₂. This is wrong, use 1.008 for H.
- 3. Divide each molar amount by the lowest value, seeking to modify the molar amounts into small, whole numbers.

Steps 2 and 3 are the technique for determining the empirical formula. Step one is required because you have all your carbon, for example, in the form of CO_2 instead of a simpler problem where it tells you how much carbon is present.

Finally, a common component of this type of problem is to provide the molecular weight of the substance and ask for the molecular formula. For example, the empirical formula of benzene is CH while the molecular formula is C_6H_6 .

Problem #1: A 1.50 g sample of hydrocarbon (containing only carbon and hydrogen) undergoes complete combustion to produce 4.40 g of CO₂ and 2.70 g of H₂O. What is the empirical formula of this compound?

Solution:

- Determine the grams of carbon in 4.40 g CO₂ and the grams of hydrogen in 2.70 g H₂O. carbon: 4.40 g x (12.011 g / 44.0098 g) = 1.20083 g C hydrogen: 2.70 g x (2.0158 g / 18.0152 g) = 0.3021482 g H
- Convert grams of C and H to their respective amount of moles. carbon: 1.20083 g / 12.011 g/mol = 0.09998 mol C hydrogen: 0.3021482 g / 1.0079 g/mol = 0.2998 mol H
- 3. Divide each molar amount by the lowest value, seeking to modify the above molar amounts into small, whole numbers.

carbon: 0.09998 mol / 0.09998 mol = 1 hydrogen: 0.2998 mol / 0.09998 mol = 2.9986 = 3

We have now arrived at the answer: the empirical formula of the substance is CH₃

Problem #2: A 0.2500 g sample of a compound known to contain carbon, hydrogen and oxygen undergoes complete combustion to produce 0.3664 g of CO₂ and 0.1500 g of H₂O. What is the empirical formula?

Solution:

- Determine the grams of carbon in 0.3664 g CO₂ and the grams of hydrogen in 0.1500 g H₂O. carbon: 0.3664 g x (12.011 g / 44.0098 g) = 0.1000 g C hydrogen: 0.1500 g x (2.0158 g / 18.0152 g) = 0.01678 g H
- 2. Determine the grams of oxygen in the sample by subtraction.
 0.2500 sample (0.1000 g C + 0.01678 g H) = 0.1332 g O Notice that the subtraction is the mass of the sample minus the sum of the carbon and hydrogen in the sample.
- 3. Convert grams of C, H and O to their respective amount of moles. carbon: 0.1000 g / 12.011 g / mol = 0.008325 mol hydrogen: 0.01678 g / 1.0079 g/mol = 0.01665 mol oxygen: 0.1332 g / 15.9994 g/mol = 0.008327 mol
- 4. Divide each molar amount by the lowest value, seeking to modify the molar amounts into small, whole numbers.

carbon: 0.008325 mol / 0.008325 mol = 1 hydrogen: 0.01665 mol / 0.008325 mol = 2 oxygen: 0.008327 mol / 0.008325 mol = 1

We have now arrived at the answer: the empirical formula of the substance is CH₂O

Reference Sheets for Unit #9 – Gas Laws

Kinetic Molecular Theory Assumptions

- 1. Gases are "Ideal Gases" meaning they do not interact with each other.
- 2. Gases consist of large numbers of tiny particles that are far apart relative to their size the volume of each gas molecule is considered negligible, they are treated as point particles.
- 3. Gas particles undergo elastic collisions meaning they do not lose energy when colliding.
- 4. Gas particles are in a constant, rapid, straight line motion they possess kinetic energy (motion energy).
- 5. The average kinetic energy of the particles is proportional to temperature T \uparrow , KE \uparrow
- 6. There is a distribution of speeds some go faster than others but overall there is an average kinetic energy of the sample.

Gas Pressure

- Pressure The force per unit area on a surface
- $pressure = \frac{force}{area}$
- Gas particles exert force, and therefore pressure, on any surface with which they collide.
- You can use various units for pressure. You may have to convert from one to the other occasionally.
- Air pressure is classically found using a mercury barometer

Unit	Symbol	Definition
Pascal	Ра	SI unit for pressure $1 Pa = \frac{1N}{m^2}$
Millimeter of mercury	mmHg	The amount of pressure that will support a 1 mm column of mercury in a barometer. Most barometers are no longer mercury based but they still use mmHg as their units.
Atmosphere	atm	Average atmospheric pressure at sea level and 0°C
Torr	torr	Torr is the same as mm Hg, just named after Evangelista Torricelli who discovered the concept of the barometer.
Pounds per square inch	psi	The pressure exerted when one pound of force is applied to a one square inch area. A common unit in "real life."



Conversions				
	1.01325 x 10⁵ Pa			
	101.325 kPa			
1 atm =	760 mmHg			
	760 torr			
	14.7 psi			

Temperature Unit for Gas Laws

- Temperature is a measure of the molecular movement.
- For gas law problems we want a temperature of "0" to mean the molecules are not moving.
- Kelvin is a unit for temperature that is scaled so that at 0 K there is no molecular movement.
- 0 Kelvin is called "Absolute Zero"
- $K = {}^{\circ}C + 273$

Water Boils 212 °F 100 °C 373 K Water Freezes 32 °F 0 °C 273 K Absolute Zero -459 °F -273 °C 0 K Fahrenheit Celsius Kelvin



Standard Temperature	0°C = 273 K
Standard	1 atm
Pressure	i adri

Standard Conditions

Many problems require that you use "standard conditions" for the given scenario. Rather than having to give you the temperature and pressure, they may just tell you to use "standard conditions" or that it is under "STP." You can convert the pressure unit to whatever is convenient, but the temperature must be in Kelvins.

Kinetic Energy and Temperature

- $KE = \frac{1}{2}mv^2$
- A sample of gas will have molecules traveling at slightly different speeds. We use the <u>average</u> speed.
- The temperature is a measure of the average kinetic energy
- All gases at the same temperature have the same *average* KE
- Small molecules (small mass, *m*) have higher average speeds

Properties of Gases



Factors Affecting Gas Pressure

Amount of Gas

- \uparrow molecules = \uparrow collisions with walls = \uparrow pressure
- \downarrow molecules = \downarrow collisions with walls = \downarrow pressure

<u>Volume</u>

 \uparrow volume = \uparrow surface area = \downarrow collisions *per unit of surface area* = \downarrow pressure \downarrow volume = \downarrow surface area = \uparrow collisions *per unit of surface area* = \uparrow pressure

Temperature

- ↑ temperature = ↑ molecule speed = ↑ frequency and force of collisions = ↑ pressure
- \downarrow temperature = \downarrow molecule speed = \downarrow frequency and force of collisions = \downarrow pressure

Ideal Gases vs. Real Gases

- Ideal Gas An imaginary gas that perfectly fits all the assumptions of the kinetic molecular theory
- Real Gases A gas that does not behave completely according to the assumptions of the kinetic molecular theory. They occupy space and exert attractive and/or repulsive forces on one another

Likely to behave nearly ideally	Likely to NOT behave ideally
Gases at high temperature	Gases at low temperature
and low pressure	and high pressure
Small non-polar gas	Large, polar gas
molecules	molecules

Diffusion and Effusion

- Diffusion
 - Spontaneous mixing of particles from two (or more) substances
 - Caused by their random motion
 - Rate of diffusion is dependent upon:
 - Speed of particles
 - Diameter of particles
 - Attractive/repulsive forces between the particles
- Effusion
 - Process by which particles under pressure pass through a tiny opening
 - Rate of effusion is dependent upon:
 - Speed of particles (small molecules have greater speed than large molecules at the same temperature, so they effuse more rapidly)





Gas Laws Summary Chart

P = Pressure

T = Temperature

V = Volume

n = # moles

Subscript 1 = initial/starting value

Subscript 2 = final/ending value

	Boyle's Law	Charles' Law	Gay-Lussac Law	Avogadro's Law	Combined Law
Variables	Pressure, Volume	Volume, Temperature	Pressure, Temperature	Moles, Volume	Pressure, Volume, Temperature
Constant	Temperature, # moles	Pressure, # moles	Volume, # moles	Pressure, Temperature	# moles
Equation	$P_1V_1 = P_2V_2$	$\frac{V_1}{T_1} = \frac{V_2}{T_2}$	$\frac{P_1}{T_1} = \frac{P_2}{T_2}$	$\frac{V_1}{n_1} = \frac{V_2}{n_2}$	$\frac{P_1V_1}{T_1} = \frac{P_2V_2}{T_2}$
Type of Relationship	Indirect ↑ P= ↓ V ↓ P= ↑ V	Direct ↑ T = V ↑ ↓ T = V ↓	Direct ↑ T = P ↑ ↓ T = P ↓	Direct ↑ n = V ↑ ↓ n = V ↓	
Law Stated in Words	The volume of a fixed amount of gas varies inversely with pressure if temperature is kept constant.	The volume of a fixed amount of gas is directly proportional to its Kelvin temperature if the pressure is kept constant	The pressure of a fixed amount of gas is directly proportional to the Kelvin temperature if the volume is kept constant	The volume of a gas is directly proportional to the number of moles present in the container if the temperature and pressure are kept constant.	Boyle's, Charles' and Gay- Lussac Laws combine the relationship between pressure, volume, and temperature
Real Life Example	Putting oxygen gas under high pressure will compress it and shrink the volume so it can fit into an oxygen tank to use while scuba diving.	A hot air balloon works by heating the gas inside the balloon causing it to expand.	A can of hairspray in a hot car will cause the gas particles to move faster which increases collisions, so the pressure increases, risking exploding the can.	When blowing up a balloon, you add more moles of gas and the volume increases.	A bag of chips is brought from sea level to the top of Mount Everest. As it gets closer to the summit the atmospheric pressure and temperature change which causes the volume of the bag to change also.
Graphical Depiction	1200 900 70 900 0 900 0 900 0 900 900 900	1100 900 700 300 100 0 200 400 600 Temperature (K)	6.0 5.0 (iii) 4.0 3.0 2.0 1.0 100 200 300 400 500 Temperature (K)	Tolune moles (n)	

	Dalton's Law of Partial Pressures	Ideal Gas Law
Equation	$P_{total} = nP_1 + nP_2 + nP_3 \dots$ $P_{atm} \text{ or } P_{total} = P_{gas} - P_{vapor}$	PV = nRT
Law Stated in Words	At constant volume and temperature, the total pressure exerted by a mixture of gases is equal to the sum of the pressures of each gas. When collecting a gas over a liquid (usually water) you need to subtract how much pressure is coming from the liquid's vapor in order to determine what the pressure of the gas you are interested in is.	The Ideal Gas Law relates the pressure, temperature, volume and moles of a gas through the gas constant "R." The ideal gas law reduces to the other gas laws when you start holding different variables constant. The Ideal Gas Law can also relate density and molar mass with the gas constant "R."

	Water Vapor Pressure at Various Temperatures for								
	Dalton's Partial Pressure Problems - Collecting a Gas Over Water								
	Temperature	Pressure	Temperature	Pressure	Temperature	Pressure	Temperature	Pressure	
	(°C)	(mmHg)	(°C)	(mmHg)	(°C)	(mmHg)	(°C)	(mmHg)	
	0.0	4.6	21.0	18.6	27.0	26.7	50.0	92.5	
	5.0	6.5	22.0	19.8	28.0	28.3	60.0	149.4	
	10.0	9.2	23.0	21.1	29.0	30.0	70.0	233.7	
	15.0	12.8	24.0	22.4	30.0	31.8	80.0	355.1	
	18.0	15.5	25.0	23.8	35.0	42.2	90.0	525.8	
	20.0	17.5	26.0	25.2	40.0	55.3	100.0	760.0	
Finding Molar Mass Using Ideal Gas Law	$\begin{array}{c c} \textbf{lar} \\ \textbf{g} \\ \textbf{aw} \end{array} n = \frac{mass}{molar \ mass} \text{So rearrange ideal gas equation} \dots PV = \frac{mRT}{M} \text{ therefore} \dots M = \frac{mRT}{PV} \qquad (m = mass, M = Molar \ Mass) \end{array}$								
Finding Density Using Ideal Gas Law	$D = \frac{m}{V}$ and	$M = \frac{mRT}{PV}$	Substituting D f	for $\frac{m}{v}$ you end	up getting M =	$= \frac{DRT}{P}$ then re	earrange to soly	ve for D you g	et $D = \frac{MP}{RT}$

Key Info	STP = Standard Temperature and Pressure 1 atm and 0°C (273 K) Molar Volume at STP = 22.4L/1mol	Ideal Gas Constant = R	Pick the right R value!
		$0.0821 \underline{L \bullet atm} = 8.31 \underline{L \bullet kPa} = 62.4 \underline{L \bullet mmHg}$	You can always change
		$mol \cdot K \qquad mol \cdot K \qquad mol \cdot K$	the pressure units and use your favorite R value!
	K = 273 + °C Change ALL temperatures to Kelvin!!!!	1 atm = 1.01325 x 10 ⁵ Pa = 101.325 kPa = 760 mmHg	g = 760 torr = 14.7 psi

Water Vapor Pressure at Various Temperatures

Calculator for Water Vapor at Various Temperatures:

https://tinyurl.com/watervaporpressure

Walkthrough about how to collect a gas over water:

https://tinyurl.com/collectinggasoverwater

Т	Р	Т	Р	Т	Р	Т	Р
(°C)	(mmHg)	(°C)	(mmHg)	(°C)	(mmHg)	(°C)	(mmHg)
1	4.8853	26	25.1370	51	96.9771	76	300.7990
2	5.2511	27	26.6642	52	101.8561	77	313.5406
3	5.6409	28	28.2715	53	106.9439	78	326.7347
4	6.0559	29	29.9623	54	112.2477	79	340.3943
5	6.4975	30	31.7402	55	117.7751	80	354.5323
6	6.9673	31	33.6089	56	123.5336	81	369.1619
7	7.4667	32	35.5723	57	129.5310	82	384.2966
8	7.9973	33	37.6344	58	135.7755	83	399.9502
9	8.5608	34	39.7993	59	142.2751	84	416.1368
10	9.1588	35	42.0711	60	149.0384	85	432.8706
11	9.7932	36	44.4543	61	156.0740	86	450.1661
12	10.4659	37	46.9533	62	163.3906	87	468.0383
13	11.1787	38	49.5729	63	170.9974	88	486.5021
14	11.9337	39	52.3178	64	178.9036	89	505.5729
15	12.7330	40	55.1928	65	187.1186	90	525.2664
16	13.5787	41	58.2032	66	195.6521	91	545.59.85
17	14.4732	42	61.3541	67	204.5142	92	566.5854
18	15.4189	43	64.6509	68	213.7147	93	588.2434
19	16.4180	44	68.0992	69	223.2643	94	610.5894
20	17.4733	45	71.7046	70	233.1733	95	633.6405
21	18.5872	46	75.4730	71	243.4526	96	657.4138
22	19.7626	47	79.4105	72	254.1137	97	681.9270
23	21.0023	48	83.5232	73	265.1667	98	707.1980
24	22.3092	49	87.8175	74	276.6242	99	733.2450
25	23.684	50	92.2999	75	288.4977	100	764.2602



Gas Stoichiometry Videos



Tyler DeWitt Video Part 1 https://tinyurl.com/kehnhf4



Example Video https://tinyurl.com/ycnnt37x



Tyler DeWitt Video Part 2 https://tinyurl.com/mb78shh



CK-12 Reading and Videos https://tinyurl.com/yc534qfx



Chem Academy Video https://tinyurl.com/ybwbrxrl



Two Examples with Keys https://tinyurl.com/yarykkg3

Reference Sheets for Unit #10 – Thermochemistry

$\mathbf{Q} = \mathbf{m}\mathbf{C}\Delta\mathbf{T}$

Variable	Meaning	Common Units	
Q	Energy being transferred	Joules, kJ, calorie, Calorie	
m Mass of sample		grams, kilograms	
С	Specific Heat	J/g°C, kJ/g°C	
ΔΤ	Change in Temperature	°C	

Note:

- Temperature is in CELSIUS not Kelvins for this topic!
- Q can be positive or negative, energy absorbed or energy released
- 1 kJ = 1000 J
- 1 calorie = 4.18 J
- 1 Calorie = 1 kcal = 1000 calories
- Specific Heat values can vary a little bit from source to source. That is ok. Depends on the exact type that was measured. Feel free to add other specific heats you come across to this list!

Common Specific Heat Values

Substance	J/g°C	cal/g°C
Aluminum	0.90	0.23
Copper	0.39	0.093
Ethanol	2.44	0.58
Glass	0.50	0.12
Gold	0.13	0.031
Graphite	0.71	0.17
Ice	2.09	0.50
Iron	0.45	0.11
Lead	0.13	0.031
Mercury	0.14	0.033
Silver	0.24	0.057
Steam	1.87	0.45
Water (liq)	4.18	1.00
Wood	1.8	0.42

Calorimetry

- The fundamental idea behind calorimetry is that sometimes we cannot measure things directly, we need to measure them indirectly.
- If we know that the energy INTO a system equals the energy OUT of a system then we can set up an equation where we know the amount of energy being released is equal in number but opposite in sign to the energy being absorbed.

Remember that the negative sign just means

• Remember that the negative sign just means OPPOSITE sign to the other Q value, not necessarily an actual negative value.

 $Q_{substance 1} = -Q_{substance 2}$

• Remember that T_{final substance 1} and T_{final substance 2} are the same! They reach the same ending temperature

For a metal object dropped into water, i.e., a typical calorimetry problem

 $Q_{water} = (m_{water})(C_{water})(\Delta T_{water})$ $Q_{metal} = (m_{metal})(C_{metal})(\Delta T_{metal})$ $Q_{water} = -Q_{metal}$

Therefore...

 $(m_{water})(C_{water})(\Delta T_{water}) = -(m_{metal})(C_{metal})(\Delta T_{metal})$

Same as...

 $(m_{water})(C_{water})(T_{final\ temp\ of\ both} - T_{initial\ water}) = -(m_{metal})(C_{metal})(T_{final\ temp\ of\ both} - T_{initial\ metal})$

Heating and Cooling Curves

This is an example of water – can be done with any substance but the temperature values will be different. Also please note that the slope and length of lines are not drawn to scale. It is traditional to just draw the heating curve. A cooling curve would just be the opposite direction!



Calorimetry Problems Involving Phase Changes

If a calorimetry problem involves phase changes you need to be careful to use the right equations in your calorimetry calculations – you may not use $Q = mC\Delta T$ on both sides, you may need to use Q = mL for part of it!

Example: 75g of ice (0°C) are dropped into 300g of water (25°C). What is the final temperature of the mixture when the ice is finished melting?

You would need to use Q=mL for the ice melting values and $Q = mC\Delta T$ for the water values

mL_{fus} = - (mC_{liq}
$$\Delta$$
T)
(75g)(334 $\frac{J}{g}$) = -(300g)(4.18 $\frac{J}{g^{\circ}c}$)(Tf - 25°C)

Reference Sheets for Unit #11 – Solutions

Definitions

<u>Solute</u> The substance that is being dissolved in a solution

<u>Solvent</u> The substance that something is being dissolved into

<u>Solution</u> The solute and solvent combined

Solubility

The amount of solute that can be dissolved at a given temperature

<u>Saturated solution</u> Maximum amount of solute dissolved

<u>Unsaturated solution</u> Less than the maximum amount of solute dissolved

<u>Supersaturated solution</u> More than the maximum amount of solute dissolved

Dissolve

When molecules of solute are surrounded by molecules of solvent and are pulled apart from other solute molecules

Dissociate

When an ionic compound has it's ionic bond disrupted by solvent molecules and it breaks into individual ions

<u>Electrolytes</u> Ionic solutes that dissociate into ions in a solution

<u>Non-electrolytes</u> Covalent compounds that do not dissociate into ions in a solution

<u>Heat of solution</u> The energy involved when solute dissolves/dissociates

Equations

 $Mass Percent = \left(\frac{mass of solute}{mass of solution}\right) x \ 100$

Parts per Million =
$$\left(\frac{mass \ of \ solute}{mass \ of \ solution}\right) x 1,000,000$$

$$Grams \ per \ Liter = \left(\frac{mass \ of \ solute}{volume \ of \ solution}\right)$$

Mole fraction of
$$A = Xa = \left(\frac{n_A}{n_A + n_B}\right)$$

$$Molarity = M = \left(\frac{moles \ of \ solute}{Liters \ of \ solution}\right)$$

 $Dilutions = M_1 V_1 = M_2 V_2$

Reference Sheets for Unit #12 – Kinetics

Wait...which rate thing this time?



Can't see it?

Then you need to memorize it or actually do the math in your calculator $\ensuremath{\mathfrak{S}}$

Change factor to Rate = [Change factor for Concentration]^x

Examples of determining the orders by actually plugging in				
rate doesn't change	$1 = 2^{x}$	concentration doubles	x = 0	
rate doubles	2 = 2×	concentration doubles	x = 1	
rate quadruples	4 = 2×	concentration doubles	x = 2	
rate increases x8	8 = 2×	concentration doubles	x = 3	
rate is cut in half	$\frac{1}{2} = 2^{x}$	concentration doubles	x = -1	
rate doesn't change	1 = 3×	concentration triples	x = 0	
rate triples	3 = 3×	concentration triples	x = 1	
rate increases by x9	9 = 3×	concentration triples	x = 2	
rate is cut in thirds	$^{1}/_{3} = 3^{x}$	concentration triples	x = -1	
rate quadruples	4 = 4×	concentration quadruples	x = 1	
	Etcetc	cetc		

Finding Units for k

Remember: $rate = k[A]^{x}[B]^{y} etc \dots$

> Rearrange: $k = \frac{rate}{[A]^{x}[B]^{y} etc...}$

Remember:

 $rate units = \frac{M}{s}$ Concentration unts = M Overall Order = (x + y + etc ...)

Substitute in your units and rewrite: $k = \frac{M/s}{M^{(x+y+etc...)}} \rightarrow k = \frac{M}{M^{(x+y+etc...)*s}} \rightarrow \text{then cancel out units}$

	Units for k based on overall order of reaction				
	$k = \frac{M}{M^{(x+y+etc)\cdot s}}$				
Overall Order	Example of Units Plugged In	Final Units for k			
0	$k = \frac{M}{M^{(0)} \cdot s} \qquad = \frac{M}{1 \cdot s}$	$\frac{M}{s} = Ms^{-1}$			
1	$k = \frac{M}{M^{(1)} \cdot s} \qquad = \frac{M}{M \cdot s}$	$\frac{1}{s} = s^{-1}$			
2	$k = \frac{M}{M^{(2)} \cdot s} = \frac{M}{M \cdot M \cdot s}$	$\frac{1}{M \cdot s} = M^{-1}s^{-1}$			
3	$k = \frac{M}{M^{(3)} \cdot s} \qquad = \frac{M}{M \cdot M \cdot M \cdot s}$	$\frac{1}{M^2 \cdot s} = M^{-2}s^{-1}$			
4	$k = \frac{M}{M^{(4)} \cdot s} = \frac{M}{M \cdot M \cdot M \cdot M \cdot s}$	$\frac{1}{M^3 \cdot s} = M^{-3}s^{-1}$			
	Etcetc				

Remember: $M = \frac{mol}{L}$ $\frac{1}{M} = M^{-1} = \frac{L}{mol}$

You may see this substituted into k units.

For example:
$$M^{-1}s^{-1} = \frac{L}{mol \cdot s}$$

Reference Sheets for Unit #13 – Equilibrium

<u>Equilibrium</u>

Equilibrium is a condition that occurs when a chemical reaction is reversible, and the forward and reverse reactions occur simultaneously, at the same rate. Chemical reactions can be classified into one of two broad categories: those reactions that "go to completion" and those reactions that establish equilibrium".

Burning methane in oxygen to form carbon dioxide and water is a "goes to completion" reaction, and is indicated using a single---headed reaction arrow: $CH_4 + 2O_2 \rightarrow CO_2 + 2H_2O$

Equilibrium reactions do not go to completion; instead, the two reactions (arbitrarily labeled "forward" and "reverse") occur simultaneously. The forward and reverse reactions are opposite; the reverse reaction is the forward reaction written backwards. For example, nitrogen gas and hydrogen gas react to form ammonia: $N_{2(g)} + 3H_{2(g)} \rightarrow 2NH_{3(g)}$

This reaction requires pressures between 2100 and 3600 psi and temperatures between 300 and 550 °C. At these pressures and temperatures ammonia spontaneously decomposes into nitrogen and hydrogen gas: $2NH_{3(g)} \rightarrow N_{2(g)} + 3H_{2(g)}$

Eventually, these two reactions occur at the same rate. When this happens, the concentrations of nitrogen, hydrogen, and ammonia become constant, and the system is said to be "at equilibrium". Typically, we represent chemical equilibrium using a double-headed reaction arrow:

$$N_{2(g)} + 3H_{2(g)} \longleftrightarrow 2NH_{3(g)}$$

Chemical Equilibrium, the Equilibrium Constant (Keq), and the Reaction Quotient (Q)

As a reaction proceeds in a closed system, the concentrations of reactants and products change until, at **equilibrium**, the <u>rate</u> of the forward reaction is equal to the <u>rate</u> of the reverse reaction. At that point there is **no net change** in the concentrations of reactants or products. For example, for the simple reaction: A = B:



Regardless of the initial concentrations, the concentrations of reactants and products present **at equilibrium** are related through the equilibrium constant (K_{eq}).

If all reactants and products are in solution for the reaction: $aA + bB \stackrel{\Rightarrow}{\Rightarrow} cC + dD$

$$K_{eq} = K_{c} = \frac{[C]_{eq}^{c}[D]_{eq}^{d}}{[A]_{ea}^{a}[B]_{eq}^{b}}$$

The subscript "eq" indicates "equilibrium" concentration. The subscript "c" stands for "concentration". The exponents are the coefficients from the balanced equation.

If all reactants and products are ideal gases then **either** partial pressures or concentrations can be used in the equilibrium constant expression. You can convert back and forth from Kc to Kp. In terms of partial pressures:

$$K_{eq} = K_p = \frac{(P_C)_{eq}^c (P_D)_{eq}^d}{(P_A)_{eq}^a (P_B)_{eq}^b}$$

The subscript "p" stands for "pressure"

Important points about equilibrium constants

- 1. Every reaction has a characteristic equilibrium constant that <u>depends only on temperature</u>. The value of the equilibrium constant is independent of the initial concentrations of reactants and products.
- 2. Equilibrium constants have no units. This is because each concentration or pressure is actually a <u>ratio</u> of the concentration or pressure to their <u>standard</u> values. Standard concentration for a substance in solution is 1 M and standard pressure for a gas is 1 atm. This portion of the equation that we don't see in our level of chemistry allows the units to cancel out.
- 3. The concentrations of pure solids or pure liquids do not appear explicitly in the equilibrium constant expression. The concentrations of pure solids and liquids are constant (density divided by molar mass). These constant concentrations are simply incorporated into the equilibrium constant. Their concentrations do not change so they do not mathematically affect the equilibrium expression.
- 4. All reactants and products must be present at equilibrium. Even though pure solids and liquids do not appear in an equilibrium constant expression, if they participate in a reaction they must be present to establish equilibrium. You still need them for the reaction to occur, they just are not going to affect the position of equilibrium.
- 5. The numerical value of the equilibrium constant expresses the tendency for reactants to be converted to products. If K_{eq} >> 1 we say that "products are favored". If K_{eq} << 1 we say that "reactants are favored". A very large equilibrium constant (>1000) means that the reaction will nearly go to completion. A very small equilibrium constant (<0.001) means very little product will be present at equilibrium (or that the reverse reaction goes to completion).
- 6. Changing the initial concentrations of reactants or products does NOT change the equilibrium constant. The equilibrium concentrations of reactants and products will change but K_{eq} will not change. The value of K_{eq} depends only on the identity of the reactants and products and on temperature. Every reaction has a unique K_{eq} at a fixed temperature.
- 7. A catalyst does NOT change the equilibrium constant. A catalyst increases the rates of the forward and reverse reactions by the same factor.
- 8. Changing the temperature DOES change the equilibrium constant. Increasing the temperature increases the equilibrium constant for an endothermic reaction and decreases the equilibrium constant for an exothermic reaction. Increasing the temperature shifts equilibrium in the endothermic direction.
- 9. The equilibrium constant for a reaction in the reverse direction is the reciprocal of the equilibrium constant in the forward direction. K' = 1/K, when K is for the forward direction and K' is for the backwards direction.
- 10. Equilibrium constants are independent of mechanism! (Unlike rate laws!) At equilibrium <u>every</u> step is at equilibrium. We are concerned about the place we end up equilibrium, we don't care how the reaction gets to that point. In kinetics the rate is determined by the way we get to our end point.

Reaction quotient (Q)

At any point during a reaction, if we know the concentrations of reactants and products, we can calculate the **reaction quotient** (**Q**).

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

notice that the concentrations are NOT necessarily equilibrium concentrations

Make sure you understand the difference between Q and Keq:

Q tells you how far a reaction is from equilibrium. K_{eq} tells you where the reaction is trying to go.

By comparing the values of Q and K_{eq} , you can determine whether or not the reaction is at equilibrium. If the reaction is not at equilibrium, you can decide whether the (net) reaction is going in the forward or reverse direction.

If $\mathbf{Q} = \mathbf{K}_{eq}$, then rate_{forward} = rate_{reverse}, the **reaction is at equilibrium** and no net change If $\mathbf{Q} < \mathbf{K}_{eq}$ then rate_{forward} > rate_{reverse}, reaction will **proceed to the right** (\rightarrow more products) If $\mathbf{Q} > \mathbf{K}_{eq}$, then rate_{forward} < rate_{reverse}, reaction will **proceed to the left** (\leftarrow more reactants)

Le Chatelier's Principle

Arguably, the single most important rule governing chemical equilibrium is Le Chatelier's Principle, promulgated by French chemist Henry Louis Le Chatelier (1850 – 1936):

"If a chemical system at equilibrium experiences a change in concentration, temperature or total pressure, the equilibrium will shift in order to minimize that change."

Consider the following equilibrium:

 $N_{2(g)} + 3H_{2(g)} \leftrightarrow 2NH_{3(g)} + Heat$

Any change in concentration, temperature, or total pressure can affect this equilibrium by increasing or decreasing the relative amounts of "reactants" and "products". When the equilibrium is disturbed, it re-establishes a <u>new</u> equilibrium. Since equilibrium point is simply a ratio of products to reactants where the rate forward equals the rate backwards, there are infinite ratios that can reach an equilibrium point. You will never quite get back to where you started, but it will find a new equilibrium point. The easiest way to think about which way a reaction will shift is:

If you add something, use it up – move away from it. If you take something away, make more of it – move towards it.

You can treat heat the same way you would one of the chemicals – if you raise the temperature you have added more heat and you need to use it up. If you lower the temperature you have taken away some heat and you need to make more of it.

A couple examples for the equation above. See your notes for more detailed explan	ations as always! 🔅	9
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Add N ₂	Shift Right	[N ₂] slight \uparrow	$[H_2]\downarrow$	[NH ₃] ↑	Temp. ↑	
Take away H ₂	Shift Left	[N ₂] ↑	[H ₂] slight \downarrow	[NH ₃]↓	Temp. ↓	Kemember –
Take away NH ₃	Shift Right	$[N_2] \downarrow$	$[H_2]\downarrow$	[NH ₃] slight ↓	Temp. ↑	solias, liquias,
Raise the Temp	Shift Left	[N ₂] ↑	$[H_2]\uparrow$	[NH ₃] ↓	Temp. slight ↑	ineri gases, ana
Lower Pressure	Shift Left	[N ₂] ↑	[H ₂] ↑	[NH ₃]↓	Temp. ↓	shift equilibrium
Add Argon gas	No change	No change	No change	No change	No change	sniji equilibrium:



ICE Tables

If you know information about the balanced equation, Keq, and initial concentrations of the chemicals before equilibrium is reached, you can determine the concentrations of each chemical once the reaction has reached equilibrium. Due to the fact that there are so many numbers involved, an ICE Table can be very useful. ICE stands for -

Initial – Write down the initial [] of each chemical

Change – Use a variable (x) to represent the change each chemical will undergo. Take into account the coefficients in the balanced eq, and if the chemical is being used up or made. *Example:* $A + B \rightarrow 2C$ $\Delta[A] = -x \Delta[B] = -x \Delta[C] = +2x$

Equilibrium – The initial plus any change that took place.

Example:

A container has a mixture of 1 M carbon monoxide gas, 2 M water vapor, and 1 M carbon dioxide gas. The following reversible reaction takes place $CO + H_2O \leftrightarrow CO_2 + H_2$ This reaction has an equilibrium constant of 0.64. What is the concentration of each chemical once equilibrium is obtained?

Since there is initially no water vapor in the vessel the reaction must begin going from the left to the right. So let the right side be the reactants and the left the product. So now consider the rows of the ICE table.

	CO +	- H ₂ O ←	\rightarrow CO ₂ +	- H ₂
Initial	1	2	1	0
Change	- X	- X	+ x	+ x
Equilibrium	1 - x	2 - x	1 + x	0 + x

Once you have your ice table values, plug them into the equilibrium expression and solve for x.

Note – Sometimes solving for x is straightforward, sometimes you have to use the quadratic formula, and sometimes you can use the "5% Rule" as a shortcut to not have to use the quadratic formula. To use the 5% rule, K < 1 (and a better rule of thumb is K must be 1000x smaller than your initial concentrations, that is more likely to follow the 5% rule)

You must always check to see if $\frac{x}{[initial]} x 100 < 5\%$ in order to use the rule! If not, then solve with quadratic.

Example using 5% rule

	CO +	⊢ H ₂ O ←	\rightarrow CO ₂ -	+ H ₂
Initial	1	2	1	0
Change	- X	- X	+ x	+ x
Equilibrium	1 - x	2 - x	1 + x	0 + x
5% Rule	~ 1	~ 2	~1	х

 $\operatorname{Keq} = \frac{[CO_2][H_2]}{[CO][H_2O]} \quad plug \ numbers/variables \ in \ \rightarrow \quad 0.64 = \frac{[1][x]}{[1][2]} \quad solve \ for \ x \ \rightarrow 1.28 = x$

Check 5% Rule $\rightarrow \frac{1.28}{1} \neq 5\%$ so you cannot use the rule...have to do it the long way unfortunately \otimes

Example finished NOT Using 5% Rule

	CO +	- H ₂ O ←	\rightarrow CO ₂ +	⊢ H ₂
Initial	1	2	1	0
Change	- X	- X	+ x	+ x
Equilibrium	1 - x	2 - x	1 + x	0 + x

 $\operatorname{Keq} = \frac{[CO_2][H_2]}{[CO][H_2O]} \quad plug \ numbers/variables \ in \ \rightarrow \ 0.64 = \frac{[1+x][x]}{[1-x][2-x]} \quad move \ denominator \ to \ the \ other \ side \ by \ multiplying$

→ 0.64(1-x)(2-x) = (1+x)(x) start simplifying by distributing → $(0.64-0.64x)(2-x) = x + x^2$ use "F.O.I.L" →

 $1.28 - 0.64x - 1.28x + 0.64x^2 = x + x^2$ simplify and move everything to one side $\rightarrow 0 = 0.36x^2 + 2.92x - 1.28 \rightarrow 0 = 0.36x^2 + 2.92x - 1.28$

Now you can use the quadratic formula to solve for $x = \frac{-b \pm \sqrt{b^2 - 4ac}}{2a}$ when $(ax^2 + bx + c)$

$$\frac{-2.92 \pm \sqrt{2.92^2 - 4(0.36)(-1.28)}}{2(0.36)}$$
 Our two solutions are x = -8.5, 0.42

The first solution cannot be physically realistic as it would mean that hydrogen would end up with a negative concentration. So 0.42 must be the desired answer for x.

Now that you have x you can finish your ICE Table! If the question asks for moles make sure to change your M into moles by using M = mol/L

	CO +	H₂O ←	\rightarrow CO ₂ -	H H ₂
Initial	1	2	1	0
Change	- X	- X	+ x	+ x
	1 - x	2 - x	1 + x	0 + x
Equilibrium	= 1 - 0.42	= 2 - 0.42	= 1 + 0.42	= 0 + 0.42
	= 0.58 M	= 1.58 M	= 1.42 M	= 0.42 M

<u>Ksp – Solubility Product Constant</u>

The solubility product constant, Ksp, is the equilibrium constant for a solid substance dissolving in an aqueous solution. It represents the level at which a solute dissolves in solution. The more soluble a substance is, the higher the Ksp value it has.

In order to determine the Ksp of a substance you need to write the dissociation reaction first. You write the Ksp equation the same way as a normal Keq - products over reactants, and solids do not factor into the equations.

Do not forget to balance your dissociation reactions! The coefficients are the exponents in your Ksp equations so the balancing matters!

Examples:

$Sn(OH)_2$ (s) \leftrightarrow Sn^{2+} (aq) + $2OH^-$ (aq)	$K_{sp} = [Sn^{2+}][OH^{-}]^{2}$
$Ag_2CrO_4(s) \leftrightarrow 2Ag^+(aq) + CrO_4^{2-}(aq)$	$K_{sp} = [Ag^+]^2 [CrO_4^2]$
$Fe(OH)_3$ (s) \leftrightarrow $Fe^{3+}(aq) + 3OH^-(aq)$	$K_{sp} = [Fe^{3+}] [OH^{-}]^{3}$

Reference Sheets for Unit #14 – Acids and Bases

ACIDS AND BASES REFERENCE SHEET



7 Strong Acids (H ⁺) All other acids are weak		8 Strong Bases (OH ⁻) All other bases are weak	
Hydrochloric acid	HCI	Lithium hydroxide LiOH	
Hydrobromic acid	HBr	Sodium hydroxide NaOH	
Hydroiodic	HI	Potassium hydroxide KOH	
Perchloric acid	HCIO ₄	Rubidium hydroxide RbOH	
Chloric acid	HCIO ₃	Cesium hydroxide CsOH	
Nitric acid	HNO ₃	Calcium hydroxide Ca(OH) ₂	
Sulfuric acid	H_2SO_4	Strontium hydroxide Sr(OH) ₂	
		Barium hydroxide Ba(OH) ₂	

Memorize these 15, ALL ELSE ARE considered WEAK









pH = 14 - pOH



<u>Arrhenius</u>

- Acids make H⁺ ions in aqueous solutions
- Bases make OH⁻ ions in solution

Bronsted-Lowry

- Acids donate protons
- Bases accept protons

<u>Lewis</u>

- Acids accept electron pairs
- Bases donate electron pairs





STRONG ACIDS			
Acid	Formula	Conj. Base	Ka
Perchloric	HCIO ₄	CIO4 ⁻	Very large
Hydriodic	н	ŀ	Very large
Hydrobromic	HBr	Br⁻	Very large
Hydrochloric	HCI	Cl-	Very large
Nitric	HNO3	NO ₃ -	Very large
Sulfuric	H ₂ SO ₄	HSO4 ⁻	Very large
Hydronium ion	H ₃ O ⁺	H ₂ O	1.0

COMMON WEAK ACIDS			
Acid	Formula	Conj.Base	Ka
Iodic	HIO₃	IO ₃ -	1.7 x 10 ⁻¹
Oxalic	$H_2C_2O_4$	$HC_2O_4^-$	5.9 x 10 ⁻²
Sulfurous	H ₂ SO ₃	HSO3 ⁻	1.5 x 10 ⁻²
Phosphoric	H ₃ PO ₄	H ₂ PO ₄ ⁻	7.5 x 10 ⁻³
Citric	$H_3C_6H_5O_7$	$H_2C_6H_5O_7^{-1}$	7.1 x 10 ⁻⁴
Nitrous	HNO ₂	NO ₂ -	4.6 x 10 ⁻⁴
Hydrofluoric	HF	F ⁻	3.5 x 10⁻⁴
Formic	НСООН	HCOO ⁻	1.8 x 10 ⁻⁴
Benzoic	C ₆ H₅COOH	C ₆ H₅COO ⁻	6.5 x 10⁻⁵
Acetic	CH₃COOH	CH₃COO ⁻	1.8 x 10⁻⁵
Carbonic	H ₂ CO ₃	HCO3 ⁻	4.3 x 10 ⁻⁷
Hypochlorous	HCIO	CIO-	3.0 x 10 ⁻⁸
Hydrocyanic	HCN	CN⁻	4.9 x 10 ⁻¹⁰

COMMON WEAK BASES			
Base	Formula	Conj. Acid	Kb
Ammonia	NH₃	NH4 ⁺	1.8 x 10 ⁻⁵
Methylamine	CH ₃ NH ₂	CH₃NH₃⁺	4.38 x 10 ⁻⁴
Ethylamine	$C_2H_5NH_2$	C ₂ H ₅ NH ₃ +	5.6 x 10 ⁻⁴
Diethylamine	(C ₂ H ₅) ₂ NH	$(C_2H_5)_2NH_2^+$	1.3 x 10 ⁻³
Triethylamine	(C ₂ H ₅) ₃ N	(C ₂ H ₅) ₃ NH ⁺	4.0 x 10 ⁻⁴
Hydroxylamine	HONH ₂	HONH ₃ +	1.1 x 10 ⁻⁸
Hydrazine	H ₂ NNH ₂	H ₂ NNH ₃ +	3.0 x 10 ⁻⁶
Aniline	$C_6H_5NH_2$	$C_6H_5NH_3^+$	3.8 x 10 ⁻¹⁰
Pyridine	C₅H₅N	C₅H₅NH⁺	1.7 x 10 ⁻⁹









WEAK ACIDS AND BASES CALCULATIONS

- Dissociation is a reversible reaction!
- So use Equilibrium Expressions, K values, and Ice Tables to find []'s before doing pH type calculations
- Equilibrium Expression still $\frac{Products}{Reactants}$ which will be $\frac{[Dissociated Ions]}{[Undissociated Molecule]}$
- To find pH (or pOH) of something you first have to know the [H₃O₊] (or [OH⁻])
 - For weak acids/bases you need to do the following steps to find those []'s
 - Step 1 ICE Table
 - Step 2 Write a Ka expression (or Kb depending on the problem)
 - Step 3 Solve for x using either quadratic or 5% rule
 - Step 4 put x back into ICE Table to find the actual [] answers
 - Step 5 use your [H₃O⁺] (or [OH⁻]) to find the pH (or pOH)

MONOPROTIC VS. POLYPROTIC – HOW MANY IONS COME OFF?

- Monoprotic acids/bases → only have one H⁺ or OH⁻
- Diprotic acids/bases → have two H⁺ or OH⁻
- Triprotic acids/bases → have three H⁺ or OH⁻
- Strong Bases
 - \circ all OH⁻ come off
 - Take that into account with your stoichiometry when finding the [OH⁻]
 - 1 M Ca(OH)₂ = 2 M of OH⁻ ions
- Strong Acids
 - \circ The first H+ comes off and it would be a normal strong acid type pH calculation
 - No Ka value needed
 - No ICE Table needed.
 - The second/third/etc H⁺ might come off <u>BUT</u>
 - That would be a weak reaction and you would need:
 - Ka value for that second H⁺ coming off
 - Would need to do an ICE table
 - Then add the [H⁺] from the ICE Table calculation to the [H⁺] you found from the first H⁺ coming off.
 - Example: H₂SO₄ → H⁺ + HSO₄⁻
 Only assume one H⁺ comes off unless given Ka value for HSO₄⁻ → H⁺ + SO₄²⁻
- Weak Acids/Bases
 - \circ For the given Ka or Kb value assume only one H⁺/OH⁻ comes off.
 - You would need a second Ka or Kb value to do a second ICE Table for the second H⁺/OH⁻ coming off, and then would need to add your []'s from each ICE Table calculation.

Salts Reference Sheet



Salts Reference Sheet

7 Strong Acids (H ⁺) All other acids are weak		8 Strong Bases (OH ⁻) All other bases are weak
Hydrochloric acid	HCI	Lithium hydroxide LiOH
Hydrobromic acid	HBr	Sodium hydroxide NaOH
Hydroiodic	HI	Potassium hydroxide KOH
Perchloric acid	HCIO ₄	Rubidium hydroxide RbOH
Chloric acid	HCIO ₃	Cesium hydroxide CsOH
Nitric acid	HNO ₃	Calcium hydroxide Ca(OH) ₂
Sulfuric acid	H ₂ SO ₄	Strontium hydroxide Sr(OH) ₂
		Barium hydroxide Ba(OH) ₂

Dougherty Valley High School Chemistry — Weak Acid/Base Reference Sheet Acid Dissociation Constant (K_a) Values for Some Weak Acids

Weak Acid	Chemical Formula	Ka
acetic	HC ₂ H ₃ O ₂	1.8 x 10⁻⁵
arsenic	H ₃ AsO ₄	5.6 x 10 ⁻³
arsenous	HAsO ₂	6 x 10 ⁻¹⁰
ascorbic	$H_2C_6H_6O_6$	8.0 x 10 ⁻⁵
benzoic	C ₆ H ₅ COOH	6.5 x 10 ⁻⁵
boric	H ₃ BO ₃	5.8 x 10 ⁻¹⁰
carbonic	H ₂ CO ₃	4.3 x 10 ⁻⁷
chloroacetic	CH ₂ CICOOH	1.4 x 10 ⁻³
citric	$H_3C_6H_5O_7$	7.4 x 10 ⁻⁴
formic	HCOOH	1.8 x 10 ⁻⁴
hydrazoic	HN ₃	1.9 x 10⁻⁵
hydrocyanic	HCN	4.9 x 10 ⁻¹⁰
hydrofluoric	HF	6.8 x 10 ⁻⁴
hydrosulfuric	H ₂ S	5.7 x 10 ⁻⁸
hypobromous	HBrO	2 x 10 ⁻⁹
hypochlorous	HCIO	3.0 x 10 ⁻⁸
hydrogen peroxide	H ₂ O ₂	2.4 x 10 ⁻¹²
iodic	HIO ₃	1.7 x 10 ⁻¹
malonic	$H_2C_3H_2O_4$	1.5 x 10 ⁻³
nitrous	HNO ₂	4.5 x 10 ⁻⁴
oxalic	$H_2C_2O_4$	5.9 x 10 ⁻²
phosphoric	H ₃ PO ₄	7.5 x 10 ⁻³
selenous	H ₂ SeO ₃	5.3 x 10 ⁻⁹
sulfurous	H ₂ SO ₃	1.7 x 10 ⁻²
tartaric	$H_2C_4H_4O_6$	1.0 x 10 ⁻³

Base Dissociation Constant (K_b) Values for Some Weak Bases

Weak Base	Chemical Formula	Kb
ammonia	NH ₃	1.8 x 10⁵
aniline	C ₆ H ₅ NH ₂	4.3 x 10 ⁻¹⁰
dimethylamine	(CH ₃) ₂ NH	5.4 x 10 ⁻⁴
ethylamine	C ₂ H ₅ NH ₂	6.4 x 10 ⁻⁴
hydrazine	N ₂ H ₄	1.3 x 10 ⁻⁶
hydroxylamine	HONH ₂	1.1 x 10 ⁻⁸
methylamine	CH ₃ NH ₂	4.4 x 10 ⁻⁴
pyridine	C ₅ H ₅ N	1.7 x 10 ⁻⁹
trimethylamine	(CH ₃) ₃ N	6.4 x 10 ⁻⁵