AP Chemistry
Topic Worksheets

Contents

Unit 1

Section 1.01

Topic 1.3 and 1.4 First Year Review

1. Give either the element name or the element symbol for the following.

a.	Sb	Antimony	b.	C	Carbon
c.	Titanium	Ti	d.	Oxygen	O
e.	H	Hydrogen	f.	Zinc	Zn
g.	Mercury	Hg	h.	Gold	Au
i.	Cr	Chromium	j.	Krypton	Kr
k.	He	Helium	l.	Silver	Ag
m.	Xenon	Xe	n.	Platinum	Pt
o.	Li	Lithium	p.	Br	Bromine
q.	Cobalt	Co	r.	Vanadium	V
s.	Y	Yttrium	t.	Al	Aluminum
u.	Argon	Ar	v.	Copper	Cu
w.	Arsenic	As	x.	Calcium	Ca
y.	Ba	Barium	z.	Rn	Radon
aa.	Sulfur	S	bb.	Neon	Ne
cc.	Silicon	Si	dd.	Fr	Francium
ee.	Nickel	Ni	ff.	U	Uranium
gg.	Rb	Rubidium	hh.	Manganese	Mn
ii.	Gallium	Ga	jj.	Pu	Plutonium
kk.	Cadmium	Cd	ll.	Potassium	K
mm.	Se	Selenium	nn.	Ge	Germanium
oo.	Molybdenum	Mo	pp.	Bismuth	Bi

qq.	Lead	Pb	rr.	Sodium	Na
ss.	Sn	Tin	tt.	Osmium	Os
uu.	P	Phosphorus	vv.	F	Fluorine
ww.	B	Boron	xx.	Radium	Ra
yy.	Cs	Cesium	zz.	Cl	Chlorine
aaa.	Tungsten	W	bbb.	Magnesium	Mg
ccc.	Sr	Strontium	ddd.	Be	Beryllium
eee.	Iron	Fe	fff.	N	Nitrogen
ggg.	Scandium	Sc	hhh.	I	Iodine

2. Determine the number of each type of atom for each formula.

a.	NaCl Na = 1 Cl = 1
b.	$CuSO_4$ Cu = 1 S = 1 O = 4
c.	H_2CO_3 H = 2 C = 1 O = 3
d.	$(NH_4)_3P$ N = 3 H = 12 P = 1
e.	$Ca(NO_2)_2$ Ca = 1 N = 2 O = 4
f.	CH_3COCl C = 2 H = 3 O = 1 Cl = 1

3. Determine the molar mass, in g/mol, of the following substances.

   a. Zinc, Zn

$Zn = 65.38$ g/mol

   b. Copper(II) chloride, $CuCl_2$

$Cu = 63.55$

$2 \times Cl = 35.45 \times 2 = 70.9$

135.45 g/mol

   c. Magnesium nitrate, $Mg(NO_3)_2$

$Mg = 24.30$

$2 \times N = 14.01 \times 2 = 28.02$

$6 \times O = 16 \times 6 = 96$

148.32 g/mol

   d. Diethyl ether, $CH_3CH_2OCH_2CH_3$

$4 \times C = 12.01 \times 4 = 48.04$

$10 \times H = 1.008 \times 10 = 10.08$

$1 \times O = 16 \times 1 = 16$

74.12 g/mol

   e. Copper sulfate pentahydrate, $CuSO_4{\cdot}5H_2O$

$Cu = 63.55$

$S = 32.06$

$4 \times O = 16 \times 4 = 64$

$10 \times H = 1.008 \times 10 = 10.08$

$5 \times O = 16 \times 5 = 80$

249.69 g/mol

4. Determine if the following describes an element, a compound, both, or neither.

	Description	Element	Compound	Both	Neither
a.	Found on the periodic table	X			
b.	A pure substance			X	
c.	Can be broken down by chemical means		X		
d.	Can be broken down by physical means				X
e.	Possible different ratio by mass				X
f.	Always has a chemical bond		X		
g.	Only one type of atom			X	

depends on your definition of compound. Could be just element

5. Answer the following questions that deal with pure substances and mixtures.
   a. Does a pure substance always have the same composition? Explain your reasoning.

Yes. A pure substance is always found in the same ratio of elements by mass.

   b. Does a mixture always have the same composition? Explain your reasoning.

No A mixture does not always have the same ratio by mass. A mixture may contain more or less of any substance that makes up the mixture.

   c. A student combusts an unknown solid and finds it is 35% oxygen by mass. The student carries out the same experiment but finds it is 45% oxygen by mass.
      i. What data would support the solid being a pure substance?

There is no data that supports this being a pure substance.

      ii. What data would support the solid being a mixture?

Since the % of oxygen by mass changed from exp 1 to exp 2 the unknown could be a mixture. A mixture does not always have the same composition by mass.

6. Determine if the following is an empirical or molecular formula by placing an “X” in the appropriate box. Then, explain your reasoning.

	Substance	Empirical	Molecular	Reasoning
a.	CH_3	X		smallest ratio of elements
b.	CH_4	X		smallest ratio of elements
c.	C_6H_6		X	not smallest ratio of elements. Can be CH
d.	CH_3OCH_3	X		Smallest ratio
e.	$KMnO_4$	X		Smallest ratio
f.	$C_{12}H_{26}$		X	not Smallest ratio can be C_6H_{13}
g.	$C_{12}H_{25}O$	X		Smallest ratio

7. Convert the following substance from grams to moles.

a. 120. grams of silicon dioxide, SiO_2

MM $_{SiO_2}$

Si : 28.02

2xO : 32

$\frac{60.09\,g}{mol}$

$$\frac{120\,g\ SiO_2 \mid 1\,mol}{\mid 60.09\,g} = 2.00\ mols\ SiO_2$$

b. 1.3 x 10^3 grams of potassium permanganate, $KMnO_4$

MM $_{KMnO_4}$

K : 39.10

Mn : 54.94

4xO : 64

$\frac{158.04\,g}{mol}$

$$\frac{1.3 \times 10^3\,g\ KMnO_4 \mid 1\,mol}{\mid 158.04\,g} = 8.2\ mols\ KMnO_4$$

c. 4.95 x 10^{-4} grams of strontium chlorate, $Sr(ClO_3)_2$

MM $_{Sr(ClO_3)_2}$

Sr : 87.62

2xCl = 70.9

6xO = 96

$\frac{254.52\,g}{mol}$

$$\frac{4.95 \times 10^{-4}\,g\ Sr(ClO_3)_2 \mid 1\,mol}{\mid 254.52\,g} = 1.94 \times 10^{-6}\ mols\ Sr(ClO_3)_2$$

8. Determine what the solid of the following substances would have at their lattice points by placing an “X” in the appropriate box.

Substance	Molecules	Atoms	Ions
H_2O	X		
$C_6H_{12}O_6$	X		
C		X	
NaCl			X
Al		X	
$Al(NO_3)_3$			X

9. Explain how to calculate the empirical formula of a compound when given …
   a. Grams of each element in the compound.

Convet grams to moles. Divide by the smallest # of moles. If needed, multiply all values to get only whole numbers.

   b. Moles of each element in the compound.

Divide by the smallest # of moles. If needed, multiply all values to get only whole numbers.

   c. Percentage of each element in the compound.

Assume 100g of the compound. % convets to grams. Follow procedure from #9a above.

10. Determine the percent composition of every element in each compound.

a. SF_4

MM
S = 32.06
4 x F = 76
108.06

$\frac{S}{SF_4} \times 100 = \frac{32.06}{108.06} \times 100$

= 29.67% S

$\frac{4 \times F}{SF_4} \times 100 = \frac{76}{108.06} \times 100$

= 70.33% F

b. SF_6

MM
S = 32.06
6 x F = 114
146.06

$= \frac{32.06}{146.06} \times 100$

= 21.95% S

$= \frac{114}{146.06} \times 100$

= 78.05% F

c. CO_2

MM
C = 12.01
2 x O = 32
44.01

$\frac{12.01}{44.01} \times 100 = 27.29\%$ C

$\frac{32}{44.01} \times 100 = 72.71\%$ O

Note: For binary compounds like #10 a-c you can also determine the % of the 2nd element by subtracting the % of element 1 from 100.

ex. 10c) C = 27.29%   O = 100 - 27.29 = 72.71%

11. Perform the following calculations.

a. A 66.0 g sample of a compound contains 36.0 g of C, 6.00 g of H, and 24.0 g of O. Determine the empirical formula.

$\begin{array}{l|l} 36\text{g C} & 1\text{mol} \\ \hline & 12.01\text{g} \end{array} = 3.00\text{ mols C}/1.5 = 2\text{ C}$

C_2H_4O = emp. formula

$\begin{array}{l|l} 6.00\text{g H} & 1\text{mol} \\ \hline & 1.008\text{g} \end{array} = 5.95\text{ mol H}/1.5 = 4\text{ H}$

$\begin{array}{l|l} 24.0\text{g O} & 1\text{mol} \\ \hline & 16\text{g} \end{array} = 1.5\text{ mol O}/1.5 - 1\text{ O}$

b. A compound contains 0.75 moles of K, 0.75 moles of Cr, and 2.6 moles of O. What is the simplest formula of the compound?

$.75\text{ mols K}/.75 = 1\text{ K}$

$KCrO_{3.5}$

$\times 2$

$K_2Cr_2O_7$

$.75\text{ mols Cr}/.75 = 1\text{ Cr}$

$2.6\text{ mols O}/.75 = 3.5\text{ O}$

c. A compound is made of $\underbrace{12.67\%\text{ Al, } 19.73\%\text{ N, and } 67.60\%\text{ O.}}_{\text{assume } 100\text{g}}$ Determine the empirical formula of the compound.

$\begin{array}{l|l} 12.67\text{g Al} & 1\text{mol} \\ \hline & 26.98\text{g} \end{array} = 0.4696\text{ mol Al}/.4696 = 1\text{ Al}$

AlN_3O_9

or

$Al(NO_3)_3$

$\begin{array}{l|l} 19.73\text{g N} & 1\text{mol} \\ \hline & 14.01\text{g} \end{array} = 1.408\text{ mols N}/.4696 = 3\text{ N}$

$\begin{array}{l|l} 67.60\text{g O} & 1\text{mol} \\ \hline & 16.00\text{g} \end{array} = 4.225\text{ mols O}/.4696 = 9\text{ O}$

12. A student is given a mixture of NaCl*(s)* and $NaNO_3$*(s)* and is tasked with determining the percent of NaCl in the mixture. The student dissolves 3.613 g of the mixture in 50 mL of DI water. The student then adds excess $AgNO_3$*(aq)* to precipitate the chloride ion as AgCl*(s)*. The student determines that 2.268 g of AgCl is formed. (Hint: There are several ways to do this problem. In this instance we will only use percent composition by mass and NOT traditional stoichiometry with the balanced equation.)

a. Determine the percent composition by mass of chlorine in AgCl.

$$\frac{Cl}{AgCl} \times 100 = \frac{35.45}{143.32} \times 100 = 24.73\%\ Cl$$

b. Determine the grams of chlorine in the 2.268 g of AgCl.

$$2.268\,g\ AgCl \times 24.73\%\ Cl = 0.5610\,g\ Cl$$

c. All of the chlorine in the AgCl came from the NaCl in the mixture. Determine the number of moles of chlorine.

$$\frac{.5610\,g\ Cl \mid 1\,mol}{\mid 35.45\,g} = 0.01582\ mols\ Cl$$

d. The mole ratio of Na to Cl is 1:1. In other words, for every one mole of Na there is one mole of Cl. Determine the number of moles of sodium.

0.01582 mols Cl = 0.01582 mols Na

e. Determine the mass of sodium. To do so, multiply the moles by the molar mass.

$$\frac{0.01582 \text{ mo's Na}}{} \bigg| \frac{22.99\text{g}}{1 \text{ mol}} = 0.3638\text{g Na}$$

f. Determine the total mass of sodium chloride, NaCl, in the mixture.

Na = 0.3638g

Cl = 0.5610g

$$\begin{array}{lr} \text{Na} & 0.3638\text{g} \\ +\ \text{Cl} & +\ 0.5610\text{g} \\ \hline & 0.9248\text{g NaCl} \end{array}$$

g. Determine the percent by mass of NaCl in the original mixture.

NaCl = 0.9248g

mixture = 3.613g

↳ mixture = NaCl + NaNO_3

$$\frac{\text{NaCl}}{\text{mixture}} \times 100 = \%\ \text{NaCl}$$

$$\frac{0.9248}{3.613} \times 100 = 25.60\%\ \text{NaCl}$$

13. In an experiment, a student is assigned the task of determining the number of moles of water in one mole of the hydrate $Na_2SO_4 \bullet nH_2O$. The student collects the data shown in the following table.

Mass of empty container	22.347 g
Initial mass of sample and container	25.959 g
Mass of sample and container after first heating	24.677 g
Mass of sample and container after second heating	23.941 g
Mass of sample and container after third heating	23.940 g

a. Explain why the sample was heated three times.

The sample is heated 3 times to be certain that all water is out of the hydrate. Since the mass shows little change after heating #3 we can be certain that all water is out of the hydrate.

b. Explain why the student can conclude that all of the water was driven off of the hydrate.

There is very little change in mass between heating #2 and #3. Since the mass did not change significantly no more water was in the compound

c. Use the data above to …

i. Determine the mass of the sample before heating.

$25.959 - 22.347 = 3.612g$ $Na_2SO_4 \bullet nH_2O$

ii. Determine the mass of water in the sample.

$25.959 - 23.940 = 2.019\ g$ H_2O

iii. Determine the moles of water in the sample.

$$2.019\text{g } H_2O \left| \frac{1\text{mol}}{18.016\text{g}} \right. = .01121 \text{ mols } H_2O$$

iv. Determine the mass of anhydrate in the sample.

$3.612 - 2.019 = 1.593\text{g } Na_2SO_4$

↳ the "an hydrate" is the compound w/o water. "an" means without.

v. Determine the moles of anhydrate in the sample.

$$1.593\text{g } Na_2SO_4 \left| \frac{1\text{mol}}{142.05\text{g}} \right. = 0.01121 \text{ mols } Na_2SO_4$$

vi. Determine the formula of the hydrated compound.

$$\frac{.01121 \text{ mols } H_2O}{.01121 \text{ mols } Na_2SO_4} = 10$$

$1\ Na_2SO_4 : 10 \text{ mols } H_2O$

$Na_2SO_4 \cdot 10\ H_2O$

14. In an experiment, a student is assigned the task of determining the number of moles of water in one mole of the hydrate $CuSO_4 \bullet nH_2O$. The student collects the data shown in the following table.

Mass of empty container	22.347 g
Initial mass of sample and container	25.959 g
Mass of sample and container after first heating	25.700 g
Mass of sample and container after second heating	25.046 g
Mass of sample and container after third heating	25.045 g

a. Use the data above to …

i. Determine the mass of water in the sample.

$25.959g - 25.045g = 0.914g\ H_2O$

ii. Determine the moles of water in the sample.

$\frac{0.914g\ H_2O \,|\, 1mol}{|\, 18.016g} = 0.0507\ mols\ H_2O$

iii. Determine the formula of the hydrated compound.

$25.959g - 22.347g$

$3.612g$ hydrate

$3.612g$ hydrate

$- 0.914g\ H_2O$

$2.698g$ anhydrate

$\frac{2.698g\ CuSO_4 \,|\, 1mol}{|\, 159.62g} = 0.0169\ mols\ CuSO_4$

$\frac{0.0507\ mols\ H_2O}{0.0169\ mols\ CuSO_4} = 3 \Rightarrow CuSO_4 \bullet 3H_2O$

b. Determine if the calculated mass of the water would increase, decrease, or remain the same if …

i. while heating the substance some solid spattered out. Explain your reasoning.

The mass of water would increase. If solid spattered out the final mass would be lower than the expected therefore the mass lost would increase. The mass lost is the mass of water.

ii. after heating the hydrate completely it was left out on the counter for an entire day before the final weighing. Explain your reasoning.

The mass of water would decrease. As the sample is left out it absorbs water from the air to replace the water driven off. thus increasing the mass of the final sample. This decreases the mass lost which is the mass of water.

iii. the sample was heated too long and some of the anhydrate vaporized and left the container. Explain your reasoning.

The mass of water would increase. If the anhydrate was vaporized before the last weighing the final mass would be lower than expected therefore the mass lost would increase. The mass lost is the mass of the water.

15. Answer the following questions about a 1.745 g sample of $CaSO_4 \cdot 2H_2O$.

a. What percent of the hydrate is water?

$$\frac{2H_2O}{CaSO_4 \cdot 2H_2O} \times 100 = \frac{2 \times 18.04}{172.182} \times 100 = 20.93\%\ H_2O$$

b. How many grams of water are present in the compound?

$$1.745\,g \times 20.93\% = 0.3652\,g\ H_2O$$

c. The sample is placed in a crucible that weighs 22.35 g. The crucible is heated to constant mass. What would be the mass of the crucible and anhydrate?

$$22.35 + (1.745 - .3652) = 23.73\,g$$

OR

$$22.35 + (1.745 \times .7907) = 23.73\,g$$

16. Answer the following questions about different mixtures of chloride compounds.

   a. A mixture of NaCl and KCl are in a container. The percent of chloride in NaCl is 60.6%. Would the percent of chloride in the mixture be greater than, less than, or equal to the percent of chloride in NaCl. Explain your reasoning.

The percent of chloride would be less than 60.6%. The percent of chloride in KCl is less than the percent of chloride in NaCl.

NaCl = 60.6%   KCl = 47.6%

Therefore the percent of chloride will be between 60.6 and 47.6.

   b. A mixture of NaCl and LiCl are in a container. The percent of chloride in NaCl is 60.6%. Would the percent of chloride in the mixture be greater than, less than, or equal to the percent of chloride in NaCl. Explain your reasoning.

The percent of chloride would be grater than 60.6%. The percent of chloride in LiCl is 83.63%. Therefore the percent of chloride in the mixture will be between 60.6 and 83.63.

17. Determine if the following impurities would increase, decrease, or not change the percent of carbon in a mixture with $C_6H_{12}O_6$, which is about 40% carbon by mass.

   a. Water, H_2O

The impurity would decrease the percent of carbon. With no carbon in the compound the total mass would increase but the mass of carbon would remain the same

   b. Ribose, $C_5H_{10}O_5$

The impurity would not change the percent of carbon because ribose has the same percent of carbon, 40%

   c. Fructose, $C_6H_{12}O_6$ (an isomer of glucose)

The impurity would not change the percent of carbon because fructose and glucose have the same percent of carbon.

   d. Sucrose, $C_{12}H_{22}O_{11}$

The impurity would increase the percent of carbon,

18. Convert the following from either grams to moles or from moles to grams.

a. How many grams are in 1.200 moles of hydrogen cyanide, HCN?

MM HCN

H= 1.008
C = 12.01
N = 14.01
27.028g / 1mol

1.200 mols HCN | 27.028g / 1mol = 32.43 g HCN

b. Determine the number of moles in 3.55 grams of selenium hexafluoride, SeF_6.

MM SeF_6

Se = 78.97
6xF = 19 x 6
192.97g / mol

3.55g SeF_6 | 1mol / 192.97g = 0.0184 mols SeF_6

c. How many grams would be in 0.75 moles of ammonium sulfate, $(NH_4)_2SO_4$?

MM $(NH_4)_2SO_4$

2xN = 14.01 x 2
8xH = 1.008 x 8
S = 32.06
4xO = 64
132.144g / mol

0.75 mols $(NH_4)_2SO_4$ | 132.144g / 1mol = 99 g $(NH_4)_2SO_4$

d. Determine the number of grams in 2.04 moles of antimony pentafluoride, SbF_5.

MM_{SbF_5}

$Sb = 121.76$

$5 \times F = 19 \times 5$

$\frac{216.76g}{mol}$

$$\begin{array}{l|l} 2.04 \text{ mols } SbF_5 & 216.76g \\ \hline & 1mol \end{array} = 442\,g\ SbF_5$$

e. How many moles are in 175 grams of $NH_4[Cr(SCN)_4(NH_3)_2]$?

MM

$N = 14.01$

$4 \times H = 1.008 \times 4$

$Cr = 52.00$

$4 \times S = 32.06 \times 4$

$4 \times C = 12.01 \times 4$

$4 \times N = 14.01 \times 4$

$2 \times N = 14.01 \times 2$

$6 \times H = 1.008 \times 6$

$\frac{336.43g}{mol}$

$$\begin{array}{l|l} 175g\ NH_4[Cr(SCN)_4(NH_3)_2] & 1mol \\ \hline & 336.43g \end{array} = 0.520 \text{ mols}$$

f. Determine the number of moles in 1.45 x 10^{-3} grams of potassium dichromate, $K_2Cr_2O_7$.

MM

$2 \times K = 78.2$

$2 \times Cr = 104$

$7 \times O = 112$

$\frac{294.2g}{mol}$

$$\begin{array}{l|l} 1.45 \times 10^{-3} g\ K_2Cr_2O_7 & 1mol \\ \hline & 294.2g \end{array} = 4.93 \times 10^{-6} \text{ mol } K_2Cr_2O_7$$

19. Calculate the number of atoms in 5.00 g of …
    a. Ca

$$5.00\text{g Ca}\left|\frac{1\text{ mol}}{40\text{g}}\right. = .125\text{ mols Ca}\left|\frac{6.022\times10^{23}\text{ atoms}}{1\text{ mol}}\right. = 7.53\times10^{22}\text{ atoms Ca}$$

    b. N_2

$$5.00\text{g } N_2\left|\frac{1\text{ mol}}{28.02\text{g}}\right. = 0.178\text{ mols } N_2\left|\frac{2\text{ N atoms}}{1 N_2}\right. = 0.357\text{ mol N}\left|\frac{6.022\times10^{23}}{}\right. = 2.15\times10^{23}\text{ atoms N}$$

    c. Ne

$$5.00\text{g Ne}\left|\frac{1\text{ mol}}{20.18\text{g}}\right. = .248\text{ mols Ne}\left|\frac{6.022\times10^{23}}{1\text{ mol}}\right. = 1.50\times10^{23}\text{ atoms Ne}$$

20. Perform the following calculations.
    a. How many grams of Cu are in 0.010 moles of $CuSO_4$

$$.010\text{ mols } CuSO_4\left|\frac{1\text{ Cu}}{1\text{ }CuSO_4}\right. = .010\text{ mols Cu}\left|\frac{63.55\text{g}}{1\text{ mol}}\right. = 0.6355\text{g Cu}$$

    b. How many moles are in 1.80 grams of $C_6H_{12}O_6$ (MM = 180 g/mol)

$$1.80\text{g } C_6H_{12}O_6\left|\frac{1\text{ mol}}{180\text{g}}\right. = 0.010\text{ mols } C_6H_{12}O_6$$

    c. What is the percent composition of Ca in CaF_2?

$$\frac{Ca}{CaF_2}\times100 = \frac{40.08}{78.08}\times100 = 51.3\%\text{ Ca}$$

21. The minimum energy needed to break an oxygen-oxygen bond in ozone is 387 kJ mol $^{-1}$. Determine the amount of energy needed to break 1 oxygen-oxygen bond in ozone.

$$\begin{array}{c|c} 387\text{ kJ} & 1\text{ mol} \\ \hline 1\text{ mol} & 6.022\times10^{23}\text{ atoms} \end{array} = 6.43\times10^{-22}\text{ kJ}$$

22. Hydrogen peroxide (H_2O_2) decomposes to water and oxygen, as shown below.

$$2\ H_2O_2(aq) \rightarrow 2\ H_2O(l) + O_2(g)$$

A small sample of MnO_2 is placed into a beaker of H_2O_2 while it is placed on a balance. The mass is measured over a period of 10 seconds and the data shown below.

Time (sec)	Mass (g)
0	134.45
1	134.20
2	133.95
3	133.70
4	133.45
5	133.20

Time (sec)	Mass (g)
6	132.95
7	132.70
8	132.45
9	132.20
10	131.95

a. Explain why the beaker lost mass.

$O_2(g)$ is being produced. The $O_2(g)$ leaves the beaker.

b. Determine the moles of oxygen created in the reaction.

$$134.45\text{g} - 131.95\text{g} = 2.5\text{g } O_2$$

$$\begin{array}{c|c} 2.5\text{g } O_2 & 1\text{ mol} \\ \hline & 32\text{g} \end{array} = 0.078125\text{ mols } O_2$$

23. Perform the following calculations.

a. How many molecules are in 1.8 g of H_2O?

$$\begin{array}{c|c} 1.8\text{g } H_2O & 1\text{mol} \\ \hline & 18.016\text{g} \end{array} = \begin{array}{c|c} .0999\text{mol } H_2O & 6.022\times10^{23}\text{molecules} \\ \hline & 1\text{mol} \end{array} = 6.02\times10^{22}\text{ molecules } H_2O$$

b. How many molecules are in 3.8 g of C_6H_6?

$$\begin{array}{c|c} 3.8\text{g } C_6H_6 & 1\text{mol} \\ \hline & 78.108\text{g} \end{array} = \begin{array}{c|c} .0487\text{mol } C_6H_6 & 6.022\times10^{23}\text{molecules} \\ \hline & 1\text{mol} \end{array} = 2.93\times10^{22}\text{ molecules } C_6H_6$$

c. Determine the number of oxygen atoms in 1.00 g of $CaCO_3$

$$\begin{array}{c|c} 1.00\text{g } CaCO_3 & 1\text{mol} \\ \hline & 100.09\text{g} \end{array} = \begin{array}{c|c} .00999\text{ mols } CaCO_3 & 3\ O \\ \hline & 1\ CaCO_3 \end{array} = .02997\text{ mols } \textcircled{O}$$

$$\begin{array}{c|c} .02997\text{ mols } O & 6.022\times10^{23}\text{ atoms} \\ \hline & 1\text{mol} \end{array} = 1.80\times10^{22}\text{ atoms } O$$

24. You have a 2.00 g sample of compounds X, Y, and Z. The molar mass (in g mol^{-1}) of X is 50, Y is 35, and Z is 90. Arrange the compounds from smallest number of moles present to largest number of moles present.

$$\frac{2}{90} < \frac{2}{50} < \frac{2}{35}$$

$$Z < X < Y$$

25. Four different metal oxides each have one oxygen. The number of metal atoms in each compound may vary.

a. Does the percent of oxygen in the compound increase, decrease, or remain the same as the molar mass of the compound increases?

The percent of oxygen decreases as the molar mass of the compound increases.

b. Would a compound with a high percent of oxygen produce more or less oxygen than a compound with a low percent of oxygen?

A compound w/ a high percent of oxygen would produce more oxygen than a compound w/ a low percent of oxygen only if we assume equal masses of the compounds,

26. Complete the following table that contains information about various isotopes. Assume all isotopes are atoms and NOT ions.

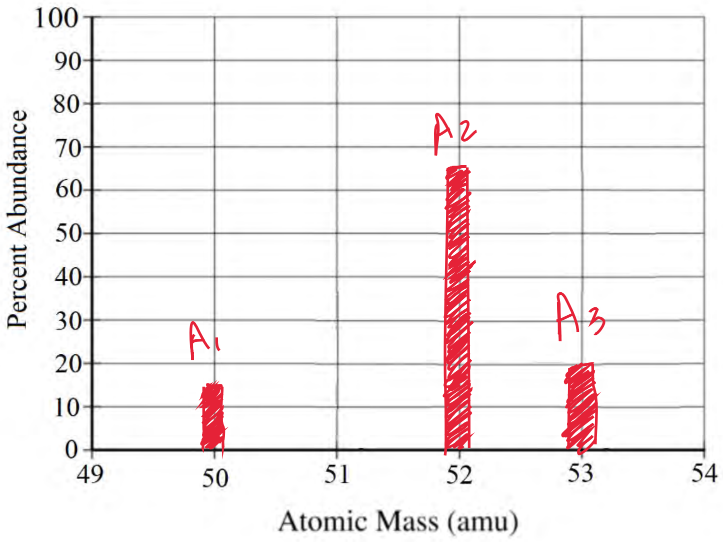
Element Name	Element Symbol	Atomic Number	Mass Number	# of Protons	# of Electrons	# of Neutrons	Nuclear Notation	Hyphen Notation
Boron	B	5	11	5	5	6	${}^{11}_{5}B$	B-11
Boron	B	5	10	5	5	5	${}^{10}_{5}B$	B-10
Krypton	Kr	36	84	36	36	48	${}^{84}_{36}Kr$	Kr-84
Krypton	Kr	36	78	36	36	42	${}^{78}_{36}Kr$	Kr-78
Zirconium	Zr	40	94	40	40	54	${}^{94}_{40}Zr$	Zr-94
Zirconium	Zr	40	96	40	40	56	${}^{96}_{50}Zr$	Zr–96

27. Briefly explain how a mass spectrometer works.

The sample is ionized and propelled down a tube. The ions are bent by a magnet. A heavier particle will bend less than a lighter particle.

28. An element is composed of three stable isotopes, A_1, A_2, and A_3. Isotope A_1 is found 15% of the time, isotope A_2 is found 65% of the time, and isotope A_3 is found 20% of the time. The atomic mass units of A_1, A_2, and A_3 are 50.0, 52.0, and 53.0 respectively.

a. Draw a mass spectrum of the element.



b. Determine the atomic mass of the element.

$$\frac{(50\times15)+(52\times65)+(53\times20)}{100} = \text{atomic mass}$$

51.9 amu

c. The atomic number of fictitious A is 27.

  i. How many protons does A_3 have?

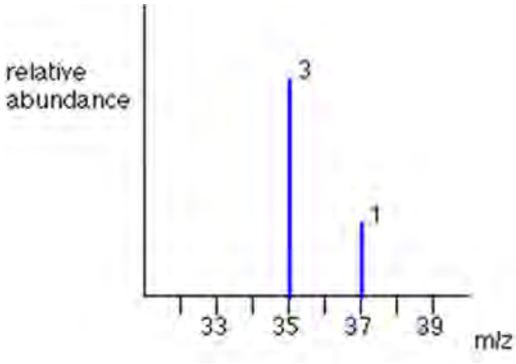
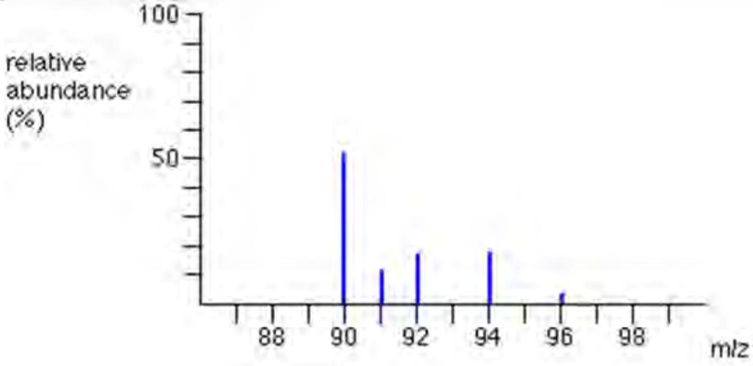
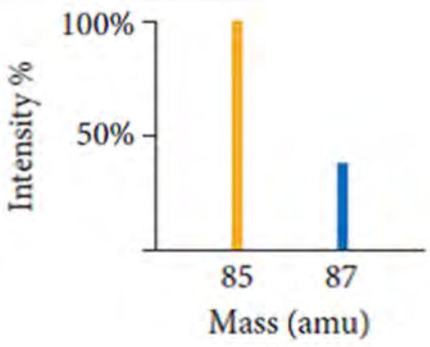
27 protons

  ii. How many neutrons does A_3 have?

A_3 would have 26 protons

(This assumes the mass number of the isotope to be equal to the mass of the isotope)

29. Determine the most likely element for the mass spectrums given below then give your reasoning.

Mass Spectrum	Element	Reasoning
	Bromine Argon [✓ Chlorine] Hafnium	Mass is somewhere b/w 35 and 37; closer to 35. MM of Cl is 35.45
	Tungsten Niobium [✓ Zirconium] Plutonium	W and Pu are way off. Nb has a mass close to 93 which would imply that the 94 isotope would be a higher %
	[✓ Rubidium] Ytterbium Krypton Strontium	The only element w/in the range of masses.

Section 1.04

Topic 1.5 First Year Review

30. Explain why an atom is electrically neutral.

An atom is electrically neutral because an atom has an equal number of p^+ & e^-. p^+ have a relative charge of $+1$ and e^- have a relative charge of -1.

31. An ion is a charged particle and is not electrically neutral.
    a. How does an atom obtain a positive charge to become a cation?

In order to become positively charged the atom must lose an e^-. After losing an e^- there are more p^+ than e^- thus resulting in a positive charge.

    b. How does an atom obtain a negative charge to become an anion?

In order to become negatively charged the atom must gain an e^-. After gaining an e^- there are more e^- than p^+ thus resulting in a negative charge.

32. Will the number of electrons in an atom be greater than, less than, or equal to the number of protons in the atom?

Equal to. Atoms are neutral $\therefore$ $e^- = p^+$

33. Will the number of electrons in a cation be greater than, less than, or equal to the number of protons in the cation?

Less than. A cation is + $\therefore$ $e^- < p^+$

34. Will the number of electrons in an anion be greater than, less than, or equal to the number of protons in the anion?

Greater than. An anion is − $\therefore$ $e^- > p^+$

35. Ozone, O_3, absorbs light with wavelengths from about 100 nm to 900 nm. One of these absorption bands is termed the Hartley band with wavelengths between 200 and 300 nm.

a. Determine the energy of a photon of UV that has a wavelength of 250 nm.

$\lambda = 250\text{nm}$

$E = ?$

$h = 6.626 \times 10^{-34}\ \text{J s}$

$c = 3 \times 10^{17}\ \text{nm/s}$

$$E = \frac{hc}{\lambda}$$

$$E = \frac{(6.626 \times 10^{-34})(3 \times 10^{17})}{250}$$

$$E = 80 \times 10^{-19}\ \text{J}$$

b. Determine the frequency of the same photon of UV light.

$\lambda = 250\text{nm}$

$\nu = ?$

$c = 3 \times 10^{17}\ \frac{\text{nm}}{\text{s}}$

$$c = \lambda \nu$$

$$3 \times 10^{17} = (250)\nu$$

$$1.2 \times 10^{15}\ \text{Hz} = \nu$$

c. Determine the energy for a mole of these photons.

$E = 8.0 \times 10^{-19}\ \frac{\text{J}}{\text{photon}}$

$$\frac{8.0 \times 10^{-19}\ \text{J}}{\text{photon}} \cdot \frac{6.022 \times 10^{23}\ \cancel{\text{photon}}}{1\ \text{mole}} = 480\,000\ \text{J}$$

or

480 kJ

d. The energy to break a bond in O_3 is 101 kJ/mol. Do photons in the Hartley band have enough energy to break one mole of O_3 bonds?

Yes a mole of photons at 250 nm has 480 kJ of energy. The energy required is less than the energy at 250nm.

36. A unit of measurement that is often used in spectral analysis is a wavenumber. A wavenumber is given the symbol $\bar{\nu}$ and has units of cm^{-1}. While wavenumbers are not a unit of energy they are directly proportional to the energy of a wave and inversely proportional to wavelength. Red light, which is considered low energy visible light, has wavenumbers between 13,300 cm^{-1} and 16,100 cm^{-1} while UV light, which is higher energy than red light, has wavenumbers between 25,000 cm^{-1} and 100,000 cm^{-1}

a. Determine the frequency, in Hz, of the red light at 15,000 cm^{-1}. To do so, multiply by the speed of light in cm/s.

$$\frac{15000}{cm}\Bigg|\frac{3\times10^{10}\,cm}{1s} = 4.5\times10^{14}\ Hz \quad or \quad 4.5\times10^{14}\ \frac{cycles}{s}$$

b. Determine the amount of energy, in J, for the same red light.

$E = ?$

$h = 6.626\times10^{-34}\ J\cdot s$

$c = 3\times10^{10}\ cm/s$

$E = h\upsilon$

$E = (6.626\times10^{-34}\ J\ s)(4.5\times10^{14}\ Hz)$

$E = 2.98\times10^{-19}\ J$

c. Determine the frequency, in Hz, of UV light at 39,000 cm^{-1}.

$$\frac{39\,000}{cm}\Bigg|\frac{3\times10^{10}\,cm}{1s} = 1.17\times10^{15}\ Hz$$

d. Determine the amount of energy, in J, for the same UV light.

$E = h\upsilon$

$E = (6.626\times10^{-34})(1.17\times10^{15})$

$E = 7.75\times10^{-19}\ J$

Topic 1.5 Worksheet

37. Determine the charge of each of the following subatomic particles.

Particle	Charge		
Electron	Positive	Neutral	[✓ Negative]
Nucleus	[✓ Positive]	Neutral	Negative
Proton	[✓ Positive]	Neutral	Negative
Neutron	Positive	[✓ Neutral]	Negative

38. According to Coulomb’s Law, does the force of attraction/repulsion increase, decrease, or remain the same when …
    a. the distance between the charges increases?

    decreases

    b. the distance between the charges decreases?

    Increases

    c. the magnitude of the charges increases?

    increases

    d. the magnitude of the charges decreases?

    decreases

39. According to Coulomb's law, which has a greater effect on the force of attraction, increasing the magnitude of the charge of oppositely charged particles or decreasing the distance between the charges? Explain your reasoning.

Decreasing the distance would have a greter effect assuming the changes are proportional. The effect of changing the distance is squared therefore the change is exponential.

40. Define the following terms that pertain to electron configuration:

a. Shell (energy level)

The principal quantum number or. energy of the e^-

b. Subshell (sublevel)

s, p, d, f

c. Core electrons

Inner e^- (not valence)

d. Valence electrons

Outermost s and p e^-

e. Electron configuration

Shows the energy and placement of e^- in an atom

f. Aufbau principle

e^- are placed in the lowest energy subshell available

41. Give the complete electron configuration of the following atoms and ions.

a. Zn

$1s^2\ 2s^2 2p^6\ 3s^2\ 3p^6 4s^2\ 3d^{10}$

b. Zn^{2+}

$1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 3d^{10}$

c. P

$1s^2\ 2s^2\ 2p^4\ 3s^2\ 3p^3$

d. P^{3-}

$1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6$

e. Mg

$1s^2\ 2s^2\ 2p^6\ 3s^2$

f. Mg^{2+}

$1s^2 2s^2 2p^6$

g. Fe

$1s^2\ 2s^2\ 2p^6 3s^2\ 3p^6\ 4s^2\ 3d^6$

h. Fe^{2+}

$1s^2 2s^2\ 2p^6\ 3s^2\ 3p^6\ 3d^6$

i. Fe^{3+}

$1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 3d^5$

42. Give the orbital notation of the valence electrons for the following atoms and ions.

a. Zn

$$\underset{4s}{\underline{\uparrow\downarrow}}\quad \underbrace{\underline{\uparrow\downarrow}\ \underline{\uparrow\downarrow}\ \underline{\uparrow\downarrow}\ \underline{\uparrow\downarrow}\ \underline{\uparrow\downarrow}}_{3d}$$

b. Zn^{2+}

$$\underbrace{\underline{\uparrow\downarrow}\ \underline{\uparrow\downarrow}\ \underline{\uparrow\downarrow}\ \underline{\uparrow\downarrow}\ \underline{\uparrow\downarrow}}_{3d}$$

c. P

$$\underset{3s}{\underline{\uparrow\downarrow}}\quad \underbrace{\underline{\uparrow}\ \underline{\uparrow}\ \underline{\uparrow}}_{3p}$$

d. P^{3-}

$$\underset{3s}{\underline{\uparrow\downarrow}}\quad \underbrace{\underline{\uparrow\downarrow}\ \underline{\uparrow\downarrow}\ \underline{\uparrow\downarrow}}_{3p}$$

e. Mg

$$\underset{3s}{\underline{\uparrow\downarrow}}$$

f. Mg^{2+}

$$\underset{2s}{\underline{\uparrow\downarrow}}\quad \underbrace{\underline{\uparrow\downarrow}\ \underline{\uparrow\downarrow}\ \underline{\uparrow\downarrow}}_{2p}$$

g. Fe

$$\underset{4s}{\underline{\uparrow\downarrow}}\quad \underbrace{\underline{\uparrow\downarrow}\ \underline{\uparrow}\ \underline{\uparrow}\ \underline{\uparrow}\ \underline{\uparrow}}_{3d}$$

h. Fe^{2+}

$$\underbrace{\underline{\uparrow\downarrow}\ \underline{\uparrow}\ \underline{\uparrow}\ \underline{\uparrow}\ \underline{\uparrow}}_{3d}$$

i. Fe^{3+}

$$\underbrace{\underline{\uparrow}\ \underline{\uparrow}\ \underline{\uparrow}\ \underline{\uparrow}\ \underline{\uparrow}}_{3d}$$

43. Give the number of unpaired valence electrons in the particles from #42

	Number of Unpaired Electrons
a.	0
b.	0
c.	3
d.	0
e.	0
f.	0
g.	4
h.	4
i.	5

44. How can you tell if an electron configuration indicates that the atom is in the excited state?

The configuration will not follow Aufbau or Hund's rule and there will be a skip in e^- filling.

45. What happens to the number of protons in the nucleus as you move from left to right across a period on the periodic table?

The # of p^+ increases.

46. An electron is removed from an atom. Would it take more energy, less energy, or the same amount of energy to remove an electron …

  a. closer to the nucleus? Explain your reasoning.

More energy. As the e^- gets closer to the nucleus the force of attraction increases exponentially according to Coulomb's Law.

  b. further from the nucleus? Explain your reasoning.

less energy. As the e^- gets further from the nucleus the force of attraction decreases exponentially according to Coulomb's Law.

47. What does a photoelectron spectrum show …
    a. on the x-axis?

The amount of energy. Typically the highest energy is on the left and the lowest energy is on the right.

    b. on the y-axis?

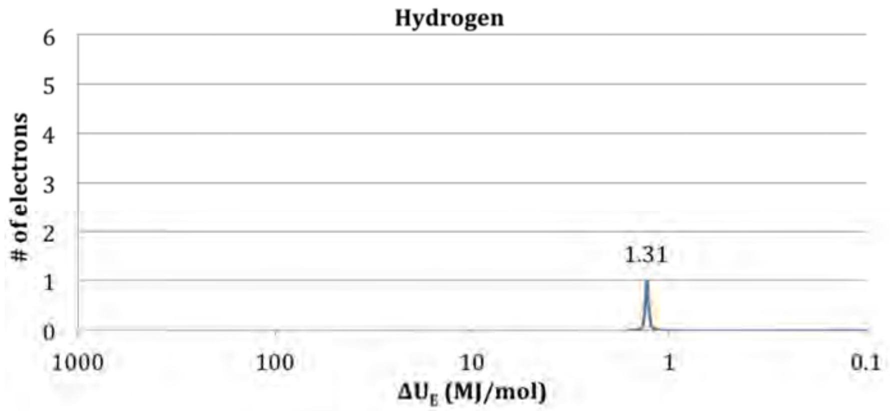
The relative number of e^- at that energy.

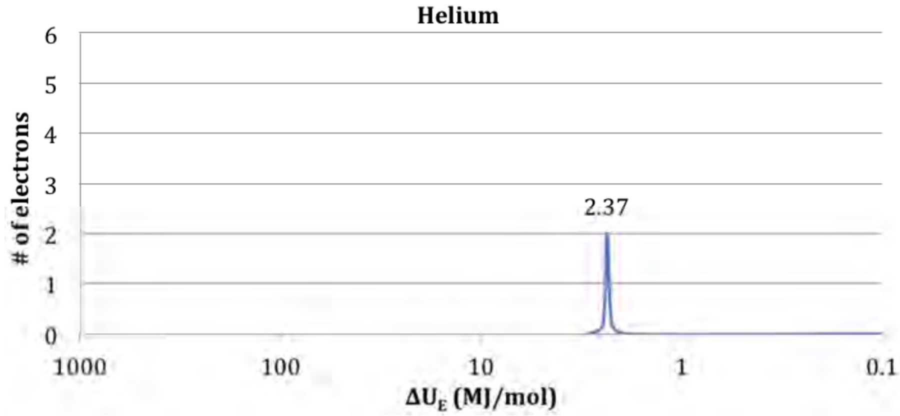
48. What does the location of a peak along the x-axis indicate about the peak on a PES?

The magnitude of energy to remove the e^- in that subshell.

49. What does the height of the peak indicate about the peak on a PES?

The number of e^- in that subshell.





50. Answer the following questions about the PES of hydrogen and helium, shown above.
    a. Why is helium's peak shifted to the left of hydrogen's peak?

Greter nuclear charge means greter force of attraction for the e^-

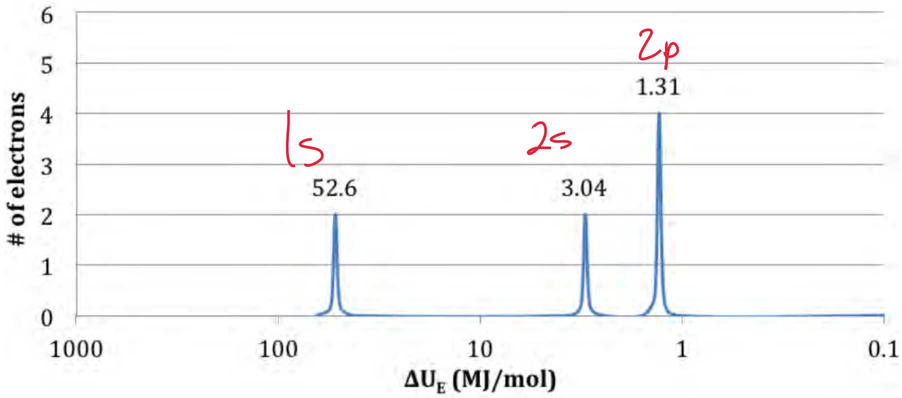
    b. Why is helium's peak higher than hydrogen's peak?

He has $2e^-$ while H has $1e^-$

    c. Would you expect lithium's first peak to be to the left or to the right of helium's first peak? Explain your reasoning.

Li would be to the left of He. Li has a greater nuclear charge than He. Greater nuclear charge means greater force of attraction which would be to the left of He's peak.

51. Use the complete PES shown below to answer the questions that follow.



a. Give the complete electron configuration of the element.

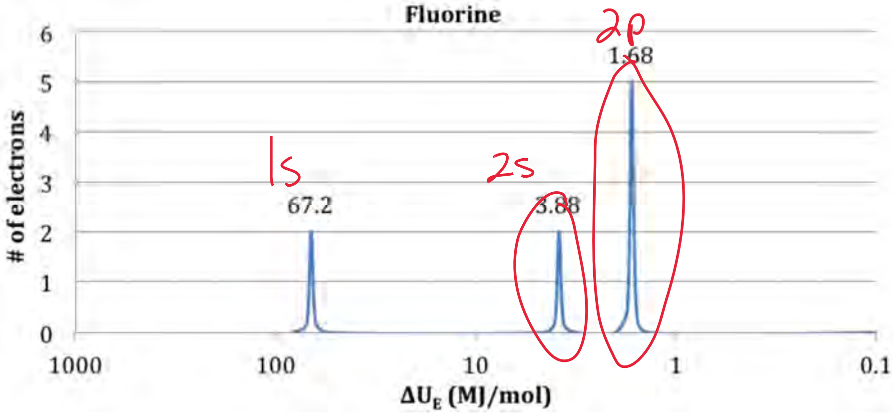
$1s^2 2s^2 2p^4$

b. What is the name of the element shown?

Carbon

c. Label each peak with the shell and subshell designation.

52. Use the complete PES shown below to answer the questions that follow.



a. Label each peak with its shell and subshell designation.

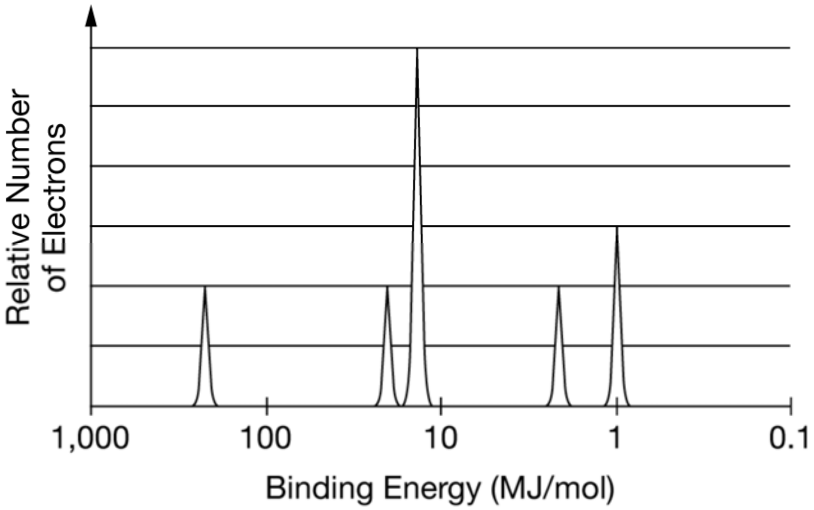
b. Circle the peak(s) of the valence electrons on the PES.

c. Would the next electron added (i.e. the last electron in Ne) be to the left, to the right, or on the peak at 1.68 MJ/mol? Explain your reasoning.

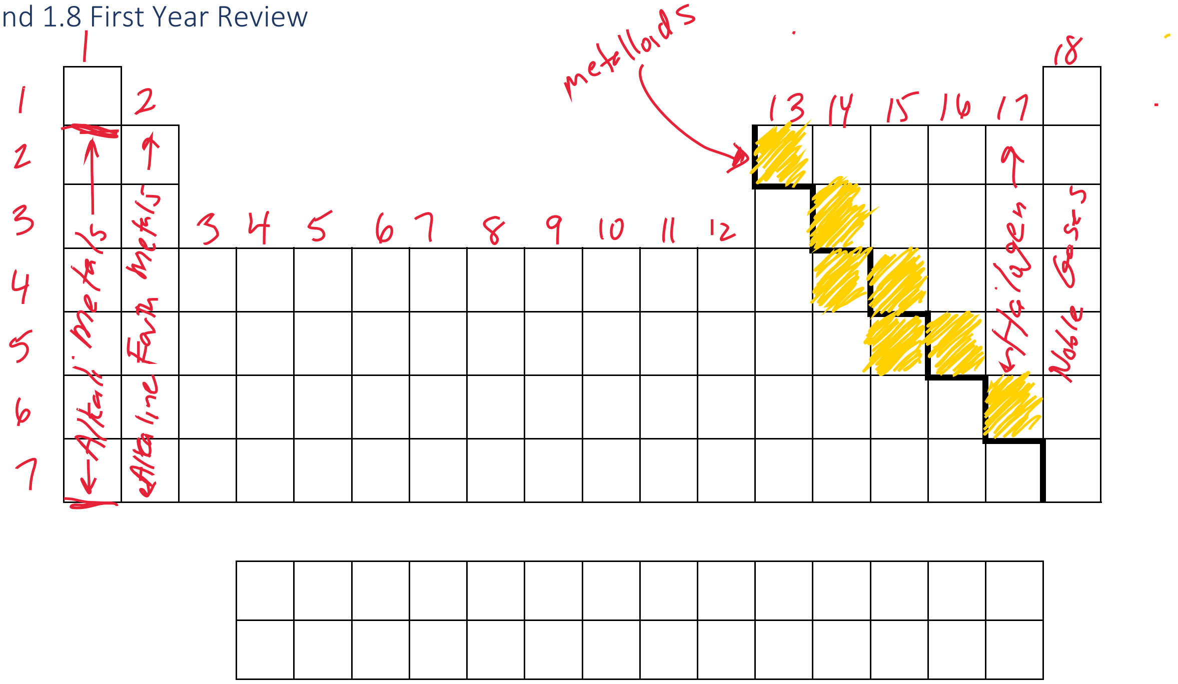
On the peak. The next e^- is still in the 2p subshell.

(The entire peak would be shifted to the left of the peak of F)

53. Give the complete electron configuration for the PES shown below.

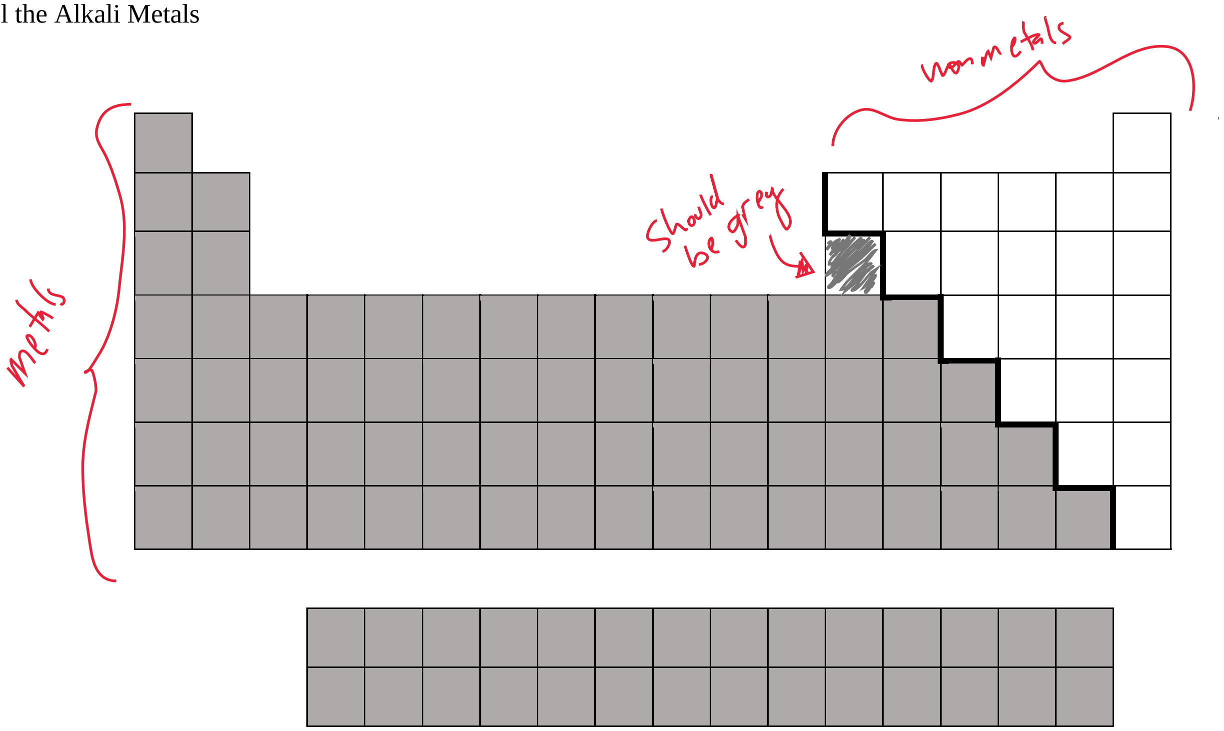


$1s^2 2s^2 2p^6 3s^2 3p^3$

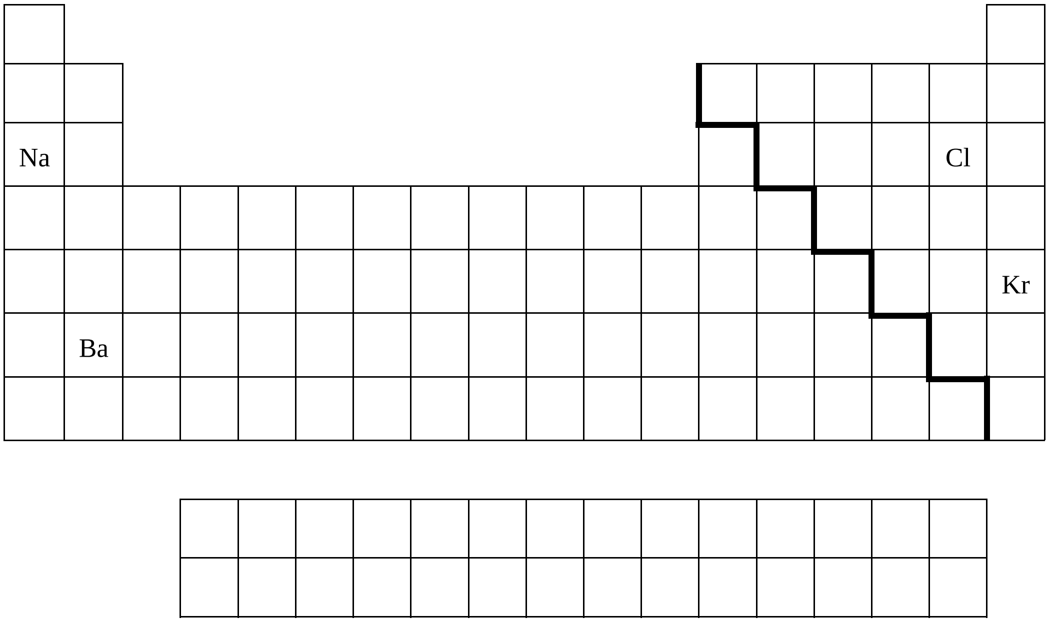


54. On the periodic table above:

Label Periods 1 – 7
Label Groups 1 – 18
Label the Alkali Metals
Label the Alkaline Earth Metals
Label the Halogens
Label the Noble Gases
Label the Metalloids



55. On the periodic table above label the metals and nonmetals.



56. Use the periodic table above to answer the questions below.

Element Symbol	Period Number	Group Number	Group Name	Metal or Nonmetal
Na	3	1	Alkali Metal	Metal
Ba	6	2	Alkaline Earth Metal	Metal
Cl	3	17	Halogen	Non metal
Kr	5	18	Noble Gas	Non metal

57. Using the concept of effective nuclear charge, explain why the size of the atom *generally* decreases when going from left to right across the periodic table.

As you move from left to right across the period, the effective nuclear charge increases. This increase in nuclear charge causes a greter force of attraction which pulls the e^- in closer.

58. Using the concept of electron shells, explain why the size of the atom increases when going from top to bottom of the periodic table.

As you move from top to bottom, the number of e^- shells increases which puts the e^- further from the nucleus

59. Explain why a cation is smaller than the atom it comes from.

A cation is the result of a loss of e^- This results in less e^- e^- repulsion and a smaller atom.

60. Explain why an anion is larger than the atom it comes from.

An anion is the result of a gain of e^- This results in more e^- e^- repulsion and a larger atom.

61. What is ionization energy?

The amount of energy required to remove an e^- from a gaseous atom in the ground state

62. Explain how Coulomb’s law can help explain ionization energy.

The energy needed to remove the e^- is used to overcome the force of attraction between the positive nuclear charge and the e^-.

63. Consider the ionization energy to remove an electron from a neutral atom, X. Would you expect the ionization energy to increase, decrease, or stay the same if …

a. an electron is removed from an identical neutral atom, X? Explain your reasoning.

Remain the same. The ionization energy remains the same for the same e^- in the same type of atom. You are removing one e^- from atom X and then one e^- from a different atom X

b. an electron is removed from an X cation, X^+ ? Explain your reasoning.

The ionization enegy would increase. An ion w/ a +1 charge has less e^- e^- repulsion. With less repulsion the ion is smaller. A smaller particle will have greter attraction between the nucleus and the e^-. Greater attraction requires greater energy.

c. an electron is removed from an X anion, X^-? Explain your reasoning.

The ionization would decrease. As an e^- is added there is greater e^- e^- repulsion. This increases the distance between the nucleus and the e^- thus reducing the force of attraction.

d. an electron is removed from a different, larger atom? Explain your reasoning.

The ionization energy would decrease. As the atom gets bigger the e^- is further from the nucleus thus reducing the force of attraction b/w the e^- and nucleus.

e. an electron is removed from a different atom in the same period with more protons. Explain your reasoning.

The ionization energy would increase. As the nuclear charge increases the force of attraction between the nucleus and the e^- increases. This requires more energy to remove the e^-.

f. an electron is removed from a different atom in the same group with fewer protons. Explain your reasoning.

The ionization energy would increase. In a group, atoms with fewer protons have a small atomic radius because of fewer energy shells. Smaller radius means a greater force of attraction between the nucleus and the e^-.

64. Explain why the first ionization energy of an atom generally increases when going from left to right across the periodic table.

Across the periodic table the nuclear charge increases. As the nuclear charge increases the force of attraction increases and therefore the energy to remove the e^- increases.

65. Explain why the first ionization energy of an atom decreases when going from top to bottom of the periodic table.

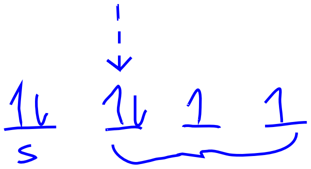
Down a group on the periodic table the atoms increase in the number of energy shells The increasing energy shells are further away from the nuclear charge. The increased distance means less force of attraction and therefore less energy to remove the e^-.

66. Explain why elements B and Al deviate from normal ionization energy trends.

B and Al have an outer np^1 electron that has a slightly higher energy level and expeiences more shielding from the ns^2 subshell. This results in a lower 1st ionization energy than expected.

67. Explain why elements O and S deviate from normal ionization energy trends.

The outermost e^- is in an ns^2np^4 configuration. The paired p electrons experiences greater repulsion and therefore requires less energy and a lower 1st ionization energy than expected



68. Explain why it takes more energy to remove the second electron of an element than the first electron.

The 2nd e^- experiences less e^- e^- repulsion. As a result, the e^- are closer to the nucleus and the force of attraction is greater. A greater force of attraction requires more energy to remove the e^-.

69. The successive ionization energies of Al are shown below.

1st ionization energy	2nd ionization energy	3rd ionization energy	4th ionization energy
578 kJ/mol	1820 kJ/mol	2750 kJ/mol	11,600 kJ/mol

a. Explain why the ionization energy jumps when removing the 4th electron.

The 4th e^- is on a lower energy level that is much closer to the nuclear charge of the nucleus. The closer to the nucleus the greater the force of attraction and therefore the greater the ionization energy.

b. Explain how the ionization energies can be used to determine the number of valence electrons.

The energy increases more when the e^- is removed from an inner shell. The e^- before that big increase are the valence e^-

70. Determine which has a greater ionization energy for removing the next electron. Then, explain your reasoning.

a. Zn or [✓ Zn^{2+}] Same nuclear charge. Zn^{2+} has less e^- e^- repulsion and is therefore smaller. Smaller particle w/ same nuclear charge has a greater force of attraction

b. Fe^{2+} or [✓ Fe^{3+}] Same answer as 70a

c. S^{2-} or [✓ Ar] Same e^- e^- repulsion. Ar has a greater nuclear charge and therefore a greater force of attraction

d. [✓ Ca^{2+}] or Ar Same answer as 70c

71. Explain why it takes more energy to remove the second electron from Na than it does from Mg.

The 2nd e^- of Na is on a lower energy level which is closer to the nuclear charge than Mg 2nd e^- which is on a higher energy level

72. Define electron affinity.

The energy change required to add an e^- to a gaseous atom to form an anion.

73. What is the relationship between atom size and electron affinity?

Generally, the larger the atom the less exothermic

74. What is the relationship between electronegativity and nuclear charge/shielding electrons?

The greater the density of nuclear charge the greater the electronegativity

75. What is the most electronegative element? Lease electronegative?

Fluorine

Cesium

76. Use the ionization energies given below to determine the identity of the second period element. Explain your reasoning.

	Ionization Energy (kJ/mol)
First	801
Second	2,430
Third	3,660
Fourth	25,000
Fifth	32,820

The element is Boron. The 4th e^- removed is from an inner shell (as evidenced by a large jump in I. E.) therefore the atom has 3 valence e^-.

77. Write the equation for the ionization of the following elements.

a. Li

$Li_{(g)} \rightarrow Li^+_{(g)} + e^-$

b. F

$F_{(g)} \rightarrow F^+_{(g)} + e^-$

c. Ar

$Ar_{(g)} \rightarrow Ar^+_{(g)} + e^-$

d. Ni

$Ni_{(g)} \rightarrow Ni^+ + e^-$

78. Explain why Cl^- is smaller than S^{2-}.

Cl^- has a greater nuclear charge. They both have the same # of e^- and thus the same e^- e^- repulsion. But Cl^- has one more p^+ and therefore a greater force of attraction.

1st electron	$X(g) \rightarrow X^{+}(g) + e^{-}$	$IE_1 = 740$ kJ/mol
2nd electron	$X^{+}(g) \rightarrow X^{2+}(g) + e^{-}$	$IE_2 = 1450$ kJ/mol
3rd electron	$X^{2+}(g) \rightarrow X^{3+}(g) + e^{-}$	$IE_3 = 7730$ kJ/mol

79. For the data above, which electron is the closest to the nucleus? Explain your reasoning.

3rd e^-. Since the 3rd e^- takes more energy to remove then the 3rd e^- must be closer to the nucleus.

80. Consider the electron configurations shown below:

Element	Electron Configuration
X	$1s^2 2s^2 2p^3$
Y	$1s^2 2s^2 2p^5$
Z	$1s^2 2s^2 2p^6 3s^1$

a. What is the typical charge of an ion of …

i. Element X

3−

ii. Element Y

1−

iii. Element Z

1+

b. What compound would X form if bonded with …

i. Mg

Mg_3X_2

ii. Li

Li_3X

iii. Fe^{3+}

FeX

iv. Z

Z_3X

c. Name another element that would bond with X similar to …

   i. Mg. Explain your reasoning.

Ca   Same group therefore same properties
b/c same # of valence e^-

   ii. Li

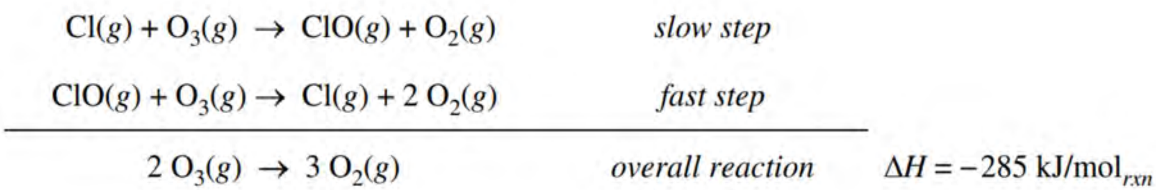
K

   iii. Z

Li

81. Does reactivity increase, decrease, or remain the same as ionization energy decreases?

Reactivity increases as I.E. decreases.



$$Cl(g) + O_3(g) \rightarrow ClO(g) + O_2(g) \quad \textit{slow step}$$

$$ClO(g) + O_3(g) \rightarrow Cl(g) + 2\,O_2(g) \quad \textit{fast step}$$

$$2\,O_3(g) \rightarrow 3\,O_2(g) \quad \textit{overall reaction} \quad \Delta H = -285 \text{ kJ/mol}_{rxn}$$

82. Shown above is the reaction mechanism of ozone, O_3, converted to oxygen, O_2. In the reaction mechanism the element Cl is a catalyst. (More on both of these topics in Unit 5.) Give two other elements that could act as a catalyst in the reaction mechanism. Explain your reasoning.

For Br. They are in the same group and therefore have similar chemical properties

Unit 2

83. What is the general trend for electronegativity across a period?

Electronegativity increases from left to right across a period.

84. What is the general trend for electronegativity down a group?

Electronegativity decreases from top to bottom down a group.

85. Is the difference in electronegative large, small, or nearly zero for …
    a. a nonpolar covalent bond?

    zero or nearly zero

    b. a polar covalent bond?

    Small

    c. an ionic bond?

    large

86. What happens to the electrons in a …
    a. covalent bond?

    e^- are shared

    b. ionic bond?

    one particle gains the e^- while the other particles loses the e^-

87. Define
    a. Electronegativity

The ability to attract an e^- when covalently bonded w/ another atom

    b. Dipole

A positive end and a negative end. Could pertain to a bond or an entire molecule

    c. Dipole moment

A measurement of the polarity of a particle. Measured in debyes (D)

88. How can Coulomb's law be used to predict the electronegativity of an element?

A smaller atom will be more electronegative because of the smaller distance between the nucleus and the other atoms e^-

89. How can the electron configuration of an element be used to predict the electronegativity of an element?

The closer to an ns^2np^6 configuration the more electronegative. The larger the value of "n" the less electronegatie.

90. A nonpolar covalent bond is made of elements with similar electronegativities.

a. Explain why the bond is nonpolar in this instance.

If atoms have similar electronegativities they will have similar attraction for the e⁻. Similar attraction cancels out.

b. Which of the bonds below would be considered nonpolar and which would be polar? Explain your reasoning.

i. Br−Br

nonpolar. Same electronegativity

ii. C−H

nonpolar Similar enough electronegativity to be considered canceling out

iii. I−F

polar Relatively large difference in electronegativity.

91. In a polar bond, which atom is the electron more attracted to, the least electronegative atom or the more electronegative atom? Justify your answer in terms of electronegativity.

The e⁻ is attracted to the more electronegative atom. A greater electronegativity attracts the e⁻ with greater force.

92. Explain how you could determine which bond would be considered more polar when given two different covalent bonds.

The further the atoms are from each other generally the more polar because of a greater difference in electronegativity

93. What types of elements generally make up a …

a. covalent bond?

Nonmetals

b. ionic bond?

Metal + nonmetal

94. What type of bonding occurs in a polyatomic ion?

Covalent bonding

95. What two properties of a bond are illustrated in a graph of potential energy versus bond distance? How are they indicated?

1. Bond Strength
2. Bond Length

96. What does bond order indicate?

Bond strength and the number of bonds
B.O. of 1 = single bond
B.O. of 2 = double bond
B.O. of 3 = triple bond

97. What influences bond length?

The point where repulsion and attraction are balanced.
The size of the atoms bonding.
The number of bonds, i.e. single bond, double bond, triple bond

98. Arrange these bonds from smallest to largest bond length.

a. Cl−Cl, F−F, Br−Br, Cl−Br

F-F < Cl - Cl < Cl - Br < Br - Br

b. C≡C, C−C, C=C

C ≡ C < C = C < C - C

99. For 98.a above, which bond would take the most energy to break? Explain your reasoning.

The F - F bond would take the most energy because the distance between the nuclear charge and other atoms electrons is the shortest therefore the force of attraction is the greatest according to Coulomb's Law.

100. For 98.b above, which bond would take the least energy to break? Explain your reasoning.

The C - C would take the least energy to break. The C - C bond is the longest bond and is therefore the weakest and takes the least amount of energy to break. The distance between the nuclear charge and the electrons is greater and therefore the force of attraction is the least.

101. How can Coulomb’s law be used to explain increasing bond strength of an ionic compound in terms of …

a. the charge of each ion?

As the charge difference increases the bond strength increases. A +1 and -1 ionic bond will be generally weaker than a +2 and -2 ionic bond.

b. the nuclear distance between each ion?

As the nuclear distance decreases the bond strength increases.

102. Does it take more energy, less energy, or the same amount of energy to break apart two ions with …

a. large charge difference compared to small charge difference. Explain your reasoning.

It will take more energy to break apart two ions with a large charge difference compared to a small charge difference. According to Coulomb's Law, as the charges increase the force of attraction increases.

b. large internuclear distance compared to small internuclear distance. Explain your reasoning.

It will take less energy to break apart two ions with a large internuclear distance. According to Coulomb's Law, as the distance between ions increases the force of attraction decreases.

103. Determine which ionic bond would have a greater bond strength. Explain your reasoning for each set.

a. RbCl, NaCl, KCl

NaCl would have a greater bond strength because of a smaller bond distance. They are all +1 ions bonded with a chloride ion, Cl^-. Na^+ is the smallest ion and therefore has a greater attraction.

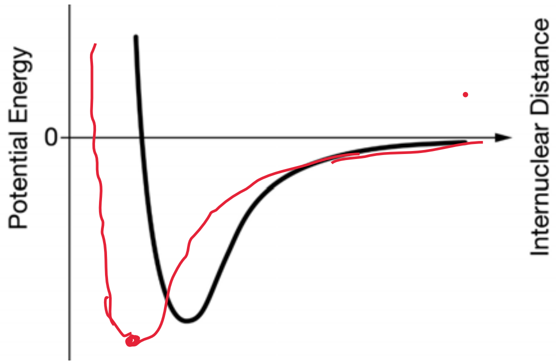
b. MgF_2, NaF, AlF_3

AlF_3 will have a greater bond strength. Al^{3+} is smaller and has a greater charge than Mg^{2+} and Na^+. Therefore the force of attraction is greater.

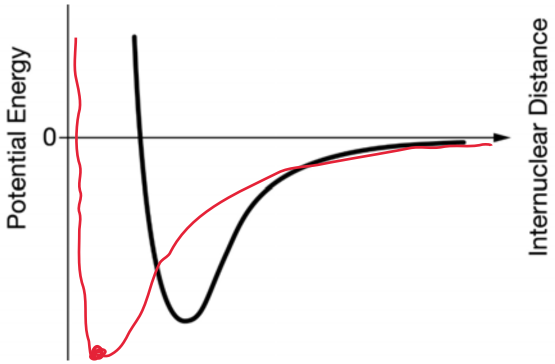
c. KF, KI, KCl

KF would have a greater bond strength because of a smaller bond distance. See answer to 103a for further explanation.

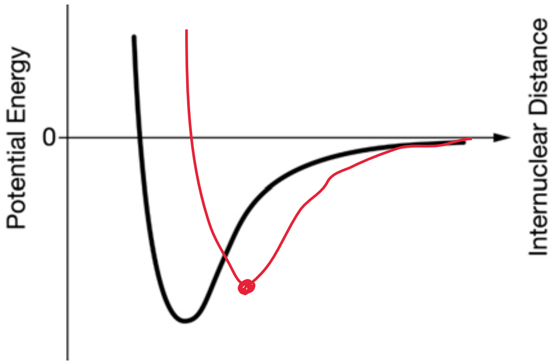
104. Shown below is a graph of potential energy (y-axis) versus internuclear distance (x-axis) for two Br atoms. On each graph, carefully sketch a curve that corresponds to potential energy versus internuclear distance for the atom indicated.



Two Cl atoms

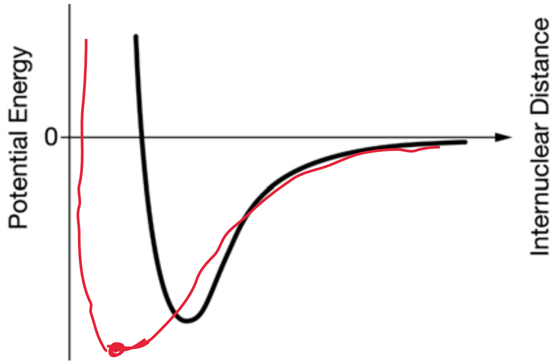


Two F atoms

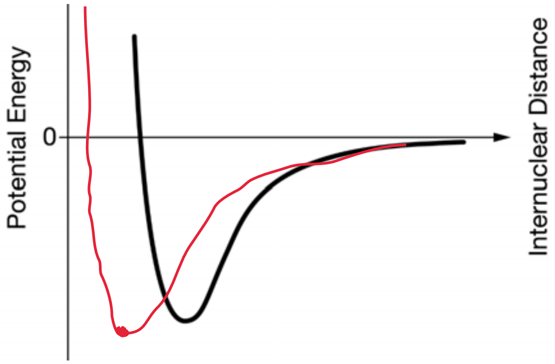


Two I atom

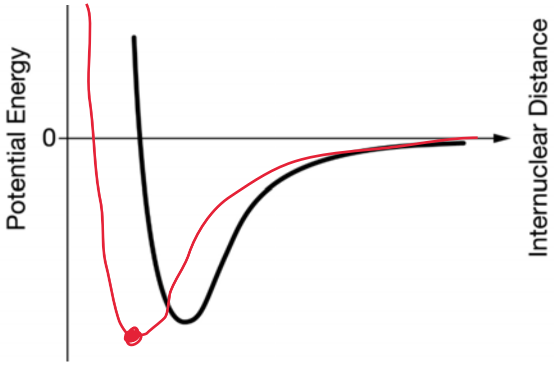
105. Shown below is a graph of potential energy (y-axis) versus internuclear distance (x-axis) for a C–C atoms. On each graph, carefully sketch a curve that corresponds to potential energy versus internuclear distance for the atom indicated.



Two single bonded F atoms

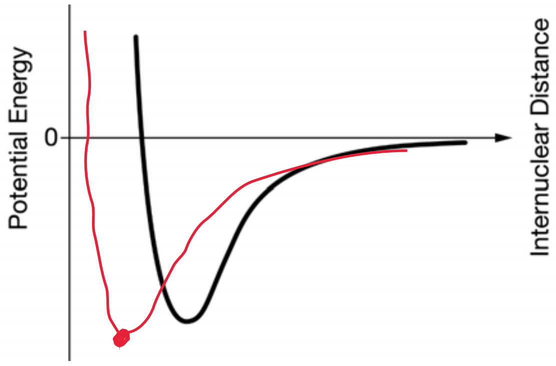


Two single bonded O atoms

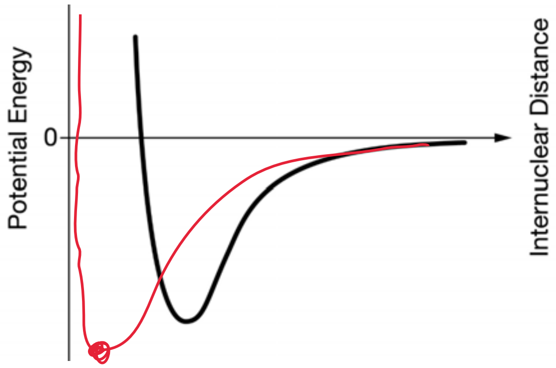


Two single bonded N atoms

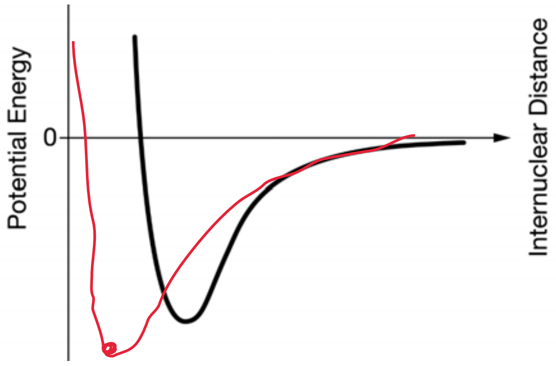
106. Shown below is a graph of potential energy (y-axis) versus internuclear distance (x-axis) for a C–C bond. On each graph, carefully sketch a curve that corresponds to potential energy versus internuclear distance for the bond indicated.



C=C bond



C≡C bond



O=O bond

107. When considering potential energy coordinate graphs as shown above, how does the following effect the potential energy minimum?

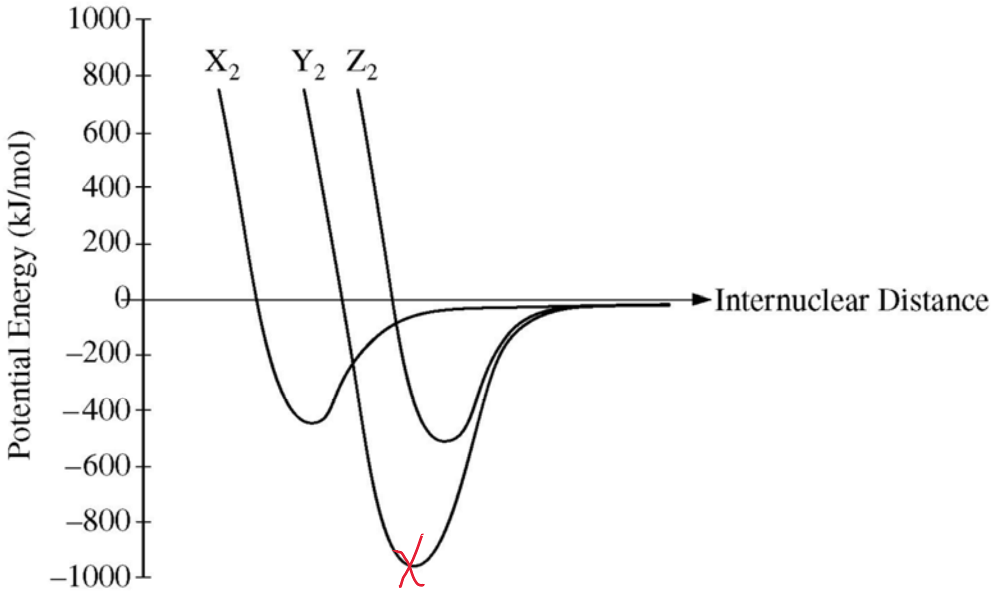
a. Bond length

As bond length increases the potential energy minimum decreases. It takes less energy to break a longer bond than it does a shorter bond.

b. Bond strength

As bond strength increases the potential energy minimum becomes more negative. A stronger bond takes more energy to break than a weaker bond.

108. Answer the questions that follow about the potential energy versus internuclear distance graph shown below.



a. Place an X at the internuclear distance where the bond forms of Y_2.

b. How much energy is released when …
  i. one mole of X bonds to another mole of X?

About 450 kJ/mol

  ii. one atom of X bonds to another atom of X?

$$\frac{450\ \text{kJ}}{\text{mol}} \left| \frac{1\ \text{mol}}{6.022 \times 10^{23}} \right. = 7.5 \times 10^{-22}\ \frac{\text{kJ}}{\text{atom}}$$

c. Which particle, X_2, Y_2, or Z_2, would take the most energy to break the bond? Explain your reasoning.

Y_2 takes the most energy to break. Y_2 has the lowest potential energy minimum (or most negative) and would therefore require the most energy to break.

Compound	Lattice Energy (kJ/mol)
LiF	1030
LiCl	834
LiI	730

Lattice Energy 1

Compound	Lattice Energy (kJ/mol)
NaCl	788
$MgCl_2$	2326
$AlCl_3$	5376

Lattice Energy 2

109. For the lattice energies given above …

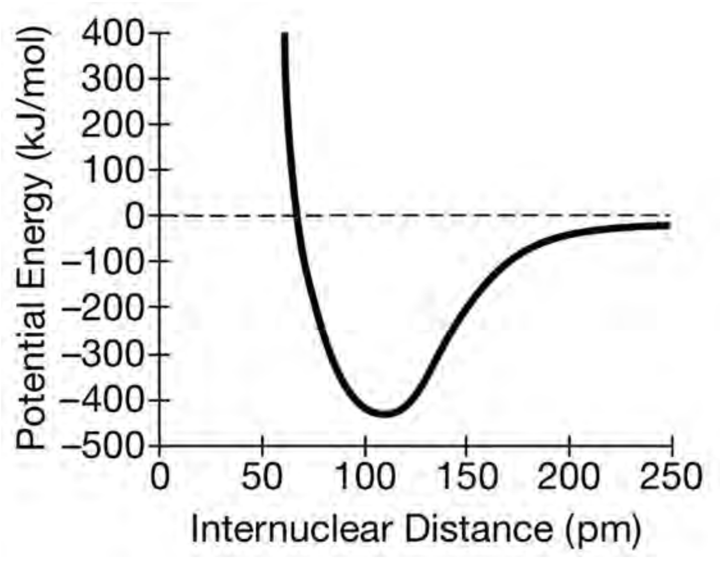
   a. in Lattice Energy 1, explain why LiF has a greater lattice energy than LiI.

Both have Li^+. The F^- is smaller than the I^- therefore the force of attraction is greater. A greater force of attraction takes more energy to break.

   b. in Lattice Energy 2, explain why NaCl has a smaller lattice energy than $AlCl_3$

Both have Cl^-. Al^{3+} has a greater charge and is a smaller ion than Na^+. Both of these increase the force of attraction according to Coulomb's Law. A greater force of attraction takes more energy to break apart.

110.Use the potential energy graph of two atoms as they approach each other to answer the questions that follow.



a. At what internuclear distance do the atoms bond?

About 110 pm.

b. How much energy is released when one mole of the atoms bond?

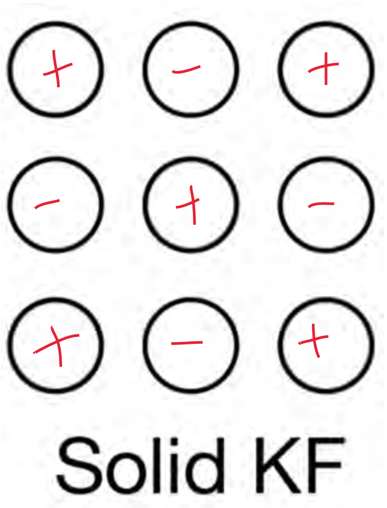
About 450 kJ

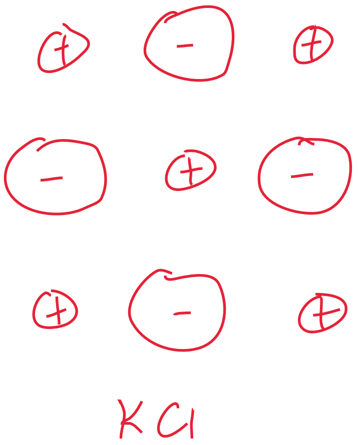
c. How much energy is required to break the bond between one mole of these atoms?

About 450 kJ

d. At what internuclear distance are the attractive and repulsive forces between the two atoms balanced?

About 110 pm





111. The structure above represents the ionic compound KF.

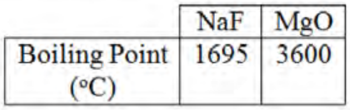
   a. Fill in the circles with + and – to indicate the charge of the ions.

   b. How would the structure differ if it were KCl instead of KF? Draw a particle picture to explain your answer.

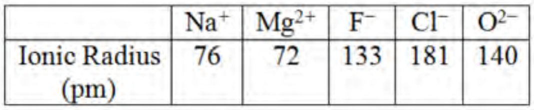
The Cl^- ion would be bigger while the K^+ would be the same size.

   c. Explain why solid ionic compounds are brittle by referring to the particle picture of KF.

When the ions are struck they are displaced/moved and the like charges will line up. This causes great repulsion which allows the solid to cleave.



	NaF	MgO
Boiling Point (°C)	1695	3600



	Na^+	Mg^{2+}	F^-	Cl^-	O^{2-}
Ionic Radius (pm)	76	72	133	181	140

112. Using the data above and Coulomb’s Law, predict whether NaCl would have a higher or lower boiling point than NaF. Explain your reasoning.

NaCl should have a lower boiling point. The Cl^- ion is larger than the F^- ion which increases the distance between charges. An increased distance between charges makes a weaker bond according to Coulomb's Law.

Topic 2.5 First Year Review

113. How many valence electrons do the following elements have?
    a. C 4
    b. N 5
    c. O 6
    d. F 7
    e. Cl 7
    f. H 1
    g. S 6

114. What does it mean when a covalent bond is polar in terms of …
    a. sharing electrons?

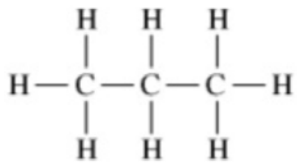
The electrons are found more often around the more electronegative atom. They share the electrons unequally.

    b. difference in electronegativity?

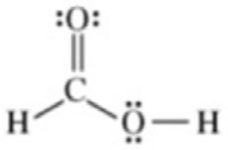
A polar covalent bond will have a difference in electronegativity. (Often reported as a difference in electronegativity greater than 0.4)

    c. electron density?

The electron density will be greater around the more electronegative atom. The more electronegative atom attracts the electrons more therefore the electron cloud will be distorted around that atom.

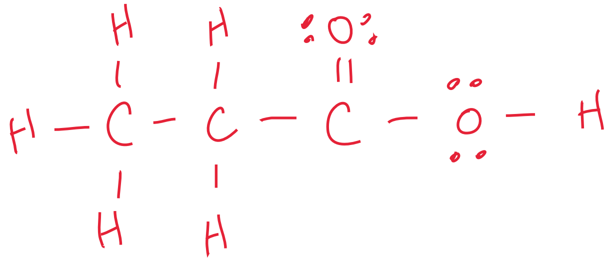


Propane

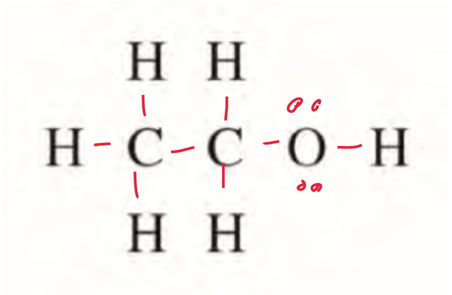


Methanoic (formic) acid

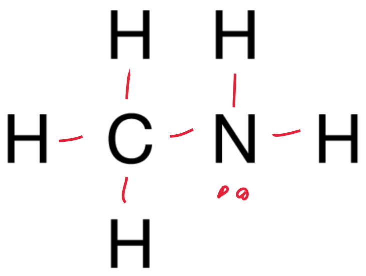
115. Shown above are the Lewis structure of propane (C_3H_8) and methanoic acid (HCOOH). Draw the Lewis structure of propanoic acid, $HC_3H_5O_2$.



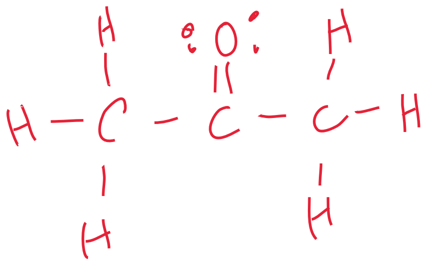
116. Complete the Lewis electron dot structure for ethanol by drawing in all of the electron pairs.



117. Complete the Lewis electron dot structure for the weak base methyl amine by showing all bonding and nonbonding electrons.



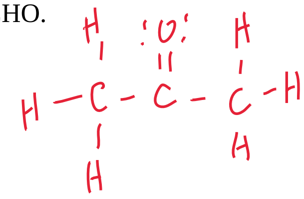
118. Draw the Lewis structure of propanone, CH_3COCH_3.

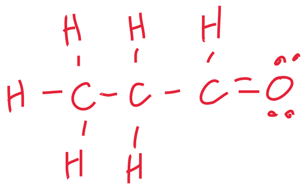


119. What is an isomer?

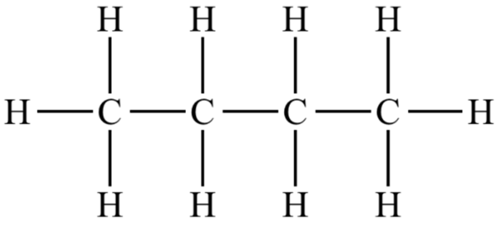
Same number and types of atoms but different bonding.

120. Explain why CH_3COCH_3 is an isomer of CH_3CH_2CHO.

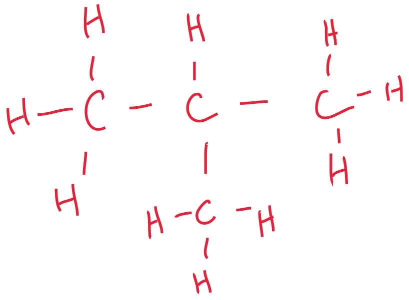




They both have the same number and types of atoms but a different arrangement of the atoms.

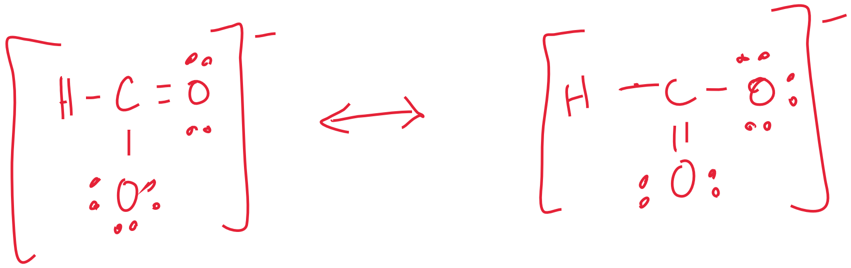


121. Butane is shown above. Draw an isomer of butane.



122. Explain the following observations about the two carbon-oxygen bonds in the methanoate (formate) anion, HCO_2^-.

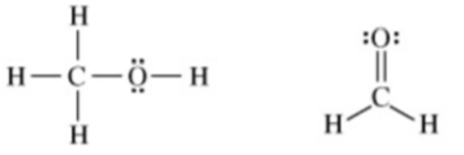
a. Draw a Lewis electron-dot diagram (or diagrams) of the methanoate ion.



b. The two carbon-oxygen bonds in the methanoate (formate) anion, HCO_2^-, have the same length. Explain why.

The methanoate ion exhibits resonance and as a result the double bond on the oxygen rotates. Each oxygen is somewhere between a double bond and a single bond, hence the same length.

c. The length of the carbon-oxygen bonds in the methanoate (formate) anion, HCO_2^-, is intermediate between the length of the carbon-oxygen bond in methanol and the length of the carbon-oxygen bond in methanal, shown below. Explain why.

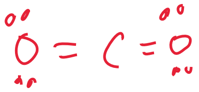
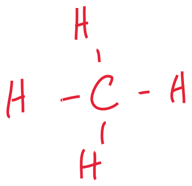
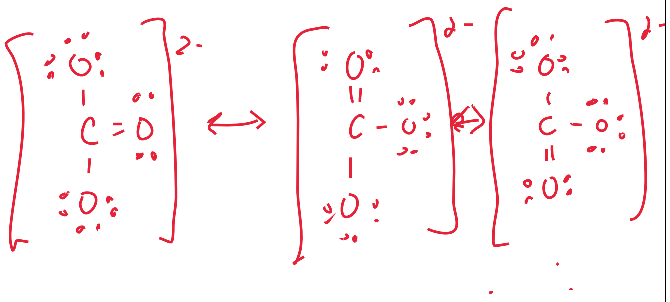
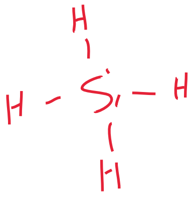
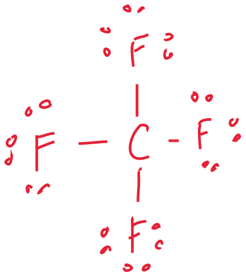
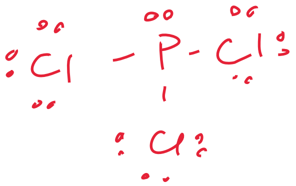
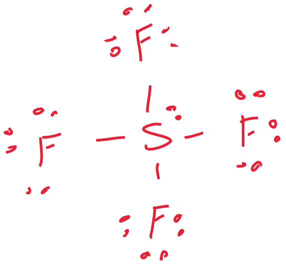
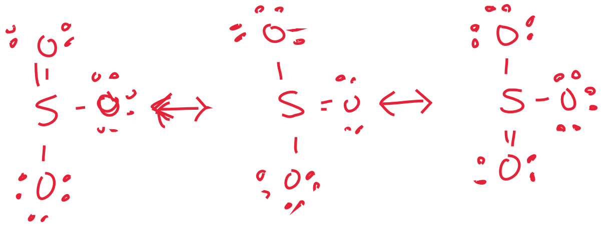
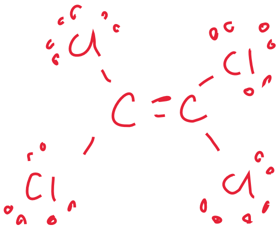
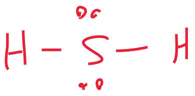


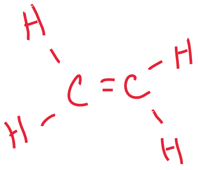
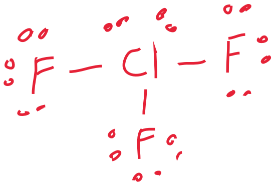
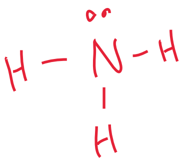
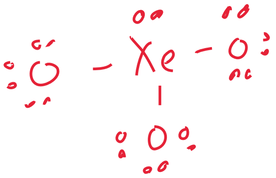
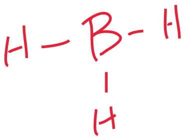
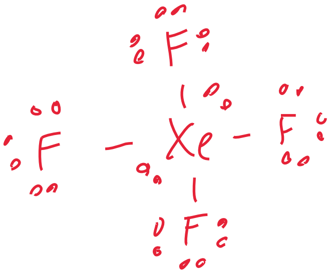
Methanol

Methanal (formaldehyde)

The C-O bond in the methanoate ion is in resonance between a double bond and a single bond. The C-O bond in methanol and methanal do not show resonance and are therefore a single bond length (for methanol) and a double bond length (for methanal).

123.Draw the following Lewis structures.

CO_2 	CH_4 
CO_3^{2-} 	SiH_4 
CF_4 	PCl_3 
SF_4 	SO_3 
C_2Cl_4 	H_2S 

C_2H_4 	HCN 
HF 	ClF_3 
H_2O 	C_2H_2 
NH_3 	XeO_3 
BH_3 	XeF_4 

Section 2.03

Topic 2.7 First Year Review

124. Determine if there is a force of attraction or a force of repulsion between the particles indicated.

a. Electron and proton

Force of attraction. (Proton is positive and electron is negative.)

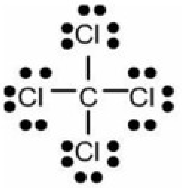
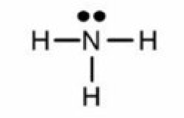
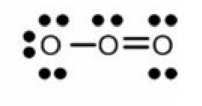
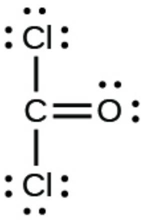
b. Electron and electron

Force of repulsion. (Electrons are negative.)

c. Electron and nucleus

Force of attraction. (Nucleus is positive due to the presence of protons and electron is negative.)

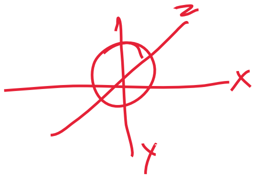
125. Determine the number of bonding and nonbonding electron domains on the central atom of the following Lewis structures.

Lewis structure	Number of bonding domains	Number of Nonbonding domains
	4	0
	3	1
	2	1
	3	0

126. In the quantum mechanical model of the atom electrons can be found in areas known as orbitals.

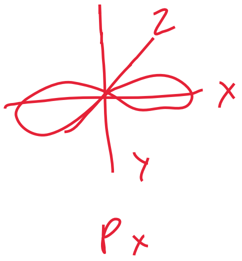
a. How many s-orbital shapes exist on a single energy level? Draw all shapes.

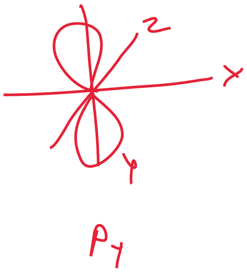
Each energy level has a single s-orbital.

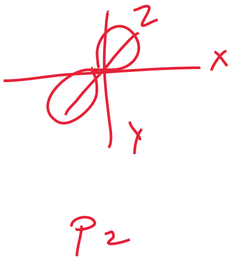


b. How many p-orbital shapes exist on a single energy level? Draw all shapes.

Each energy level has three p-orbitals; p_X, p_Y, and p_Z

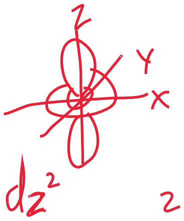


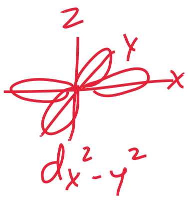


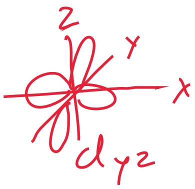


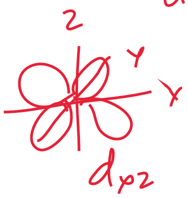
c. How many d-orbital shapes exist on a single energy level? Draw all shapes.

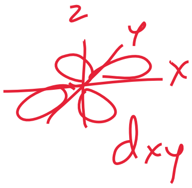
Each energy level has five d-orbitals; d_{z^2}, $d_{x^2-y^2}$, d_{xy}, d_{xz}, d_{yz}











127. Determine the molecular geometry and hybridization of each of the central atoms for the Lewis structures you drew in #123 above.

Particle	Molecular Geometry	Hybridization
CO_2	linear	sp
CO_3^{2-}	trigonal planar	sp^2
CF_4	tetrahedral	sp^3
SF_4	see-saw	sp^3d
C_2Cl_4	trigonal planar	sp^2
C_2H_4	trigonal planar	sp^2
HF	linear	sp^3
H_2O	bent	sp^3
NH_3	trigonal pyramid	sp^3
BH_3	trigonal planar	sp^2

Particle	Molecular Geometry	Hybridization
CH_4	tetrahedral	sp^3
SiH_4	tetrahedral	sp^3
PCl_3	trigonal pyramid	sp^3
SO_3	trigonal planar	sp^2
H_2S	bent	sp^3
HCN	linear	sp
ClF_3	t-shaped	sp^3d
C_2H_2	linear	sp
XeO_3	trigonal pyramid	sp^3
XeF_4	square planar	sp^3d^2

128. The bond angle of CH_4, NH_3, and H_2O decreases. Explain why.

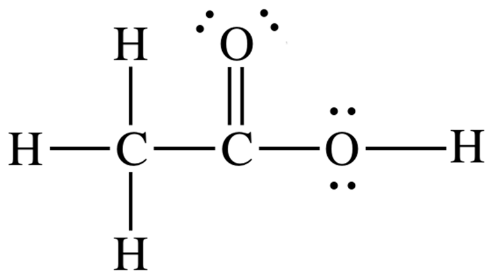
The lone pair electrons on the central atom repels with a greater force as the number of lone pairs increases. CH_4 has no lone pairs, NH_3 has one lone pair, and H_2O has two lone pairs. As the repulsion increases the bonded atoms get closer to each other.

129. Explain why the carbon-oxygen bond length in CO_3^{2-} is greater than the carbon-oxygen bond length in CO_2.

The C-O bond length in CO_2 is that of a double bond length while the bond length in CO_3^{2-} is between two single and one double bond because of resonance. All three C-O bond lengths in CO_3^{2-} are the same length. The C-O bond in CO_3^{2-} is longer because of this resonance.

130. In terms of molecular geometry, account for the fact that the CF_4 molecule is nonpolar, where as the SF_4 molecule is polar.

In CF_4 the dipoles cancel because of the tetrahedral shape. In SF_4 the dipoles do not cancel because of the see-saw shape.



131. Shown above is the Lewis structure of ethanoic acid.

a. What is the approximate angle of the H – O − C bond.

107°. (109.5° is also acceptable.)

b. What is the approximate angle of the O – C = O bond.

120°

c. What is the approximate angle of the H – C – H bond.

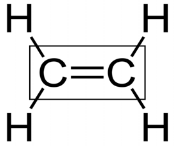
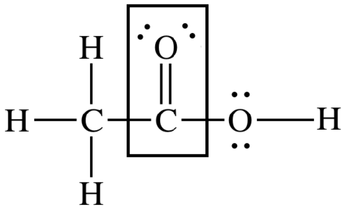
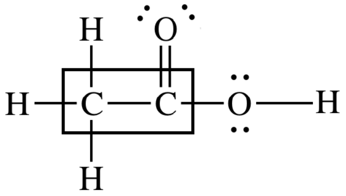
109.5°

132. Complete the table below that shows the relationship between hybridization and bond angles.

Hybridization	Bond Angle	Possible VSEPR Molecular Geometries
sp	180°	linear (180° only)
sp^2	120°	trigonal planar (120°) and bent (<120°)
sp^3	109.5°	tetrahedral (109.5°), trigonal planar (<109.5°), and bent (<<109.5°)

Note: The "Bond Angle" is assuming no lone pairs. The bond angles with lone pairs is shown with the geometry.

133. Determine the number of sigma and pi bonds in the following particles.

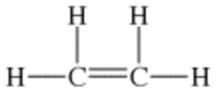
Particle	# of sigma and pi bonds on the carbon indicated
$H-\boxed{C\equiv C}-H$	1 sigma, 2 pi
	1 sigma, 1 pi
	1 sigma, 1 pi
	1 sigma

σ sigma

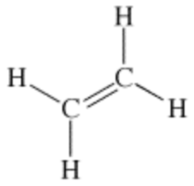
π pi

134. Shown below are two different representations of the same molecule ethene, C_2H_4. Which of the two Lewis structures is a more accurate representation of the molecular shape? Explain your reasoning.

Structure A



Structure B



Structure B is the more accurate representation of the molecular shape. Both structures are good Lewis structures. However, Structure B is a better representation of the approximate angles that the shape would take, trigonal planar around each carbon.

135. Give the reasoning for when a molecule is …

a. polar.

Dipoles do not cancel

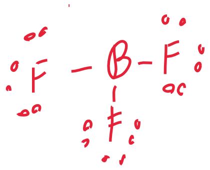
b. nonpolar.

Dipoles cancel
or
No polar bonds present

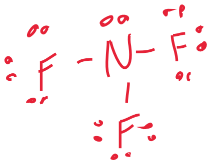
136. Determine if the particle is polar or nonpolar for the Lewis structures you drew in #123.

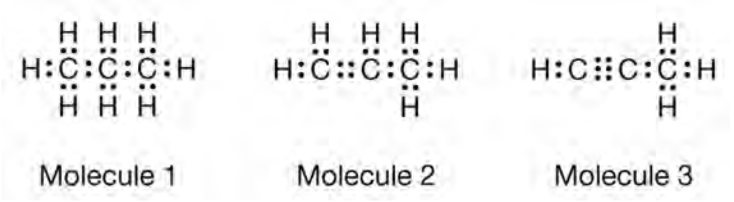
Particle	Polar or Nonpolar	Particle	Polar or Nonpolar
CO_2	NP	CH_4	NP
CO_3^{2-}	NP	SiH_4	NP
CF_4	NP	PCl_3	P
SF_4	P	SO_3	NP
C_2Cl_4	NP	H_2S	P
C_2H_4	NP	HCN	P
HF	P	ClF_3	P
H_2O	P	C_2H_2	NP
NH_3	P	XeO_3	P
BH_3	NP	XeF_4	NP

137. BF_3 is nonpolar while NF_3 is polar. Explain why.



In BF_3 the dipoles cancel because it is trigonal planar. In NF_3 the dipoles do not cancel because it is trigonal pyramid.





138. Use the three molecules above to answer the questions that follow. In each case, refer to all three molecules when asked to explain or justify.

a. Which of these molecules is polar? Explain your reasoning.

They are all nonpolar because they are all made of nonpolar bonds.

b. Which molecule has the strongest carbon – carbon bond? Justify your answer.

Molecule 3. The C-C bond is a triple bond which is the strongest C-C bond. The other molecules have a single bond and a double bond.

c. Which molecule has C – C – H bond angle of approximately 120°? Explain your reasoning.

Molecule 2. The other molecules all have a bond angle of 109.5° because of the tetrahedral structure. The C-C-H bond in molecule 2 is trigonal planar.

d. Determine the hybridization of the left most carbon for each molecule.

Molecule 1 = sp^3
Molecule 2 = sp^2
Molecule 3 = sp

e. Which molecule would have the highest percent of carbon by mass?

Molecule 3. There are more carbons per total molecule ($3xC/C_3H_4$) compared to molecule 2 ($3xC/C_3H_6$) and molecule 3 ($3xC/C_3H_8$).

Section 2.04

Topic 2.6 First Year Review

139. Carbon can bond to carbon with a single, double, or triple bond.
    a. Will a C−C bond be shorter, longer, or the same length as a C=C bond?

A C - C single bond will be longer than a C=C double bond.

    b. Will a C−C bond be shorter, longer, or the same length as a C≡C bond?

A C-C single bond will be longer than a C≡C triple bond.

    c. Which of these three bonds, C−C, C=C, or C≡C, will be the strongest?

The C≡C triple bond is the strongest.

140. How many electrons are involved in a …
    a. Single bond?

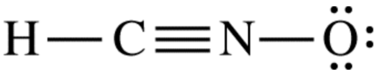
2

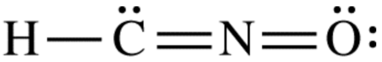
    b. Double bond?

4

    c. Triple bond?

6





141. Use the Lewis structures above to answer the questions that follow.

a. Determine the formal charge of every atom in each of the Lewis structures.

$H = 1 - (\frac{1}{2}(2) + 0) = 0$

$C = 4 - (\frac{1}{2}(8) + 0) = 0$

$N = 5 - (\frac{1}{2}(8) + 0) = +1$

$O = 6 - (\frac{1}{2}(2) + 6) = -1$

$H = 1 - (\frac{1}{2}(2) + 0) = 0$

$C = 4 - (\frac{1}{2}(6) + 2) = -1$

$N = 5 - (\frac{1}{2}(8) + 0) = +1$

$O = 6 - (\frac{1}{2}(4) + 4) = 0$

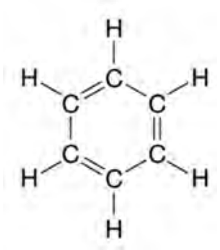
b. Based on formal charge, which of the two molecules is the best representation of HCNO? Justify your answer.

The structure on the left is the best representation. The most electronegative atom (oxygen) has a negative charge while in the structure on the right the most electronegative atom (oxygen) does not have the most negative charge.

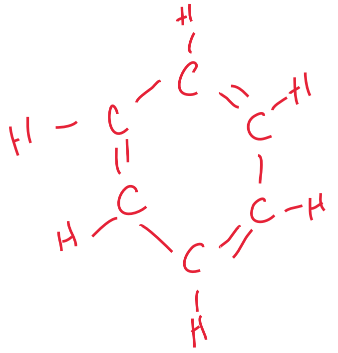
Also, the less electronegative carbon has a formal charge of zero in the structure on the left while the carbon in the structure on the right has the most negative charge.

142. Explain why every bond in CO_3^{2-} is the same length.

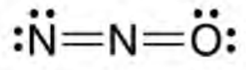
Resonance. The double bond rotates because the electrons are delocalized.



143. Shown above is benzene, C_6H_6. Draw all resonance structures that exist.



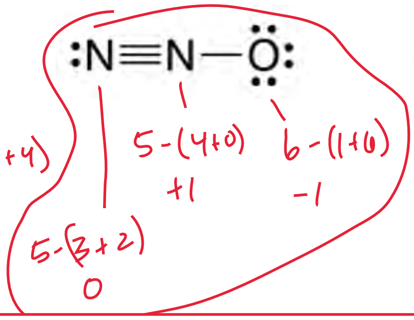
144. Determine which structure below is the best based on formal charge.



$5-(2+4)$ -1

$5-(4+0)$ $+1$

$6-(2+4)$ 0



$5-(3+2)$ 0

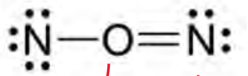
$5-(4+0)$ $+1$

$6-(1+6)$ -1

$6-(2+4)$ 0

$5-(1+6)$ -2

$5-(3+0)$ $+2$



-2 $+3$ -1

The best structure is circled. It has the smallest difference in formal charge and the most electronegative atom has the negative formal charge.

Topic 2.4 First Year Review

145. Will atoms be more similarly sized across a period or down a group? Explain your reasoning.

Atoms are more similarly sized across a period. The atoms across a period all have the same energy level being filled. Down a group the atoms all have different energy levels being filled. A change in energy level is a large difference in size.

146. Define the following terms:
   a. Malleable

Able to be bent of pounded flat

   b. Ductile

Able to be drawn into a thin wire

   c. Sea of electrons

Delocalized electrons. Electrons are free to move about the entire structure and are not constrained to a single atom.

   d. Conductor

A substance that allows electrons to flow in one or more directions.

Topic 2.4 Worksheet

147. What relative size of metallic atoms would make …

a. an interstitial alloy?

A large atom and a small atom

b. a substitutional alloy?

Relatively similar size of atoms

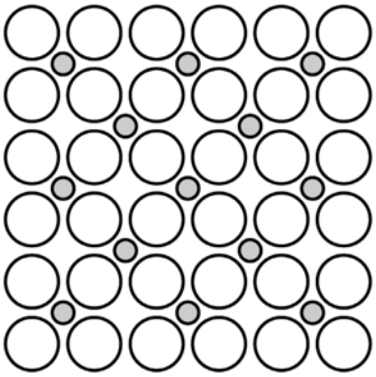


Figure A

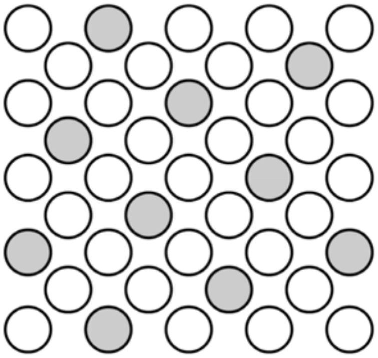


Figure B

148. Use the two figures above to answer the questions that follow.

a. Would Figure A be an interstitial or substitutional alloy? Justify your reasoning.

Interstitial alloy. The smaller atoms fit into the spaces of the larger atoms.

b. Would Figure B be an interstitial or substitutional alloy? Justify your reasoning.

Substitutional alloy. The atoms are replaced by other atoms of similar size.

c. Which figure would represent an alloy made of copper and zinc? Explain your reasoning.

Figure B, substitutional alloy. The atoms are close to each other on the periodic table and are therefore close in atomic size. This makes them a substitutional alloy.

d. Which figure would represent an alloy made of iron and carbon? Explain your reasoning.

Figure A, interstitial alloy. The atoms are far apart from each other on the periodic table and are therefore different in size. This makes them an interstitial alloy.

e. Which figure would be less malleable? Explain your reasoning.

Figure A. The small atoms in between the larger atoms makes them less malleable due to the lack of movement/rigidity.

f. The two alloys are made of the same white element. Which figure would be more dense?

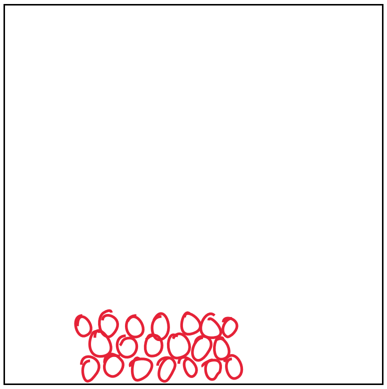
Figure A. There are more atoms in the same amount of space.

Unit 3

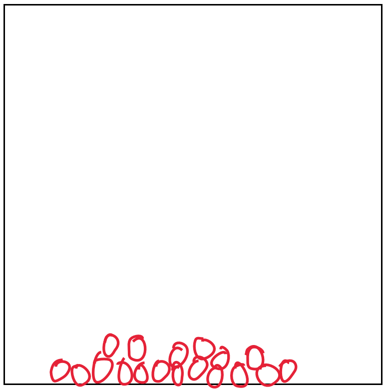
149. Determine if the following is a property of a solid, liquid, or gas by placing an X in the appropriate box. Some properties may describe multiple states of matter.

	Property	Solid	Liquid	Gas
a.	Compressible			X
b.	Definite shape	X		
c.	Definite volume	X	X	
d.	Fills entire container regardless of volume			X
e.	Greatest force of attraction between the particles	X		
f.	Heat of fusion occurs between these two states	X	X	
g.	Heat of vaporization occurs between these two states		X	X
h.	Made of particles	X	X	X
i.	No definite shape		X	X
j.	No definite volume			X
k.	Of the three, packed the loosest			X
l.	Of the three, packed the tightest	X		
m.	Particles are free to move about the most			X
n.	State of substance at temperatures above the melting point but below the boiling point		X	
o.	State of substance at temperatures below the melting point	X		
p.	State of subtsance at temperatures above the boiling point			X
q.	Takes the shape of the container		X	X
r.	Weakest force of attraction between the particles			X

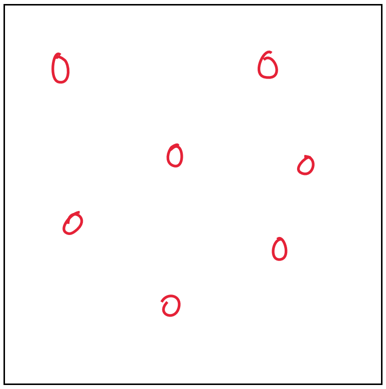
150. Draw a particulate picture of …



Solid

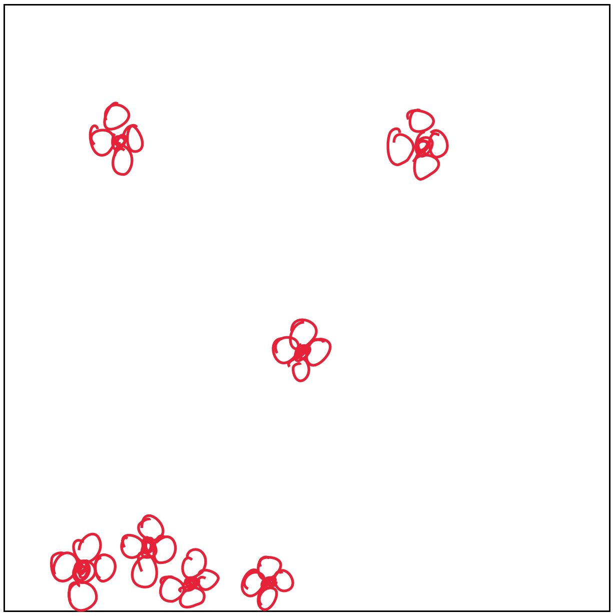


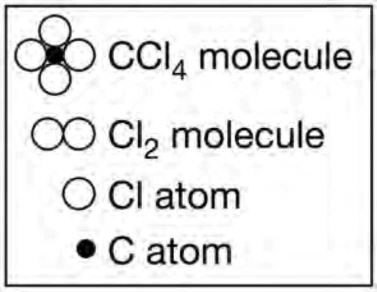
Liquid



Gas

151. CCl_4(*l*) is placed in a previously evacuated container at 30 °C, and some of the CCl_4(l) evaporates. In the box below, draw a particulate diagram to show the species in the container after some of the CCl_4(l) has evaporated.





Section 3.02

Topic 3.2 First Year Review

152. Which is expected to have the greater dipole moment? In each case justify your answer.
    a. C bonded to N or C bonded to F

C bonded to F. C and N are closer to each other on the periodic table and therefore have more similar electronegativity values (2.55 and 3.04 respectively) while C and F are further from each other on the periodic table and therefore have a larger difference in electronegativity (2.55 and 3.98 respectively). The greater the difference in electronegativity the greater the dipole moment.

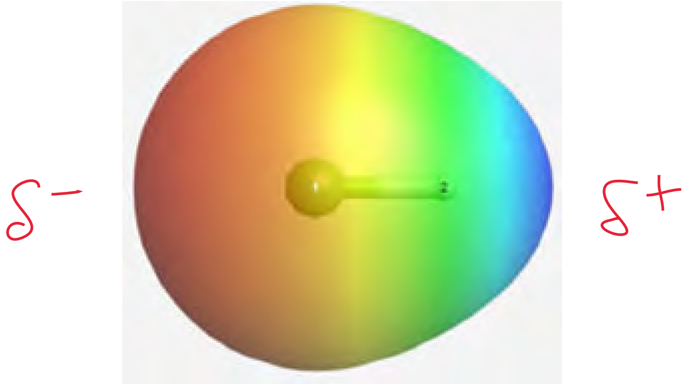
    b. N bonded to P or N bonded to Sb

N bonded to Sb. Electronegativity decreases down the group therefore Sb has a smaller electronegativity than P. The difference in electronegativity between N and Sb (3.04 and 2.05) is greater than the difference in electronegativity between N and P (3.04 and 2.19).

153. According to Coulomb's Law, does the force of attraction increase, decrease, or remain the same as the size of the charge difference increases? (Assume the distance between the charges remains the same.) Explain your answer.

The force of attraction increases as the size of the charge difference increases. According to Coulomb's Law the force of attraction is porportional to the product of the charges, as shown in the equation below with q_1 and q_2. As q_1 and q_2 get bigger the force of attraction also gets bigger if the denominator stays the same.

$$F = k\frac{q_1 q_2}{r^2}$$



154. The image above is of the electron density of HF.

a. Is the molecule polar or nonpolar? Justify your answer by referring to the electron density.

The molecule is polar. The electron density is more centered around one atom than the other. Since the electron density is not shared equally it is a polar molecule.

b. To the left and right of the molecule label one side “δ +” and the other side “δ −”. Justify your choice by referring to the electron density of the molecule.

The greater electron density will be more negatively charged than the less electron dense area. A greater electron density means that the electrons are found more often around that area of the molecule.

c. How does a neutral particle like HF gain a partial positive and partial negative charge?

An unequal sharing of electrons.

Topic 3.1 Worksheet

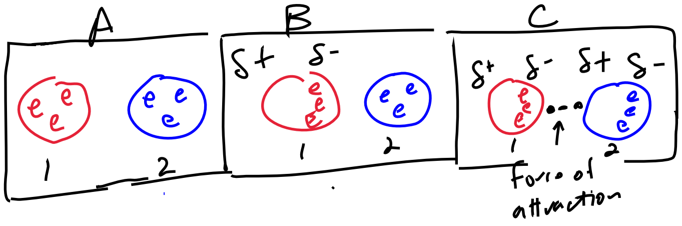
155. What's the difference between a single bond (intramolecular force) and intermolecular forces?

A single bond is within a molecule while an intermolecular force is between two different molecules.

156. Explain how …

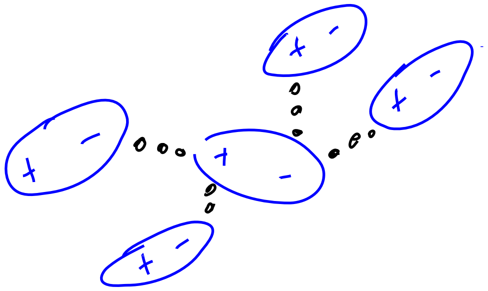
a. London dispersion forces form. Include a picture.

Random movement of electrons causes a temporary dipole. The adjacent molecule is induced to form a dipole because of the temporary dipole. The temporary dipole goes away and the induced dipole goes away as a result.



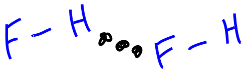
b. Dipole-dipole forces form. Include a picture.

A polar particle is attracted to another polar particle; positive end of a polar particle is attracted to the negative end of a different polar particle.



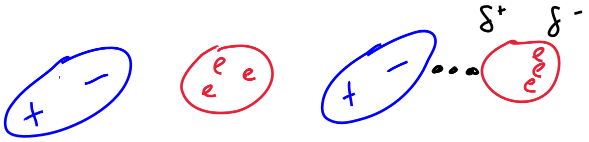
c. Hydrogen bonds form. Include a picture.

A H bound to a N, O, or F is attracted to a different particles N, O, or F. (Macromolecules can form H bonds within the same structure, e.g. proteins and DNA)



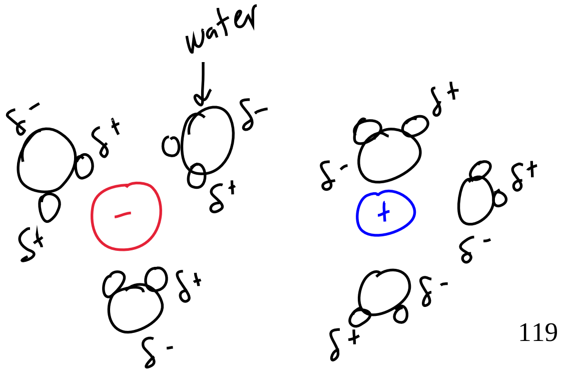
d. Dipole-induced dipole forces form. Include a picture.

A polar particle induces a nonpolar particle to become polar.



e. Ion-dipole forces form. Include a picture.

An ion (+ or -) is attracted to a polar particle.



157. Explain polarizability.

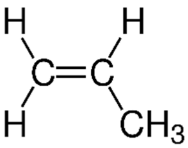
The ability of a molecule to become polar.

158. Explain how polarizability increases with an increase in the number of electrons.

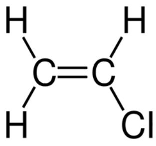
As the number of electrons increases there is a greater likelihood of being temporarily polar because there is a greater likelihood of more electrons being on one side than the other. As the chance of being polar increases the London dispersion force also increases in strength.

159. Explain how polarizability increases with an increase in surface area of a molecule.

Larger surface area tends to have more electrons. As the number of electrons increases the chance of being polar increases.



Propene



Vinyl chloride (chloroethene)

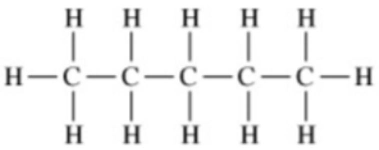
160. Answer the questions that follow about propene (CH_2CHCH_3) and vinyl chloride (CH_2CHCl), shown above.

a. Identify the intermolecular forces present in each molecule.

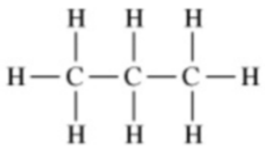
Propene has only London dispersion forces while vinyl chloride has London dispersion forces and dipole-dipole forces.

b. The boiling point of liquid propene (226 K) is lower than the boiling point of liquid vinyl chloride (260 K). Account for this difference in terms of the types and strengths of intermolecular forces present in each liquid.

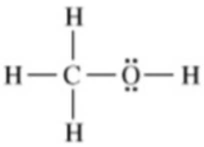
Vinyl chloride has a larger electron cloud, is more polarizable, and has a larger dipole moment. Thus the intermolecular forces in vinyl chloride are stronger than propene.



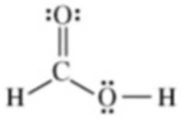
Pentane



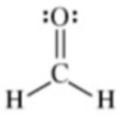
Propane



Methanol



Methanoic (formic) acid



Methanal (formaldehyde)

161. Answer the following questions in terms of principles of chemical bonding and intermolecular forces. In each explanation where a comparison is to be made, a complete answer must include a discussion of both substances. The complete Lewis electron-dot diagrams shown above may be useful in answering parts of this question.

   a. At 1 atm and 298 K, pentane is a liquid whereas propane is a gas. Explain.

Pentane and propane are both nonpolar and therefore only have London dispersion forces between their particles. Since pentane has more electrons than propane it is more polarizable and therefore has stronger London dispersion forces than propane. Stronger forces of attraction means a greater boiling point as the molecules are held together with a greater force of attraction.

   b. At 1 atm and 298 K, methanol is a liquid where as propane is a gas. Explain.

Methanol has London dispersion forces, dipole-dipole, and hydrogen bonding while propane only has London dispersion forces. Since methanol has stronger intermolecular forces it will be a liquid while propane has weak intermolecular forces and is therefore a gas.

162. Two types of intermolecular forces present in liquid H_2S are London (dispersion) forces and dipole-dipole forces.

a. Compare the strength of the London (dispersion) forces in liquid H_2S to the strength of the London (dispersion) forces in liquid H_2O. Explain.

H_2O has weaker London dispersion forces than H_2S. H_2S has more electrons and is therefore more polarizable. The more polarizable the stronger the London dispersion forces.

b. Compare the strength of the dipole-dipole forces in liquid H_2S to the strength of the dipole-dipole forces in liquid H_2O. Explain.

H_2O will have stronger dipole-dipole forces than H_2S. O is more electronegative than S and thus H-O is a more polar bond than H-S. A stronger dipole creates a stronger dipole-dipole force.

163. Consider Br_2 and Cl_2.

a. What intermolecular forces do they both exhibit? Justify your answer.

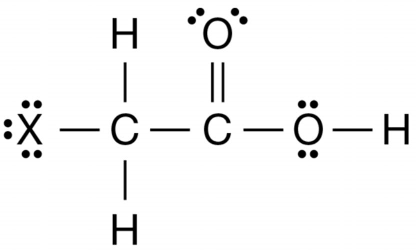
London dispersion forces only. Both are nonpolar.

b. Which has stronger intermolecular forces? Justify your answer.

Br_2 has stronger IMF because it has more electrons than Cl_2. More electrons means more polarizable.

c. Which do you expect to be a liquid at room temperature? Explain why.

Br_2 would be more likely to be a liquid at room temperature because it has stronger IMF.



164. Consider the haloacetic acids, illustrated above with an “X” in place of the halogen.

a. What intermolecular forces are present in all of the haloacetic acids?

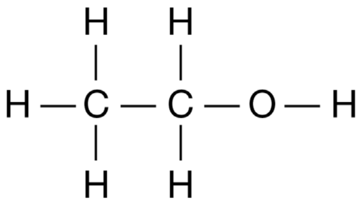
London dispersion forces, dipole-dipole forces, and hydrogen bonding.

b. Which haloacetic acid would you expect to have the highest boiling point, Cl, Br, or I?

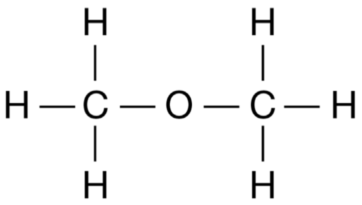
Iodine.

c. Explain your reasoning.

The only difference for all of haloacetic acids is the halogen taking the place of "X". Iodine has the most electrons of all of the halogens listed. More electrons means more polarizable which means stronger London dispersion forces than Cl or Br. Stronger IMF requires more energy to boil.



Ethanol



Dimethyl ether

165. Answer the following questions about the isomers ethanol and dimethyl ether.

a. Explain why methanol and dimethyl ether are isomers of each other.

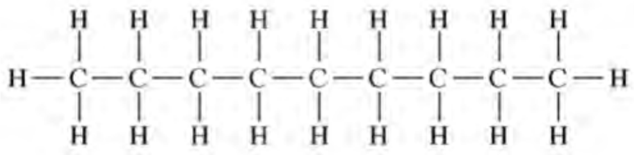
Same number and types of atoms but a different arrangement.

b. Identify all intermolecular forces in both molecules.

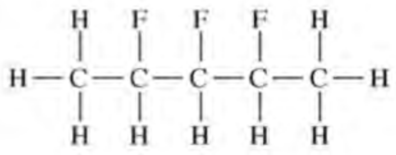
Ethanol has London dispersion forces, dipole-dipole forces, and hydrogen bonding.
Dimethyl ether has London dispersion forces and dipole-dipole forces.

c. Ethanol has a boiling point of 78 ℃ while dimethyl ether has a boiling point of −24 ℃. Identify the intermolecular force that is most responsible for the difference.

The hydrogen bonding in ethanol is responsible for the difference in boiling point. Dimethyl ether does not exhibit hydrogen bonding and therefore has weaker forces of attraction.



Nonane



2,3,4-trifluoropentane

166. Shown above are the Lewis structures for nonane and 2,3,4-trifluoropentane.

a. Identify the intermolecular forces present in both molecules.

Nonane has London dispersion forces.
2,3,4-trifluoropentane has London dispersion forces and dipole-dipole forces.

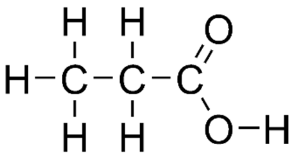
b. Nonane has a higher boiling point than 2,3,4-trifluoropentane even though they have nearly identical molar masses.

i. Which intermolecular force is most responsible for this difference?

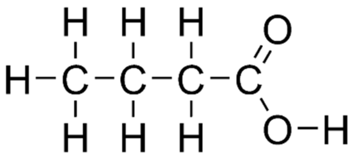
The London dispersion force is more responsible for the difference.

ii. Explain how the intermolecular force you identified in 166bi. causes the difference in boiling point.

Nonane is more polarizable because of greater surface area compared to 2,3,4-trifluoropentane. A greater polarizability increases the strength of the London dispersion forces present.



Propanoic acid



Butanoic acid

167. Shown above are the Lewis structures for propanoic acid and butanoic acid. Propanoic acid has a lower boiling point than butanoic acid.
    a. Identify all intermolecular forces present in each molecule.

Both molecules have London dispersion forces, dipole-dipole forces, and hydrogen bonding.

    b. Which intermolecular force is most responsible for the difference in boiling point?

London dispersion forces. Butanoic acid is more polarizable because it has more electrons than propanoic acid. More polarizable means stronger London dispersion forces.

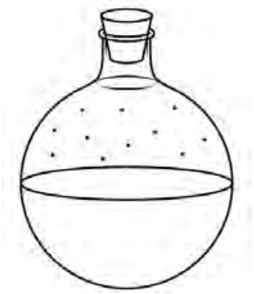
168. Br_2 has a higher boiling point than BrCl. Explain why in terms of the intermolecular forces present.

Br_2 has London dispersion forces while BrCl has London dispersion forces and dipole-dipole forces. However, the London dispersion forces in Br_2 are stronger than the forces in BrCl. This is because Br_2 is more polarizable than BrCl because Br_2 has more electrons.

NOTE: The properties of the molecules dictate the strongest force present. Since Br_2 only has LDF that means the LDF of Br_2 HAVE TO BE STRONGER than the forces present in BrCl because of its higher boiling point. Higher b.p. means stronger forces of attraction.

169. Energy is required to boil ethanol. Consider the statement "As ethanol boils, energy goes into breaking C – C bonds, C – H bonds, C – O bonds, and O – H bonds." Is the statement true or false? Justify your answer.

The statement is false. When a substance boils the intermolecular forces break and NOT the chemical bonds.



Flask A

Flask B

Flask C

Flask D

170. Shown above are four identical flasks that contain four different liquids all at the same temperature. The particles above each liquid are in the vapor phase.

a. Which flask has the weakest forces of attraction between the particles? Explain your reasoning.

Flask B. Flask B has the most number of particles in the vapor phase. The particles have to overcome the IMF in order to enter the vapor phase.

b. Which flask has the strongest forces of attraction between the particles? Explain your reasoning.

Flask C. Fewest number of particles in the vapor phase. This means it has the strongest forces of attraction and keeps the particles from entering the vapor phase.

c. If all of the particles were nonpolar, what could be a possible reason for the difference in the strength of the intermolecular forces?

Either the number of electrons making it more polarizable or the greater surface area of the particle making it more polarizable.

d. If all of the particles were polar but did not contain hydrogen bonding, …
    i. what intermolecular force of attraction would most contribute to the differences in strength?

London dispersion forces. If all particles were polar but did not have hydrogen bonding then the difference in strength of the London dispersion forces would be the most likely contributor without more information.

NOTE: A difference in polarity can also contribute to a difference in strength. A more polar particle will have stronger dipole-dipole forces compared to a less polar particle. However, that information is not provided.

    ii. would the particle in flask A have more, fewer, or the same number of electrons as the particle in flask D?

Assuming London dispersion forces are the reason for the difference in volatility, Flask A would have more electrons than Flask D. Fewer vapor particles in A than in D therefore A must have stronger IMF. If all other properties are equal then the number of electrons can strengthen the London dispersion forces present.

Container	A	B	C
Gas	Methane	Ethane	Butane
Formula	CH_4	C_2H_6	C_4H_{10}
Molar mass (g/mol)	16	30.	58
Temperature (°C)	27	27	27

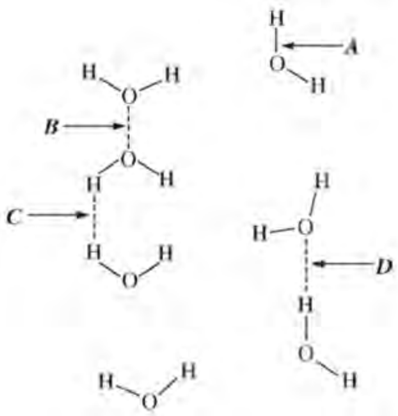
171. Consider the three gases above in three identical containers. As the pressure of the gases is increased they transition from the gas phase to the liquid phase.

a. Explain, on the molecular level, why increasing the pressure of the gas would change it from gas to liquid.

As the particles are pressed closer to each other at higher pressure the strength of their IMF would increase because of the decreased distance between particles. As the strength increases it would transition to a liquid.

b. Which gas would condense with the least amount of pressure? Explain your reasoning. You must mention all three gases in your explanation.

Butane would condense with the least amount of pressure. Butane would have stronger LDF because of more electrons and greater polarizability. Methane and ethane would both have weaker LDF because of fewer electrons and less polarizability.



172. Use the figure above to answer the questions that follow.

   a. What does arrow A indicate?

Arrow A indicates a covalent bond between H and O.

   b. Arrow D indicates a hydrogen bond. Explain why D indicates a hydrogen bond while B and C do not.

A hydrogen bond is between a H attached to a N, O, or F attracted to a different particles N, O, or F. Arrow B is between two oxygen atoms while arrow C is between two hydrogen atoms. Only arrow D is between an oxygen and a different particles hydrogen that has a O bonded to it.

173. Use the information in the table below to answer the questions that follow about three organic compounds.

Compound Name	Compound Formula	ΔH°_{vap} (kJ mol^{-1})
Propane	$CH_3CH_2CH_3$	19.0
Propanone	CH_3COCH_3	32.0
1-propanol	$CH_3CH_2CH_2OH$	47.3

Note: Students may not know what the heat of vaporization is at this point in the curriculum.

   a. Explain why propane has a smaller heat of vaporization than propanone.

Propane has weaker IMF than propanone. Propane only has LDF while propanone has LDF and dipole-dipole forces. It takes less energy to break the IMF in propane compared to propanone.

   b. Explain why propanone has a smaller heat of vaporization than 1-propanol.

1-propanol has stronger IMF than propanone. 1-propanol has hydrogen bonding, dipole-dipole, and London dispersion forces while propanone only has dipole-dipole and London dispersion forces. Weaker forces of attractoin requires less energy to go from liquid to gas.

174. Which of the figures below correctly shows a hydrogen bond? Explain your reasoning.

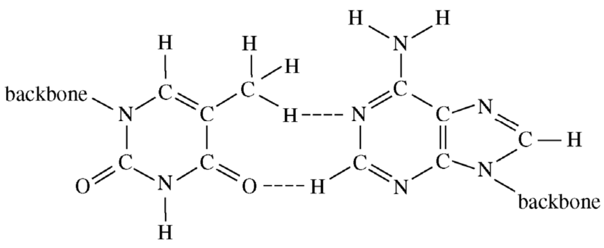


Figure A

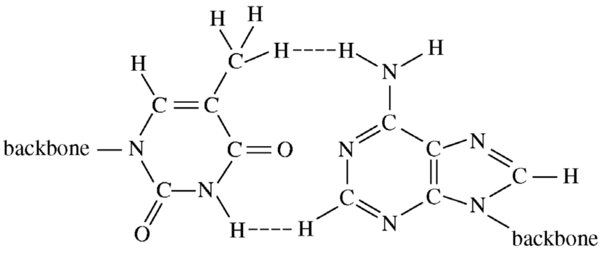


Figure B

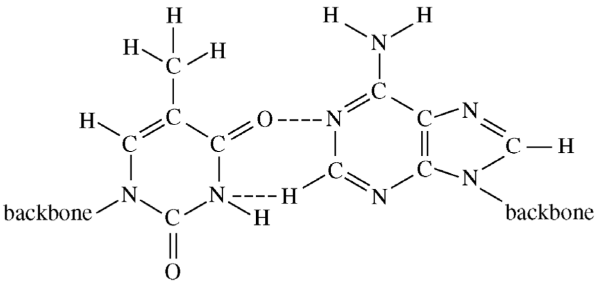


Figure C

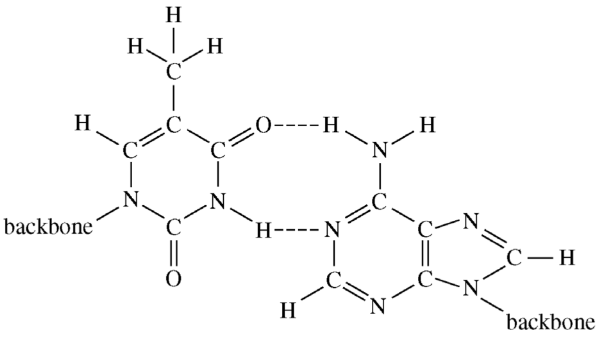


Figure D

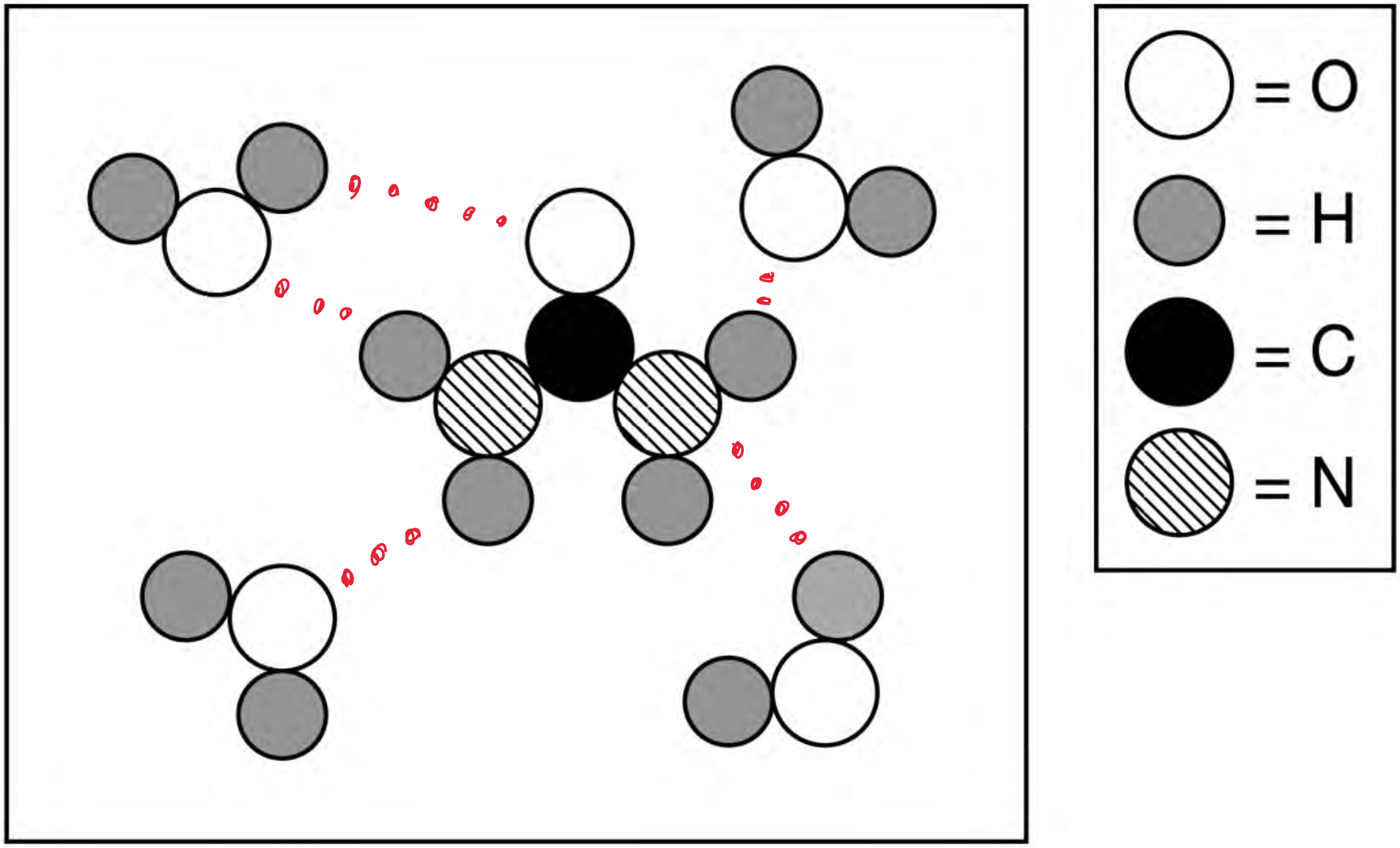
Figure A: H not bound to N, O, or F

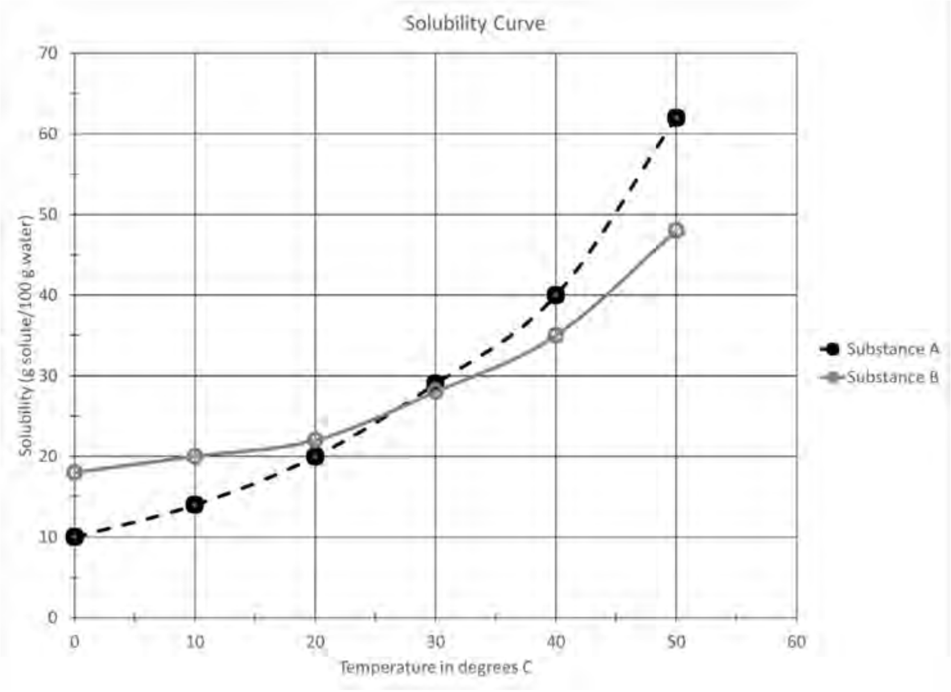
Figure B: H bound to N BUT not attracted to a different N, O, or F

Figure C: H not bound to N, O, or F

Figure D: H bound to N, O, or F and that H attracted to a different N, O, or F

175. Shown below is a molecule of urea (H_2NCONH_2) which is highly soluble in water. Draw and label the hydrogen bonds between urea and water.





176. Use the graph above to answer the questions that follow.

a. Which solid is more soluble at 45 ℃?

Substance A. The solubility of A at 45°C is about 48 g/100 g H_2O while the solubility of B at 45°C is about 40 g/100 g H_2O

b. What mass of Substance A will dissolve in 100 g of water at 10 ℃?

About 14 grams.

c. 50 grams of Substance B is placed into 100 grams of water at 40 ℃.

i. Would the solution be saturated or unsaturated?

The solution would be saturated.
Note: I don't teach supersaturated using a solubility curve. Supersaturation only occurs under special circumstances and I've found that students think it happens in every instance above the solubility curve.

ii. How many grams of Substance B would be dissolved?

40 grams

iii. How many grams of Substance B, if any, would be set on the bottom of the beaker?

10 grams

Topic 3.10 Worksheet

177. Complete the table below about the general solubility of different compounds in water (H_2O) and hexane (C_6H_{14}).

Strongest Type of IMF in the Particle	Soluble in Water (Yes or No)	Soluble in Hexane (Yes or No)
London dispersion forces	No	Yes
Dipole-dipole force	Yes	No
Hydrogen bonding	Yes	No
Ionic bonding	Yes	No

178. Define the following terms:

a. Miscible

When a liquid is soluble in another liquid

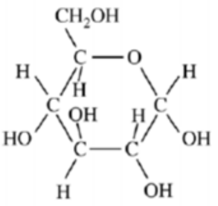
b. Soluble

Able to dissolve

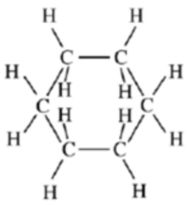
c. Solution

A homogeneous mixture of solute and solvent

179. Shown below are the structures of glucose, $C_6H_{12}O_6$, and cyclohexane, C_6H_{12}.



Glucose



Cyclohexane

a. Identify all intermolecular forces present in each compound.

Glucose has London dispersion forces, dipole-dipole forces and hydrogen bonding.
Cyclohexane has London dispersion forces only.

b. One of the molecules is soluble in water while the other is not. Which of the two would be soluble in water? Explain your reasoning.

Glucose is soluble in water. Glucose has similar IMF to water while cyclohexane does not. Similar IMF are soluble in each other.

$CH_3\text{-}CH_2\text{-}CH_2\text{-}CH_2\text{-}OH$
Compound A

$HO\text{-}CH_2\text{-}CH_2\text{-}CH_2\text{-}OH$
Compound B

$CH_3\text{-}CH_2\text{-}CH_2\text{-}CH_2\text{-}CH_3$
Compound C

180. Which of the compounds above is …

a. the most soluble in water? Explain your reasoning.

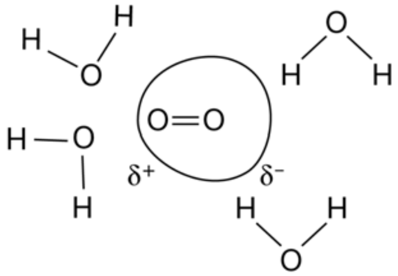
Compound B. Most similar IMF to water. Compound B has more hydrogen bonding than Compound A which makes Compound B more soluble. Compound C would be insoluble because of dissimilar IMF.

b. the least soluble in water? Explain your reasoning.

Compound C. Compound C only has LDF which is the least similar to water.

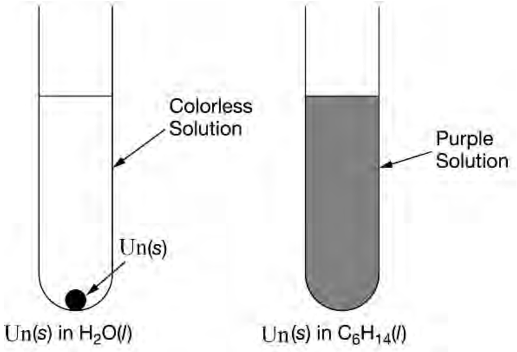
c. the most soluble in hexane, C_6H_6? Explain your reasoning.

Compound C. Compound C has the most similar IMF to hexane which is also nonpolar and has only LDF.



181. Shown above is a model that explains how oxygen is dissolved in water. Explain how water creates a temporary dipole in oxygen in order to create a dipole-induced dipole force of attraction.

Water is always polar. This dipole forces oxygen to have a temporary dipole by moving the electrons. The temporary dipole makes a force of attraction between water and oxygen.



182. An unknown solid, Un(s), was placed into two separate test tubes; one containing water and one containing hexane. What do the results indicate about the intermolecular forces in the solid?

The IMF in Un(s) are more similar to hexane than water because it dissolved in the hexane. Un(s) is more than likely nonpolar with LDF as the only forces of attraction.

183. Place an “X” in the box for the force of attraction present in each solid:

	Solid	Covalent Bonding	Intermolecular Forces	Ionic Bonding	Metallic Bonding
a.	A crystal of table salt, NaCl			X	
b.	A strip of magnesium ribbon, Mg				X
c.	A lump of sucrose, $C_{12}H_{22}O_{11}$	X	X		
d.	A bowl made of quartz, SiO_2	X			

184. Give the properties of the following types of solids. Include example molecules of the solid.
   a. Ionic solid

Made of ions. Not malleable. Not ductile. Brittle. Solid does not conduct electricity. Solution and molten conducts electricity. Low vapor pressure. Strong interaction of ions.

   b. Covalent network solid

Elements of group 14. Made of nonmetal atoms or binary nonmetal compounds. High melting points. Rigid and hard. Held together by strong covalent bonds.

   c. Molecular solid

Nonmetals. Molecules at the lattice points therefore IMF hold the solid together but those IMF are weak compared to covalent/ionic bonding. Low melting point and does not conduct electricity. Very large molecules or polymers.

   d. Metallic solid

Made of metal atoms. Conducts electricity and heat as a solid. "Sea of electrons" which means delocalized electrons. Malleable and ductile.

185. Explain why a solid ionic compound will not conduct electricity but an aqueous solution will.

As a solid the ions are unable to move because they are held rigid due to the strong electrostatic attraction between cations and anions. When aqueous the ions are able to move and thus conduct electricity.

186. Answer the following questions about the solids SO_2 and SiO_2.
   a. What type of solid is formed by SO_2?

SO_2 is a molecular solid. The solid is held together by IMF.

   b. What type of solid is formed by SiO_2?

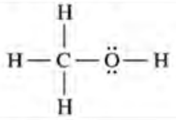
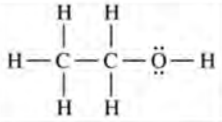
SiO_2 is a covalent network solid. The solid is held together by covalent bonds.

   c. Which solid has stronger bonding?

SiO_2 has stronger bonding.

   d. Which solid should have a lower melting point?

SO_2 would have a lower melting point because of the weak IMF holding the solid together. It will take less energy to separate the SO_2 molecules from each other.

Substance	Lewis Diagram	Boiling Point
CH_3OH		338 K
C_2H_5OH		351 K

187. Consider the information in the table above.
    a. Identify the intermolecular forces present in both substances.

CH_3OH has LDF, dipole-dipole, and hydrogen bonding.
C_2H_5OH has LDF, dipole-dipole, and hydrogen bonding.

    b. Which substance would have the lower equilibrium vapor pressure?

C_2H_5OH would have a lower vapor pressure. The stronger forces of attraction means that fewer molecules will be able to escape the surface of the liquid.

    c. Identify the intermolecular force most responsible for the lower vapor pressure.

LDF would be the IMF most responsible for the lower vapor pressure

188. The electrical conductivity of an ionic compound increases with the number of ions in the formula unit. Explain why a 1.0 M solution of NaCl would have less electrical conductivity than a 1.0 M solution of Na_3PO_4.

A 1.0 molar solution of NaCl would only have two moles of ions per liter of solution; 1 mole of Na^+ and 1 mole of Cl^- per liter of solution. A 1.0 molar solution of Na_3PO_4 would have four moles of ions per liter of solution; 3 moles of Na^+ and 1 mole of PO_4^{3-} per liter of solution. The more ions present in solution the greater the electrical conductivity.

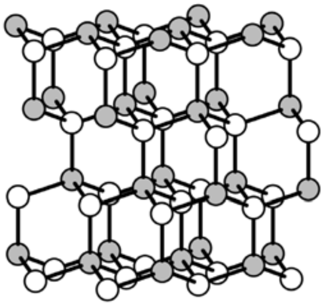
189. Consider the melting of NaCl or MgS.
    a. What types of bonds are being broken when the compounds melt? Explain your reasoning.

Ionic bonds are being broken. The compounds are made of ions.

    b. Which would you expect to have a higher melting point? Use Coulomb's Law to explain your reasoning.

MgS would have a higher melting point. Mg^{2+} and S^{2-} have higher charges than Na^+ and Cl^-. The higher charges means a greater force of attraction. A greater force of attraction requires more energy to break apart. The difference in size is relatively nonexistent as Mg^{2+} is smaller than Na^+ and Cl^- is smaller than S^{2-}.

190. Shown below is the solid boron nitride.



Cubic Form of Boron Nitride

○ Boron atom
● Nitrogen atom

a. What type of solid is boron nitride?

Covalent network solid

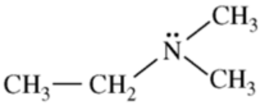
b. What evidence in the picture supports your answer?

Atoms are at the lattice point and they are covalently bonded to one another.

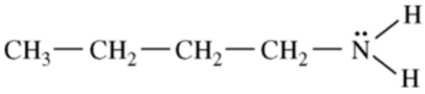
c. Boron nitride is a very hard solid. What about its structure gives it this property?

The covalent bonds make it a very strong substance.

191. Shown below are two isomers. Which would have the higher boiling point and why?



Compound 1



$CH_3-CH_2-CH_2-CH_2-\ddot{N}H_2$

Compound 2

Compound 2 would have the higher boiling point. Compound 2 has stronger IMF. Compound 2 has LDF, dipole-dipole, and hydrogen bonding while Compound 1 only has LDF and dipole-dipole forces.

Section 3.05

Topic 3.5 First Year Review

192. What is kinetic energy?

Energy in motion. In Chemistry, KE is generally tied to the motion of the particles and is proportional to the temperature in Kelvin.

193. What is potential energy?

Stored energy. In Chemistry, PE is generally considered to be in forces of attraction that include IMF and covalent/ionic bonding.

194. A 5.00 L container of methane gas is full. All of the methane gas is moved to a 20.00 L container.
   a. Is the container 100% full or only 25% full of methane? Explain your reasoning.

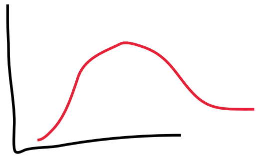
The 20.00 L container is 100% full. A gas will occupy the volume of the container.

   b. Did the mass of methane gas in the 20.00 L container increase, decrease, or remain the same as the mass in the 5.00 L container?

The mass remained the same. The volume does not effect the mass of the gas.

195. What does a Maxwell-Boltzmann distribution illustrate? Draw an example of one.

A count of particles vs. a measurement of energy/velocity.



196. Does the velocity of a gas increase, decrease, or remain the same as …

a. Temperature is increased.

As temperature increases velocity increases.

b. molar mass is increased.

As molar mass increases velocity decreases.

197. Gas A (75 °C, molar mass = 50 g/mol) and gas B (15 °C, molar mass = 25 g/mol) are placed in the same container and are allowed to come to thermal equilibrium.

a. What happens to the average kinetic energy of gas A as it approaches thermal equilibrium?

The average KE_A decreases as the sample cools. Lower temperature = lower average KE.

b. What happens to the average kinetic energy of gas B as it approaches thermal equilibrium?

The average KE_B increases as the sample heats up. Higher temperature = higher average KE

c. Compare the temperature of gas A to the temperature of gas B at thermal equilibrium. Explain your reasoning.

At thermal equilibrium the temperature of gas A equals the temperature of gas B. Thermal equilibrium occurs when the gases are at the same temperature.

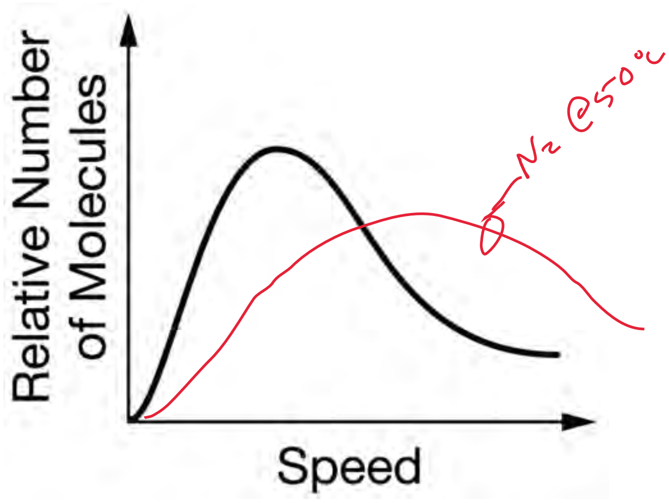
d. Compare the average kinetic energy of gas A to the average kinetic energy of gas B at thermal equilibrium. Explain your reasoning.

The average KE_A = the average KE_B. If the temperature is the same the average KE is the same. Temperature is a measure of KE.

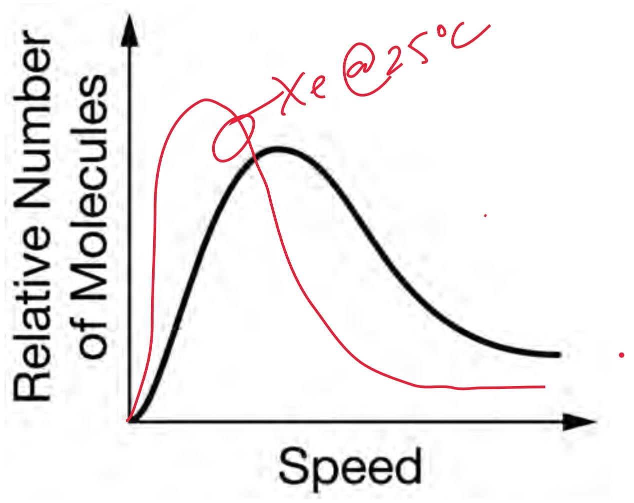
e. Compare the average speed of gas A to the average speed of gas B at thermal equilibrium. Explain your reasoning.

Gas A has a lower speed than Gas B. They have the same average KE however $KE=1/2mv^2$. Although they have the same KE their velocities (i.e. speed) would be different because their molar masses are different. As the molar mass increases the velocity decreases.

198. Use the Maxwell-Boltzmann distributions shown below to answer the questions that follow.



Shown above is the distribution for N_2 at 25 °C.
Draw and label the distribution for N_2 at 50 °C.



Shown above is the distribution for N_2 at 25 °C.
Draw and label the distribution for Xe at 25 °C

199. Describe the effect …

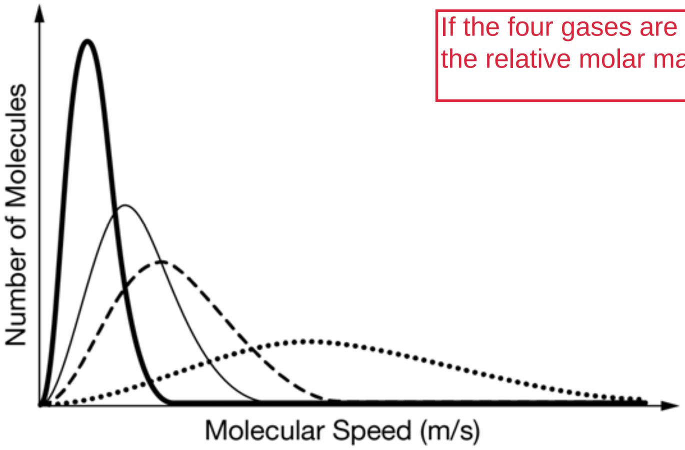
a. of raising the temperature on the motion of gas particles. Explain your reasoning.

As temperature increases the motion of the gas particles increases. According to the KMT temperature is proportional to the average KE.

b. on the pressure of a gas when the temperature is increased. Explain your reasoning.

As temperature increases the pressure of a gas increases. Since the motion of the particle increases the number of collisions with the container increases. Pressure is a measure of the number of collisions with the container.

200. The Maxwell-Boltzmann distrubition below was created using four different gases. What specific property of the gas can you determine from the graph alone?



If the four gases are at the same temperature then the relative molar mass can be determined.

(A)

T = 30°C
P = 1 atm

(B)

T = 50°C
P = 0.5 atm

(C)

T = 40°C
P = 2 atm

(D)

T = 50°C
P = 1 atm

(E)

T = 40°C
P = 0.5 atm

201. Use the five flasks above to answer the questions that follow.

a. Which flask do the molecules have the greatest average speed of the particles? Explain your answer.

Flask B. N_2 has the lower molar mass. Flask B and D have the highest temperature and are therefore have the greatest KE. N_2 has a lower molar mass than O_2 and is therefore moving faster.

b. Which nitrogen flask has the most number of particles? Explain your reasoning.

Flask A. Flask A has a higher pressure at a lower temperature therefore it has more particles than Flask B.

c. Which oxygen flask has the most number of particles? Explain your reasoning.

Flask C. Flask C has the highest pressure at the lowest temperature therefore it must have more particles.

d. If flask C were placed into flask E, and there was no change in temperature, what would be the total pressure? Explain your reasoning.

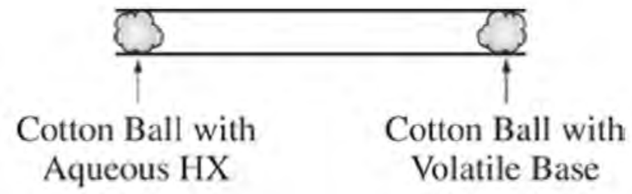
1.5 atm. The total pressure is the sum of the partial pressures since the flasks are all the same size.

e. Which flasks have the same average kinetic energy? Explain your reasoning.

Flasks B and D have the same average KE. Flask C and E have the same average KE.

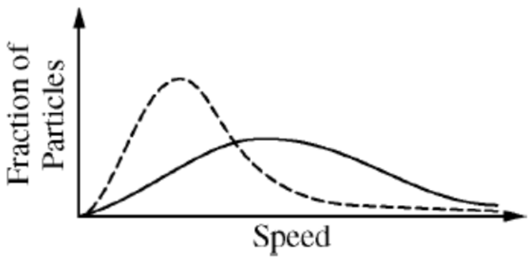
Temperature is proportional to the average KE.

202. The apparatus below is a glass tube with two cotton balls stuck in the ends. One cotton ball is soaked with an aqueous solution of HX while the other cotton ball is soaked with a volatile base. When HX and the base meet a white solid is produced and creates a vertical ring in the tube. In the following situations, determine if the white ring will form closer to HX, closer to the base, or near the middle. In each case, explain your reasoning.

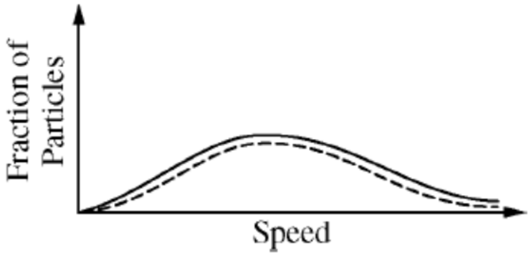


HX	Base	Location of white ring	Reasoning
HCl	CH_3NH_2	Closer to HX	CH_3NH_2 (MM = 31 g/mol) travels faster than HCl (MM = 36.5 g/mol). Therefore the gases will meet closer to HCl
HBr	CH_3NH_2	Closer to HX	Same reasoning as above.
HCl	$CH_3(CH_2)_6NH_2$	Closer to the base	HCl (MM = 36.5 g/mol) is much faster than the base (MM = 115 g/mol) because HCl's mass is much less.

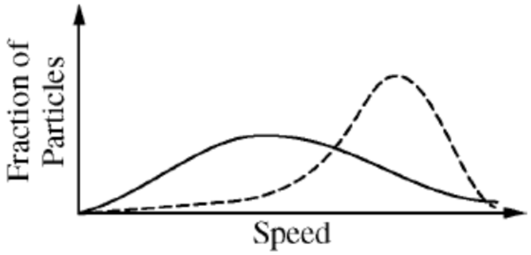
203. The Maxwell-Boltzmann distributions below show two different gases at the same temperature. Compare the molar mass of the gas depicted by the solid line (Gas X) to the molar mass of the gas depicted by the dashed line (Gas Y) by stating whether the molar mass of the Gas Y is greater than, less than, or equal to the molar mass of Gas X.



Gas Y has a greater molar mass than Gas X.



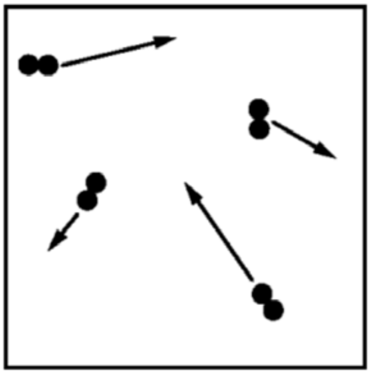
Gas Y has a molar mass equal to Gas X.



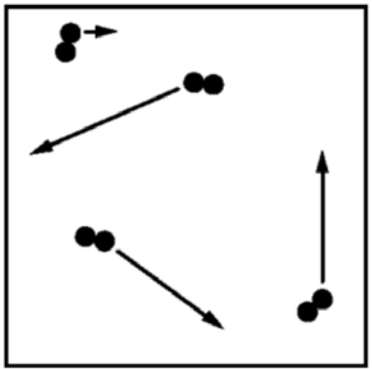
Gas Y has a molar mass lighter than Gas X.

Note: This is a weird looking Boltzman diagram but it is being used to stress the difference in speed and hence MM.

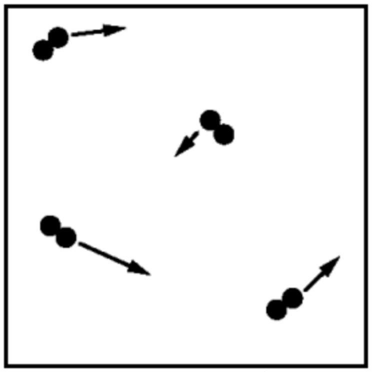
204. The particle picture on the left is a gas at a given temperature. Determine if the particle pictures are at a higher temperature or a lower temperature. Explain your reasoning.



Original temperature



[✓ Higher Temp] or Lower Temp



Higher Temp or [✓ Lower Temp]

Reasoning	The arrows are longer which means they are moving faster. The "arrows" are vectors.	The arrows are smaller which means they are moving slower.

Section 3.06

Topic 3.4 First Year Review

205. Perform the following conversions.

a. A rigid container is 450 mL. Convert the volume to liters.

$$\left.450\text{ mL}\,\right|\,\frac{1\text{ L}}{1000\text{ mL}} = 0.450\text{ L}$$

b. How many mL is 2.00 L?

$$\left.2.00\text{ L}\,\right|\,\frac{1000\text{ mL}}{1\text{ L}} = 2000\text{ mL}$$

c. Convert 125 mm Hg to atm.

$$\left.125\text{ mm Hg}\,\right|\,\frac{1\text{ atm}}{760\text{ mm Hg}} = 0.164\text{ atm}$$

d. Determine the pressure in kPa if it is measured at 1.72 atm.

$$\left.1.72\text{ atm}\,\right|\,\frac{101.3\text{ kPa}}{1\text{ atm}} = 174\text{ kPa}$$

e. A canister is measured to be at 0.892 kPa.

i. Determine the pressure in mm Hg.

$$\left.0.892\text{ kPa}\,\right|\,\frac{760\text{ mm Hg}}{101.3\text{ kPa}} = 6.69\text{ mm Hg}$$

ii. Determine the pressure in atm.

$$\left.6.69\text{ mm Hg}\,\right|\,\frac{1\text{ atm}}{760\text{ mm Hg}} = 8.81\times10^{-3}\text{ atm}$$

f. An oven is ran at 200. °C. Determine the temperature in Kelvin.

$T_K = T_C + 273$

$T_K = 200 + 273 = 473\text{ K}$

g. The flashpoint of carboard is 700. K. Determine the flashpoint in Celsius. (The flashpoint is the temperature at which the substance will auto-ignite.)

$T_K = T_C + 273$

$700 = T_C + 273$

$T_C = 427\,°C$

206. Give the following equations or definitions. Identify any variables in an equation.

a. Ideal gas law

$$PV = nRT$$

pressure (atm)

volume (L)

moles

temperature (K)

Ideal gas constant $\left(0.08206 \frac{L \cdot atm}{mol\ K}\right)$

b. Ideal gas law (solved for molar mass)

$$MM = \frac{gRT}{PV}$$

molar mass

grams

c. Ideal gas law solved for density

$$\frac{g}{V} = \frac{PMM}{RT} \quad \text{or} \quad d = \frac{PMM}{RT}$$

d. Mole fraction

$$\chi_A = \frac{n_A}{n_{total}}$$

mole fraction

e. Partial pressure

$$P_A = \chi_A P_{tot}$$

partial pressure of A

f. STP

Standard temperature and pressure

0°C or 273 K

1 atm

g. Molar volume of a gas

@ STP 22.4 L/mol

207. Determine the mass, in grams, of each gas under the conditions given.

a. 22.4 L of $CH_4(g)$ at STP

1 mole @

16.042 g

OR

$$\frac{22.4\text{ L } CH_4 \mid 1\text{ mol}}{\mid 22.4\text{ L}} = \frac{1\text{ mol} \mid 16.042\text{ g}}{\mid 1\text{ mol}} = 16.042\text{ g}$$

b. 2.2 L of $CO_2(g)$ at $\boxed{1 \text{ atm and } 0\ °C}$ STP

.1 mole

molar mass = 44.01 g/mol

4.4 g

OR

$$\frac{2.2\text{ L } CO_2 \mid 1\text{ mol}}{\mid 22.4\text{ L}} = \frac{.0982\text{ mols} \mid 44.01\text{ g}}{\mid 1\text{ mol}} = 4.32\text{ g}$$

c. .2 L of $SO_2(g)$ at STP

.01 mol @

$MM_{SO_2} = 64.06$

.641 g

.641 g

d. 4 L of $N_2O(g)$ at 273 K and 1 atm

$MM = 44.02\ \frac{g}{mol}$

.2 mol

8.8 g

208. For the following changes to a gas, determine which will have a greater effect for the property specified without using a calculator.

a. Would the pressure increase, decrease, or remain the same if the Celsius temperature is doubled from 10 ℃ to 20 ℃ while the volume the gas occupies is doubled from 2 L to 4 L? Explain your reasoning.

The pressure would decrease. The change in temperature in Kelvin is only 10 K therefore doubling of the volume will have a greater effect on pressure.

b. Would the volume of a gas increase, decrease, or remain the same if the Kelvin temperature were doubled while the pressure is decreased by a factor of four? Explain your reasoning.

The volume would increase. Both changes will increase the volume.

c. Would the temperature of a gas have to increase, decrease, or remain the same if the volume of the container is reduced by half while the pressure was doubled? Explain your reasoning.

The temperature would stay the same.

209. Perform the following calculations.

a. Determine the partial pressure of each gas if the canister contains 0.5 moles of O_2, 1.0 moles of N_2, and 0.5 moles of Ar at a total pressure of 600 mm Hg.

$X_{O_2} = \frac{.5}{2} = 0.25$

$X_{N_2} = \frac{1}{2} = 0.5$

$X_{Ar} = \frac{.5}{2} = 0.25$

$P_{O_2} = X_{O_2} P_{tot}$

$P_{O_2} = (.25)(600)$

$P_{O_2} = 150$ mm Hg

$P_{N_2} = X_{N_2} P_{tot}$

$P_{N_2} = (.5)(600)$

$P_{N_2} = 300$ mm Hg

$P_{Ar} = X_{Ar} P_{tot}$

$P_{Ar} = (.25)(600)$

$P_{Ar} = 150$ mm Hg

b. Determine the partial pressure of N_2 if the canister contains 0.5 moles of O_2, 0.5 moles of N_2, and 0.5 moles of Ar at a total pressure of 900 mm Hg.

X of each gas $= .33$

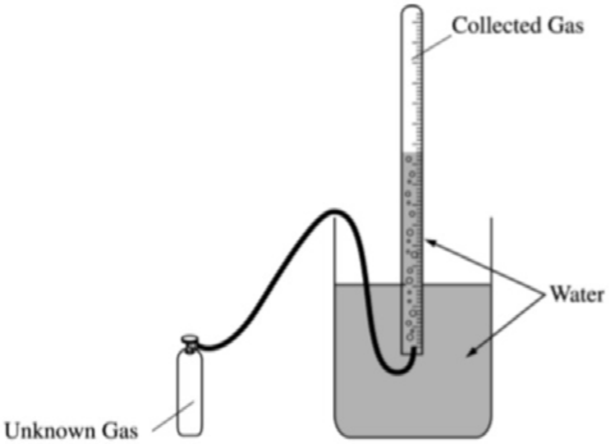
$P_A = X_A P_{tot}$

$P_A = (.33)(900)$

$P_A = 300$ mm Hg

Each gas exerts 300 mm Hg of pressure

210. A student collects a gas over water using the apparatus shown below. Answer the questions that follow about the lab experiment.



a. Explain why the water level on the inside of the collection tube and the outside of the collection tube needs to be at the same level.

If the water levels are the same then the pressure of the gas inside the tube is the same as the pressure of the room. If they are at the same level we can assume that $P_{gas} = P_{room}$.

b. Explain why the pressure of the room is the same as the pressure of the gases (assuming a is true).

If the water levels are the same then the gas and room are pushing with the same amount of force. If the water level is higher on the inside of the tube then the room pressure is pushing with more force than the gas pressure. The opposite is true if the water level on the inside of the tube is lower than the water in the beaker.

c. Explain why the pressure of the room is NOT the pressure of the collected gas.

The gas is collected over water. The water also has vapor over the surface therefore there is also water vapor collected and causing pressure inside the tube.

d. How would you determine the mass of gas delivered?

Weigh the canister. Deliver the gas to the inside of the tube. Weight the canister again.

e. Use the data below to determine the molar mass of the unknown gas.

Room Pressure	750 mm Hg
P_{H2O} at 22 °C	19.8 mm Hg
Mass of Canister Before the lab	25.100 g
Mass of Canister After the lab	24.276 g
Temperature of Water (°C)	22.0 °C
Volume of Gas Collected (mL)	358 mL

$PV = \frac{gRT}{MM}$

$MM = \frac{gRT}{PV}$

$MM = \frac{(.824)(.08206)(295)}{\left(\frac{730.2}{760}\right)(.358)}$

$P_{room} = P_{gas} + P_{H_2O}$

$750\text{ mmHg} = P_{gas} + 19.8\text{ mmHg}$

$730.2\text{ mmHg} = P_{gas}$

$g = \text{Canister before} - \text{Canister after}$

$g = 25.100 - 24.276$

$g = 0.824\text{ g}$

$MM = 58.0\ \frac{g}{mol}$

211. A 5.00 milliliter vial contains 0.750 grams of CCl_4 at 0 °C. Calculate the pressure in the vial at 25 °C.

$PV = nRT$

$P(.005) = \left(\frac{.750}{153.81}\right)(.08206)(298)$

$P = 23.9\text{ atm}$

212. A gas has a vapor density of 7.50 g L^{-1} at 350 K. If a container of the gas has a pressure of 2.22 atm what would be the molar mass of the gas?

$PV = \frac{gRT}{MM}$

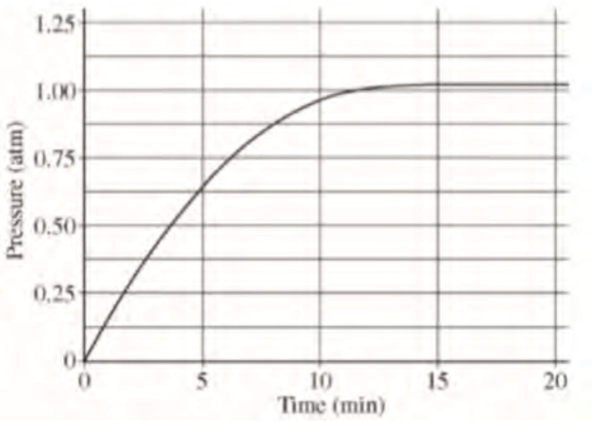
$MM = \frac{gRT}{VP}$

density

$MM = \frac{(7.50)(.08206)(350)}{2.22}$

$MM = 97.1\text{ g/mol}$

213. The pressure in a 1.00 L rigid flask at 1100 K was measured and recorded over time. Use the graph below to answer the questions that follow.



a. Determine the moles of gas present in the flask at 5 minutes.

$PV = nRT$ $(.625)(1) = n(.08206)(1100)$

$n = 0.00692$ mols

b. Determine the moles of gas present in the flask at 20 minutes.

$PV = nRT$ $(1.01)(1) = n(.08206)(1100)$

$n = 0.0112$ moles

c. Which reaction below could be occurring in the flask? Explain your reasoning. In your explanation you must discuss all three reactions.

$CaCO_3(s) \rightleftharpoons CaO(s) + CO_2(g)$
Reaction A

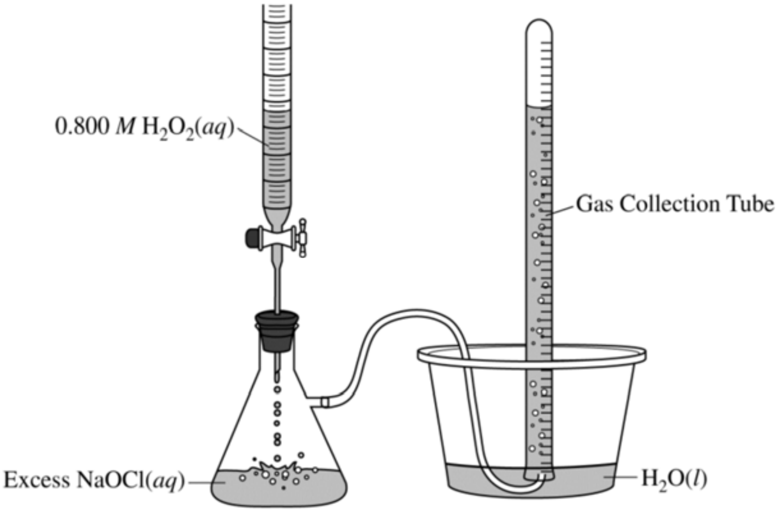
$4\,Fe(s) + 3\,O_2(g) \rightarrow 2\,Fe_2O_3(s)$
Reaction B

$CH_4(g) + 2\,O_2(g) \rightarrow CO_2(g) + 2\,H_2O(g)$
Reaction C

Reaction A is occurring in the flask. Reaction A is creating a gas therefore the change in pressure would go from "0" to a positive value. Reaction B would see a decrease in pressure as the gas is being converted to a solid. Reaction C would see no change in pressure as the moles of gas are the same on both sides of the equation.

$$H_2O_2(aq) + OCl^-(aq) \rightarrow H_2O(l) + Cl^-(aq) + O_2(g)$$

214. A student investigates the reaction between $H_2O_2(aq)$ and NaOCl(aq), which is represented by the net-ionic equation shown above. The student decides to produce 40.0 mL of $O_2(g)$ at a pressure of 0.988 atm and a temperature of 298 K using the reaction represented above. The student uses the equipment shown below. The student sets up a 250 mL Erlenmeyer flask fitted with a one-hole stopper. The flask is connected to a 50 mL gas-collection tube that initially is completely filled with water.



a. Calculate the volume of 0.800 M $H_2O_2(aq)$ that the student should add to excess NaOCl(aq) to produce 40.0 mL of $O_2(g)$ at 0.988 atm and 298 K.

$n_{O_2} = \frac{PV}{RT}$

$n_{O_2} = \frac{(.988)(.04)}{(.08206)(298)}$

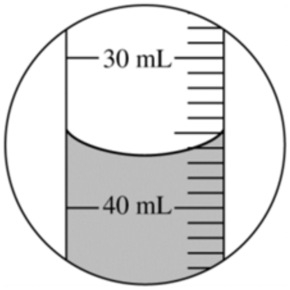
$n_{O_2} = 0.00162 \text{ mols } O_2$

$n_{O_2} = n_{H_2O_2}$

$M_{H_2O_2} = \frac{n_{H_2O_2}}{V_{H_2O_2}}$

$V = \frac{.00162}{.8} = .00203 L$

b. The student added the amount of H_2O_2(aq) calculated in part (a) to excess NaOCl(aq). However, instead of producing 40.0 mL of O_2(g), the volume indicated in the diagram below was produced.



i. Based on the diagram above, what volume of gas was produced?

36.4 mL

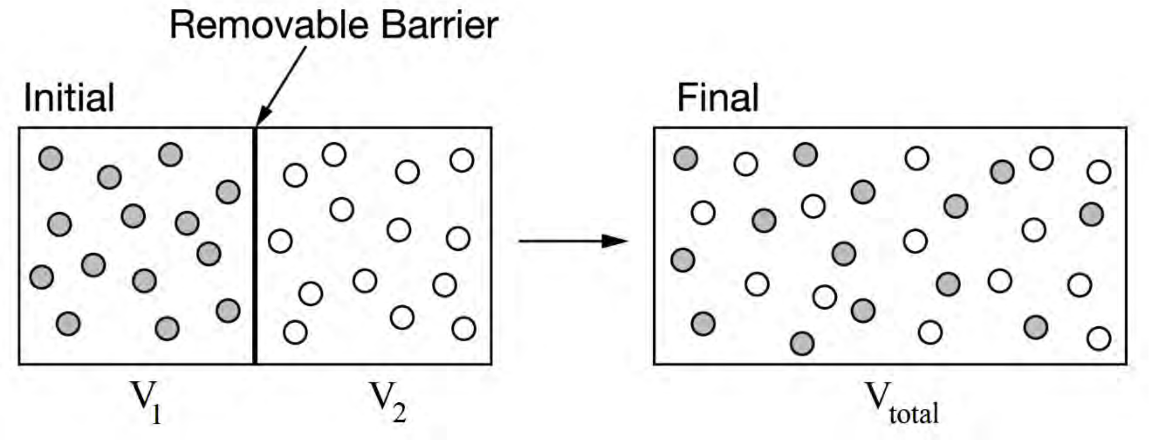
ii. Assuming that all the gas in the tube is O_2(g), calculate the percent yield of O_2(g).

$$\frac{36.4 \text{ mL}}{40.0 \text{ mL}} \times 100 = 91.0\%$$

iii. Is the assumption that all the gas in the tube is O_2(g) correct? Explain.

No. The gas is collected over water. The water vaporizes and causes pressure in the tube as well.

215. In the apparatus shown below $V_1 = V_2$. Answer the questions that follow about the apparatus assuming that there is no change in temperature throughout the experiment.



a. How can you tell that in the Initial vessel there are equal number of moles in V_1 and V_2?

There are equal number of particles in each container.

b. How would the initial pressure in V_1 compare to the initial pressure in V_2? Explain your reasoning.

The pressures would be equal. The temperature is the same in both V_1 and V_2. There are also equal numbers of particles. Therefore they would experience the same pressure.

c. How does the average KE of the particles in V_1 compare to the average KE of the particles in V_2?

They have the same average KE.

d. What information would you need to know in order to compare the speed of the particles?

The molar mass of the particles.

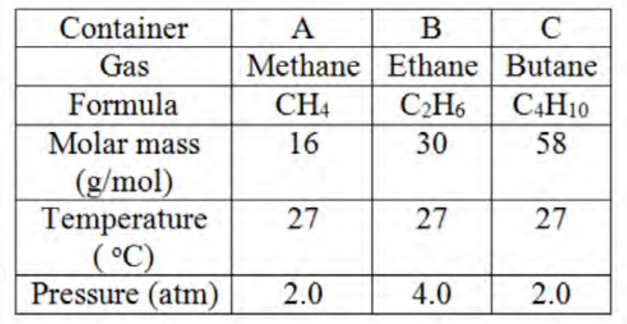
e. How does V_1 and V_2 compare to V_{total}?

$V_1 + V_2 = V_{total}$

f. What would happen to the pressure of the gas in V_1 once the barrier is removed, as shown in the Final container? Explain your reasoning.

The volume would decrease by 1/2 because the volume is doubled.

216. Use the data below for the questions that follow. All three gases are in three identical, rigid containers.



Container	A	B	C
Gas	Methane	Ethane	Butane
Formula	CH_4	C_2H_6	C_4H_{10}
Molar mass (g/mol)	16	30	58
Temperature (°C)	27	27	27
Pressure (atm)	2.0	4.0	2.0

a. Which container has the most number of particles? Explain your reasoning.

Container B. It has the greatest pressure therefore it must have more particles colliding against the container since they are all at the same temperature.

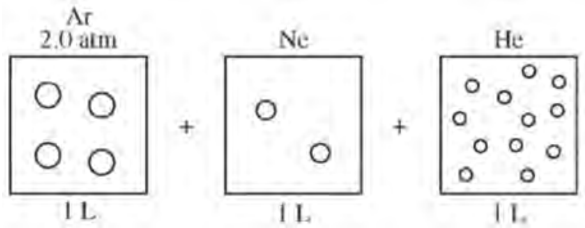
b. Which container do the particles have the lowest speed? Explain your reasoning.

Container C. Since they are all at the same temperature they all have the same average KE. Therefore the speed is inversely proportional to the molar mass. Butane has the greatest molar mass.

c. Which container has the greatest density? Explain your reasoning.

Container B. Container B has the most number of particles and therefore the greatest mass. The volume is the same for all containers so B has the greatest mass per container.

217. Consider the boxes below all at the same temperature.



a. Determine the pressure in the Ne box. Explain your reasoning.

1 atm. It has half as many particles as the Ar box which has a pressure of 2 atm.

b. Determine the pressure in the He box. Explain your reasoning.

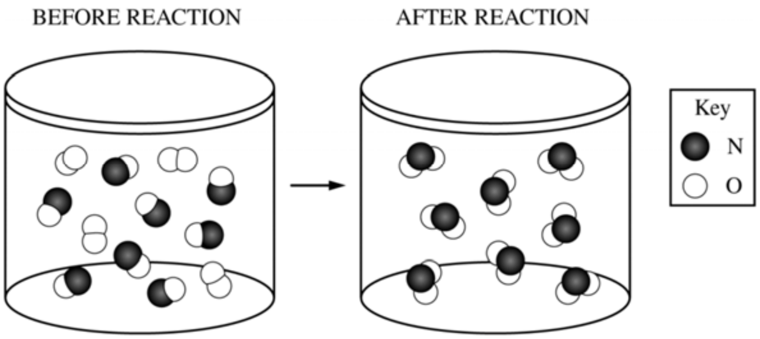
6 atm. It has 3 times the particles as Ar therefore the pressure will be 3 times the pressure of Ar.

c. What would be the total pressure if all three gases were placed in the same 1 L container.

9 atm. 2 atm from Ar, 1 atm from Ne, and 6 atm from He.

d. What would be the pressure of He if all three gases were placed into a 2 L container. Explain your reasoning.

3 atm. The pressure would decrease by 1/2 because the volume went up by 2.



218. A chemical reaction is carried out in a rigid container as illustrated above.

a. Write the balanced chemical equation.

$2NO_{(g)} + O_{2(g)} \longrightarrow 2NO_2 \ (g)$

b. How would the pressure of the container before the reaction compare to the pressure of the container after the reaction? Explain your reasoning.

The pressure of the container after the reaction would decrease. There are fewer moles of gas as products compared to the reactants. The pressure would be 2/3 of the original pressure.

219. Enough Ar(g) is pumped into a rigid container such that the pressure of Ar(g) is 0.45 atm.

a. What would be the pressure if the number of Ar(g) particles was doubled?

0.90 atm

b. The container from (a) has additional CO_2(g) pumped in so that the partial pressure of CO_2(g) is 0.30 atm.

i. What is the partial pressure of Ar(g)?

0.90 atm

ii. What is the total pressure?

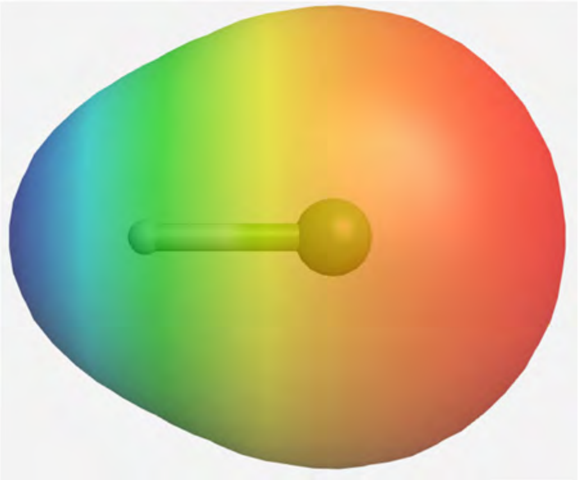
1.2 atm

Topic 3.6 First Year Review

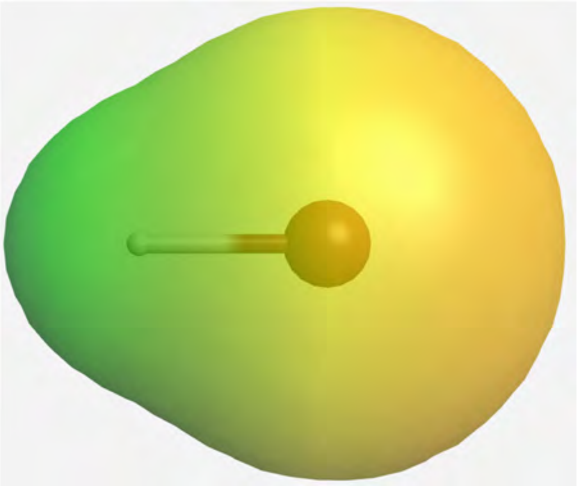
220. Does the force of attraction increase, decrease, or remain the same as the dipole moment of a particle increases?

The force of attraction increases as the dipole moment of a particle increases. The more polar the particle the stronger the force of attraction.

221. Shown below is the electrostatic potential of HF and HBr. The dipole moments are 1.972 Debyes and 1.107 Debyes, respectively.



HF. The electrostatic potential was set to a minimum of -0.1449 and a maximum of 0.3340 for comparison against HB.. The dipole moment is 1.9720 Debye.



HBr. The electrostatic potential was set to a minimum of -0.1449 and a maximum of 0.3340 for comparison against HF. The dipole moment is 1.107 Debye.

a. Which is expected to have a greater dipole-dipole force of attraction? Explain your reasoning.

HF should have a greater dipole-dipole force of attraction. HF has a greater dipole moment (1.9720 Debye) compared to HBr (1.107 Debye) which creates a larger partial charge in HF than in HBr. That larger partial charge creates a stronger dipole-dipole force of attraction according to Coulomb's Law. The electrostatic potential also shows that the electrons are more attracted to the F in HF than they are to Br in HBr.

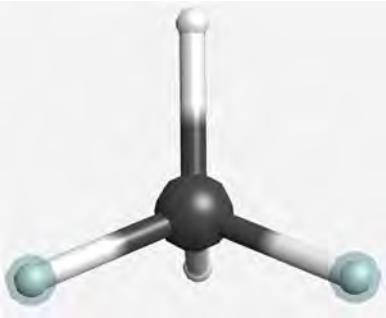
b. Based solely on the dipole dipole foces, would the boiling point of HF be greater than, less than, or equal to the boiling point of HBr? Explain your reasoning.

HF would have a boiling point that is greater than the boiling point of HBr. The dipole-dipole forces are stronger in HF than they are in HBr. The stronger dipole-dipole forces means a greater force of attraction and therefore more energy to separate the particles.

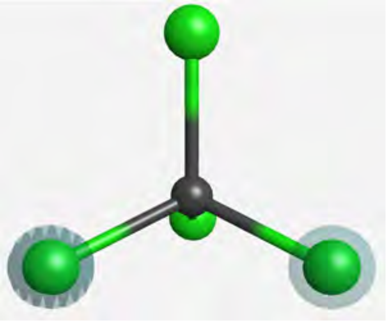
222. Does the volume a particle occupies increase, decrease, or remain the same as you move down the periodic table?

The volume a particle occupies increases as you move down the periodic table. The atomic size increases from top to bottom on the periodic table.

223. The H to H distance in CH_4 is 1.780 Å while the Cl to Cl distance in CCl_4 is 2.874 Å.



CH_4. Distance from H to H (highlighted) is 1.780 Å.



CCl_4. Distance from Cl to Cl (highlighted) is 2.874 Å.

Explain why the distance from Cl to Cl is greater than the distance from H to H.

The size of the H atom is much smaller than the size of the Cl atom. Since the Cl atom increases in size it must orient itself further from the other atoms to reduce repulsion. This results in a larger distance between atoms.

224. Shown below is the van der Waals equation for real gases. Explain each variable that factors for deviations from the ideal gas law.

$$\left(P + \underbrace{\frac{an^2}{V^2}}_{\text{Corrects for IMF}}\right)(V - \underbrace{nb}_{\text{Corrects for size of the particle}}) = nRT$$

225. Explain why a gas at low temperature deviates from the ideal gas law.

At low temperatures the IMF increase as the particles begin to slow. The strengthened IMF causes a deviation from the ideal state.

226. Explain why a gas at high pressure deviates from the ideal gas law.

At high pressure the gas takes up less space therefore the particle size is greater compared to a larger space. I.e. at 5 L the particle size is negligible but as that volume decreases because of greater pressure the particle size takes up a large percentage of the total volume.

227. Which will deviate more at high pressures, a large gas particle or a small gas particle? Explain your reasoning.

A large gas particle. At high pressure the actual size of the particle causes a deviation from the ideal gas law.

228. Which will deviate more at low temperatures, a particle with weak intermolecular forces or a particle with strong intermolecular forces? Explain your reasoning.

A particle with strong IMF. As the temperature decreases the effect of the IMF increases. Stronger IMF would cause a greater variation than particles with a weaker IMF.

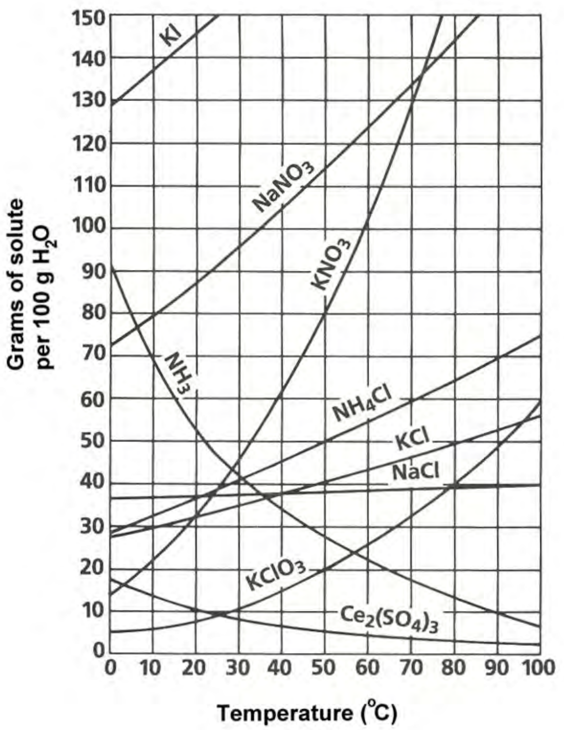
229. Would you expect the predicted pressure of a sample of CH_4 to be closer to or further from the predicted pressure of a sample of CCl_4? Explain your reasoning.

I would expect the pressure of CH_4 to be further from the pressure of CCl_4. CCl_4 is a large particle with much stronger IMF (London dispersion forces).

Section 3.08

Topic 3.7 First Year Review

Use the chart below to answer the questions that follow.



230. A saturated solution of KNO_3 is made in 100 g of water at 50 °C.

a. Determine the molarity of the $KNO_3(aq)$ at 50 °C. Assume the density of water is 1 g/mL.

80g dissolves @ 50°C

MM = $101.11 \frac{g}{mol}$

$\frac{80g\ KNO_3 \mid 1mol}{\mid 101.11g} = 0.791 mols\ KNO_3$

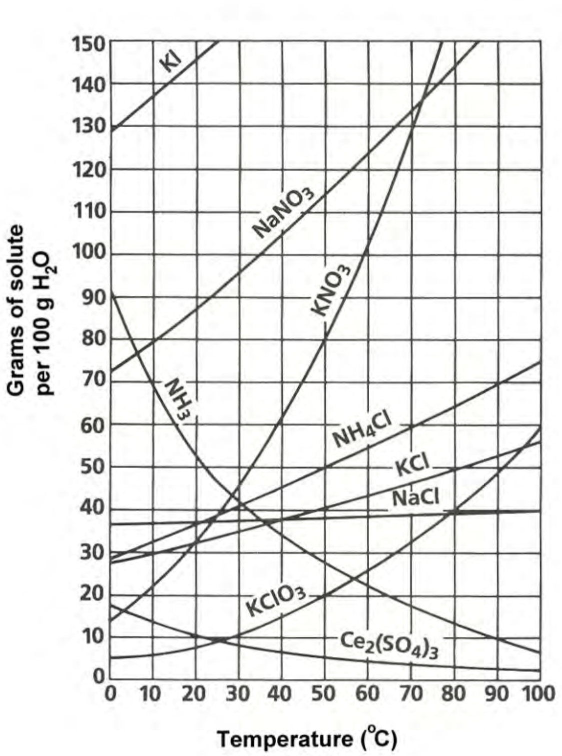
$M = \frac{n}{L} = \frac{.791\ moles}{.1\ L} = 7.91M\ KNO_3$

b. The KNO_3 solution is cooled to 25 °C. Determine the new molarity of the solution at 25 °C. Assume the density of water is 1 g/mL.

About 38g dissolved @ 25°C

$M = \frac{n}{L} = \frac{0.376}{.1\ L} = 3.76M\ KNO_3$

$\frac{38g\ KNO_3 \mid 1mol}{\mid 101.11g} = 0.376 mols\ KNO_3$



231. A saturated solution of NH_3 is made in 100 g of water at 20 ℃.

a. Determine the molarity of NH_3(aq) at 20 ℃. Assume the density of water is 1 g/mL.

About 52g NH_3 dissolves @ 20°C

$$M = \frac{n}{L} = \frac{3.05}{.1} = 30.5\,M\ NH_3$$

$$52g\ NH_3 \left|\frac{1mol}{17.034g}\right. = 3.05\ mols\ NH_3$$

b. The solution made in 231.a above is heated to 90 ℃. Determine the new molarity of NH_3(aq). Assume the density of water is 1 g/mL.

10g of NH_3 is dissolved @ 90°C

$$M = \frac{n}{L}$$

$$= \frac{.587}{.1} = 5.87\,M\ NH_3$$

$$10g\ NH_3 \left|\frac{1mol}{17.034g}\right. = 0.587\ moles\ NH_3$$

232. Describe the procedure a student should use to prepare 250. mL of 0.125 M $CuSO_4(aq)$ using appropriate equipment selected from the list below. Assume that the student uses appropriate safety equipment.

- 250 mL beaker
- 250 mL graduated cylinder
- 250 mL volumetric flask
- Eye dropper
- 500 mL wash bottle filled with distilled water
- 3.000 M $CuSO_4$ in a 50 mL buret

$M_1V_1 = M_2V_2$

$(3)V_1 = (.125)(250)$

$10.4\ mL = V_1$

1. Distribute 10.4 mL of 3.000 M $CuSO_4$ from the buret into the 250 mL volumetric flask.
2. Fill the volumetric flask to the 250 mL line using the wash bottle.
3. Use an eye dropper for the last few drops of distilled water into the flask.?

233. Describe the procedure a student should use to prepare 100. mL of 0.250 M NaOH(aq) using appropriate equipment selected from the list below. Assume that the student uses appropriate safety equipment.

- 100 mL beaker
- 100 mL graduated cylinder
- 100 mL volumetric flask
- Eye dropper
- 500 mL wash bottle filled with distilled water
- Electronic balance
- Weigh boat
- NaOH(s)

$M \times V = n_{NaOH}$

$(.250)(.100) = n_{NaOH}$

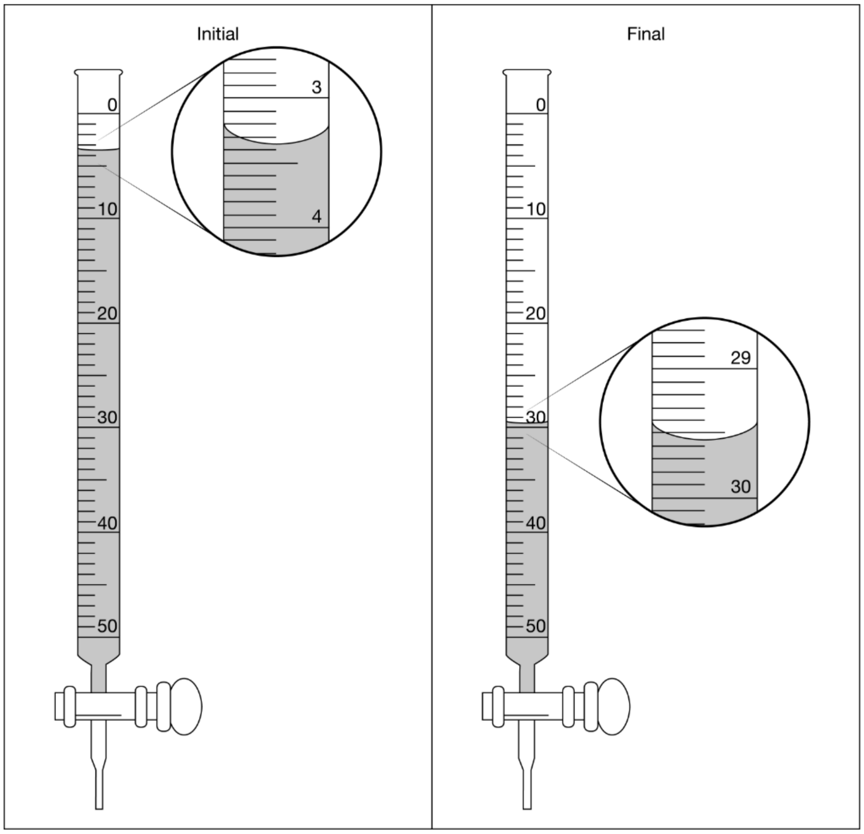
.025 mols NaOH

$\frac{.025\ mols\ NaOH}{} \left| \frac{40\ g}{1\ mol} \right. = 1.00\ g\ NaOH(s)$

1. Weigh out 1.00 g of NaOH(s). Do so by placing a weigh boat on the electronic balance, tare the mass of the weigh boat, and add 1.00 g of NaOH(s).
2. Fill the 100 mL volumetric flask halfway with distilled water.
3. Add the NaOH(s) to the flask. Stir to dissolve the NaOH(s) completely.
4. Fill the volumetric flask to the 100 mL line, using the eye drropper to get exactly 100 mL on the line.

234. A student used a 50.0 mL buret to add $KMnO_4$(aq) to $H_2C_2O_4$(aq) until a faint lavender color was observed in the flask, an indication that the end point of the titration had been reached.

a. The initial and final volume readings of the solution in the buret are shown below. Write down the initial reading and the final reading and use them to determine the volume of $KMnO_4$(aq) that was added during the titration.



Initial = 3.35 mL

Final = 29.55 mL

$$\begin{array}{r} 29.55 \\ -\ 3.35 \\ \hline 26.20 \end{array}$$

26.20 mL delivered

b. Determine the moles of $KMnO_4$ delivered if the molarity of the $KMnO_4$(aq) is 0.320 M.

$M = \frac{n}{V}$

$M \cdot V = n$

$(.320)(.02620) = n$

$.00838$ mols $= n$

235.Perform the following calculations.

a. A 100 mL sample of 0.500 M $NaNO_3(aq)$ solution is mixed with 100 mL of 0.500 M $Ca(NO_3)_2(aq)$ solution. What is the final concentration of the NO_3^- ion?

$100 \times .5M = 50$ mmols NO_3^- from $NaNO_3$

$100 \times .5M \times 2 = 100$ mmols NO_3^- from $Ca(NO_3)_2$

$$\frac{150 \text{ mmols } NO_3^-}{200 \text{ mL}} = .75 \text{ M } NO_3^-$$

b. How many grams of $CaCO_3(s)$ (molar mass 100. g) are needed to make 10. mL of 0.50 M solution?

$10 \text{ mL} \times .50 = 5$ mmols

$.005 \text{ mols} \times 100g = .50g$ $CaCO_3$

c. A 540 mg sample of glucose (molar mass 180 g) is dissolved in enough water to make 300. mL of solution. What would be the molarity of glucose in 100. mL of the solution?

$$\frac{540 \text{ mg glucose}}{} \left| \frac{1 \text{ mol}}{180g} \right. = 3 \text{ mmols or } .003 \text{ mols}$$

$$\frac{.003 \text{ mols}}{.3L} = .01M$$

The 100mL & 300mL solutions would have the same molarity

d. A student dilutes 100. mL of 2.00 M $CaCl_2(aq)$ to a final volume of 400. mL with distilled water.

i. How many moles of chloride ion are in the 100. mL solution?

$M \times V = n_{CaCl_2}$

$2 \times .1 = .2 \text{ mols } CaCl_2 \left| \frac{2Cl^-}{1CaCl_2} \right. = .4 \text{ mols } Cl^-$

ii. How many moles of chloride ion are in the 400. mL solution?

$.4 \text{ mols } Cl^-$

iii. What is the molarity of chloride ion in the 100. mL solution?

$M = \frac{n}{L} = \frac{.4}{.1} = 4.00 M\ Cl^-$

iv. What is the molarity of chloride ion in the 400. mL solution?

$M = \frac{n}{L} = \frac{.4}{.4} = 1.00 M\ Cl^-$

Section 3.09

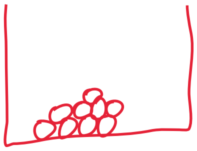
Topic 3.8 First Year Review

236. How many atoms are in one particle of $Ca(OH)_2$?

1 2 2 = 5 atoms

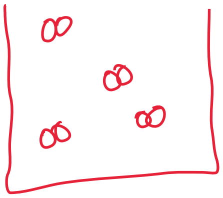
237. Draw a particle picture of the following. In each case, explain your particle picture.

a. Ag(s)



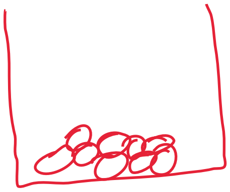
Ag(s) is a single atom and should be arranged in a crystalline structure.

b. $H_2(g)$



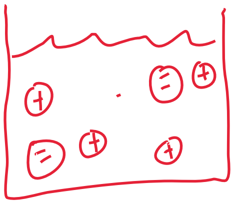
$H_2(g)$ is a diatomic gas. The particles should be all over the container.

c. $C_{12}H_{22}O_{11}(s)$



$C_{12}H_{22}O_{11}(s)$ should have a crystalline structure that is held together with IMF.

d. $K_2S(aq)$

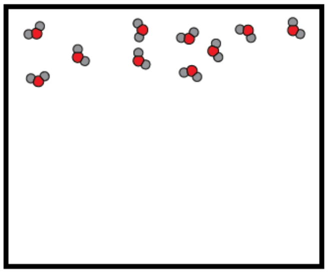
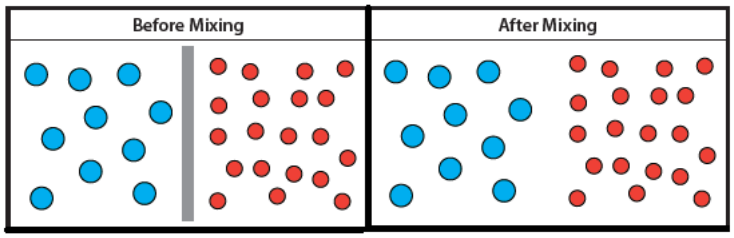
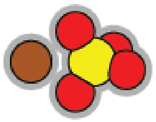
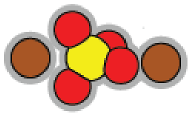
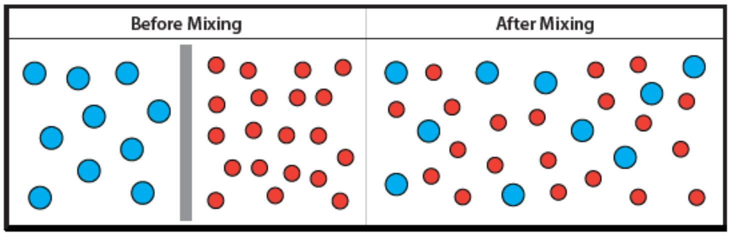
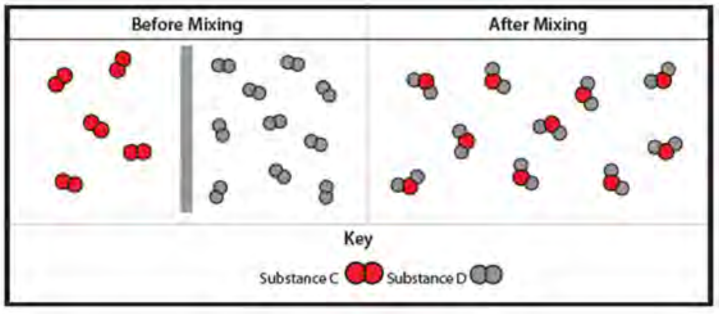


For every two potassium ions you need one sulfide ion.

e. $AlCl_3(s)$

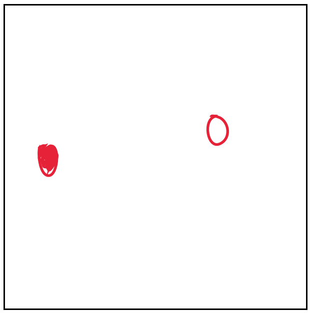
For every one aluminum ion you need three chloride ions.

238. Shown below are different particle pictures. Determine what is wrong with the particle picture and then explain how it should be fixed to be accurate.

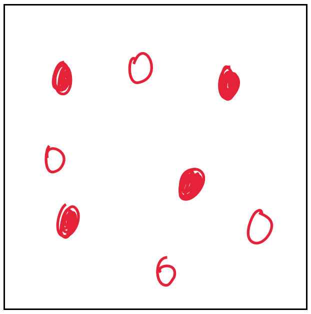
Description of Particle Picture	Particle Picture	How would you fix the picture?
A gas contained in a rigid container.		The gas should be located all throughout the container and not just in the top of the container.
Separate gases allowed to mix.		The gases should be thoroughly mixed and diffused throughout each other, not separated still.
A particle picture of Li_2SO_4.		There should be two lithium ions, not just one. (The lithium ion is brown.)
A particle picture of Iron(II) sulfate.		There should only be one iron ion, not two. (The iron ion is brown.)
A particle picture of a chemical reaction.		This is not a chemical reaction. This would be a physical change. To be a chemical reaction a new substance must be formed.
A particle picture of a physical change		This is a chemical reaction. To be a physical change you should end with substance C and substance D by themselves, not bonded to each other.

239. Use the key below to draw the solutions indicated.

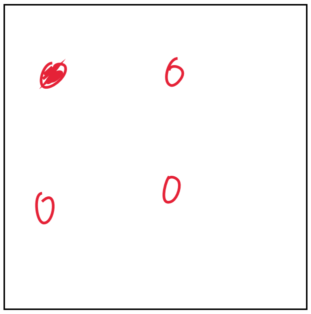
Key
● Any positive ion
○ Any negative ion
H_2O molecules are not shown.



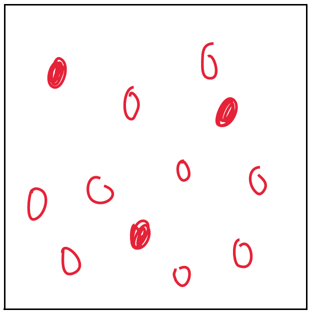
Dilute KCl(aq)



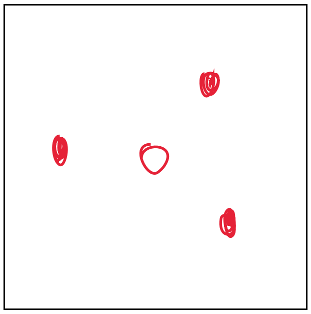
Concentrated KCl(aq)



Dilute $Al(NO_3)_3(aq)$



Concentrated $Al(NO_3)_3(aq)$



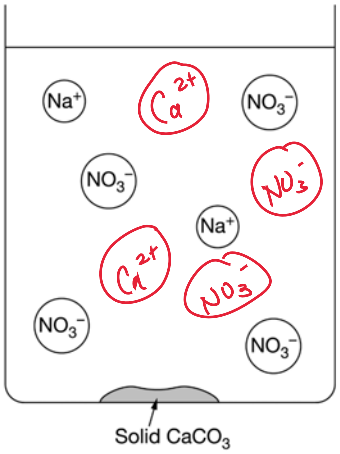
Dilute $Na_3PO_4(aq)$

240. A student mixes 100. mL of $Na_2CO_3(aq)$ with an excess amount of $Ca(NO_3)_2(aq)$, as shown in the equation below.

$$Na_2CO_3(aq) + Ca(NO_3)_2(aq) \rightarrow 2\,NaNO_3(aq) + CaCO_3(s)$$

The diagram below showing the beaker where the reaction takes place is incomplete. Draw in the species needed to accurately represent the major ionic species remaining in the solution after the reaction has been completed.

No Ca^{2+} ions are shown but it is in excess.

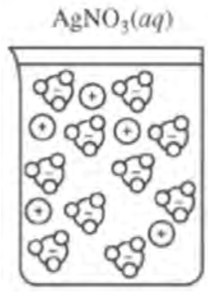


The charge of all the ions must add up to zero.

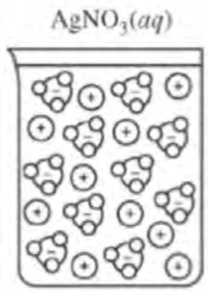
241. Beaker X contains a solution of $AgNO_3$ while Beaker Y contains a solution of $MgCl_2$. Answer the questions that follow about the reaction that occurs when Beaker X and Beaker Y are poured into Beaker Z.

$$\underset{\text{Beaker X}}{2\ AgNO_3(aq)} \quad + \quad \underset{\text{Beaker Y}}{MgCl_2(aq)} \quad \rightarrow \quad \underbrace{2\ AgCl(s) \quad + \quad Mg(NO_3)_2(aq)}_{\text{Beaker Z}}$$

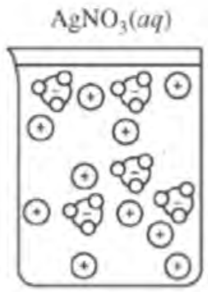
a. Determine which beaker below is correct for Beaker X. Explain your reasoning. Be certain to discuss all beakers in your explanation.



Beaker 1



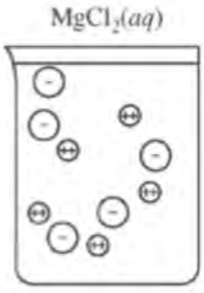
Beaker 2



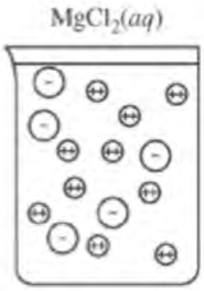
Beaker 3

Beaker X is beaker 2. Beaker 2 has the right ratio of + and - ions, 1:1.
Beaker 1 has a ratio of + to - of 1:2
Beaker 3 has a ratio of + to - of 2:1

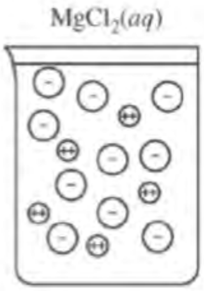
b. Determine which beaker below is correct for Beaker Y. Explain your reasoning. Be certain to discuss all beakers in your explanation.



Beaker 4



Beaker 5



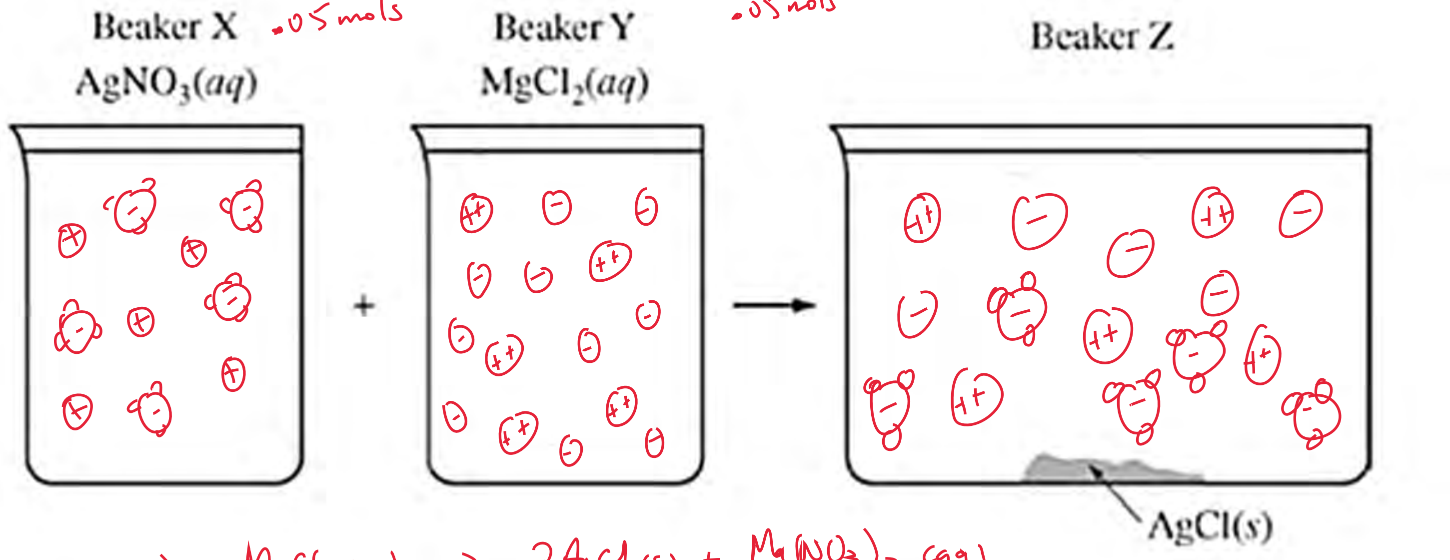
Beaker 6

Beaker Y is beaker 6. Beaker 6 has a ratio of 1++ ion to 2 -, the same as $MgCl_2$, 1 Mg^{2+} : 2 Cl^-
Beaker 4 has a ratio of 1++ to 1-
Beaker 5 has a ratio of 2++ to 1-

242. In letters a – c below amounts of a chemical reaction are given. Use the key below to draw the reactants in Beaker X, the reactants in Beaker Y, and the products in Beaker Z. In each case, calculate the mass of AgCl(s) formed in the reaction.

0.01 mol of Ag^+ 0.01 mol of NO_3^- 0.01 mol of Mg^{2+} 0.01 mol of Cl^-

a. $\underbrace{\text{50. mL of 1.00 M } AgNO_3(aq)}_{.05 \text{ mols}}$ reacts with $\underbrace{\text{50. mL of 1.00 M } MgCl_2}_{.05 \text{ mols}}$.



$2AgNO_3(aq) + MgCl_2(aq) \longrightarrow 2AgCl(s) + Mg(NO_3)_2(aq)$

$.05$ mols $.05$ mols

$.025$ mols Mg^{2+} & $.05$ mols NO_3^-

$$\begin{array}{l|l} .05 \text{ mols } AgNO_3 & Mg(NO_3)_2 \\ \hline & 2AgNO_3 \end{array} = \boxed{.025 \text{ mols } Mg(NO_3)_2}$$

$$\begin{array}{l|l} .05 \text{ mols } MgCl_2 & Mg(NO_3)_2 \\ \hline & MgCl_2 \end{array} = .05 \text{ mols } Mg(NO_3)_2$$

$AgNO_3$ is the L.R.

$$\begin{array}{l|l} .05 \text{ mols } AgNO_3 & MgCl_2 \\ \hline & 2AgNO_3 \end{array} = .025 \text{ mols } MgCl_2 \text{ used}$$

$.05$ mols $MgCl_2$ start
$.025$ mols $MgCl_2$ use
$.025$ mols $MgCl_2$ end
$.025$ mols Mg^{2+} & $.05$ mols Cl^-

Final beaker = No Ag^+
$.05$ mols Mg^{2+} (spectator ion)
$.05$ mol Cl^-
$.05$ mols NO_3^-

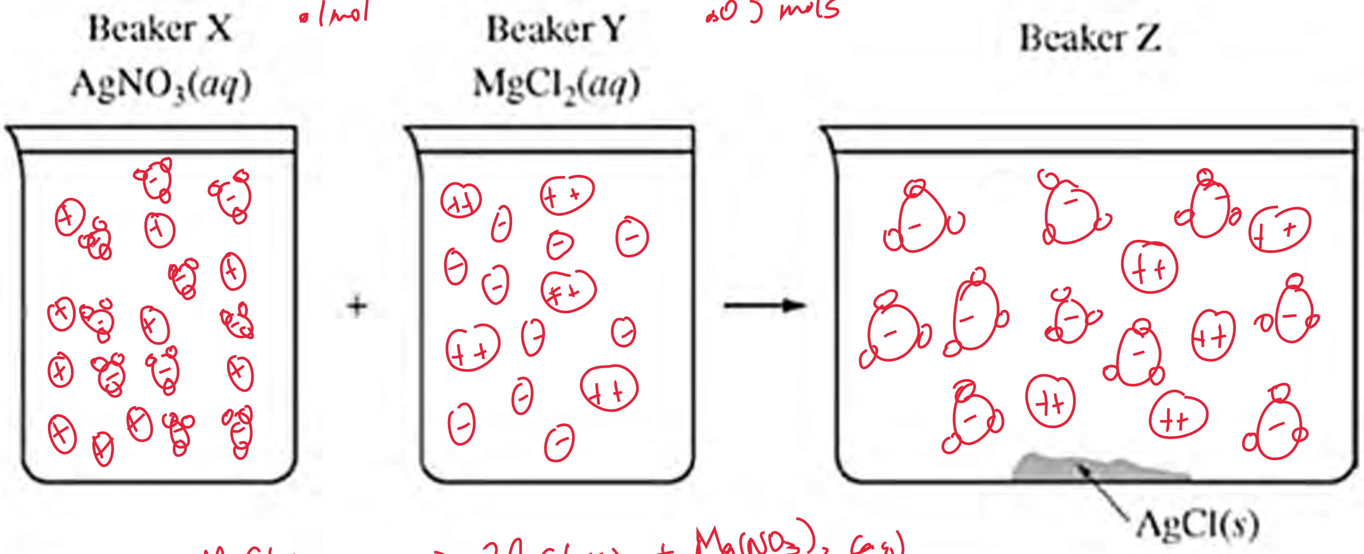
$$\begin{array}{l|l} .05 \text{ mols } AgNO_3 & 2AgCl \\ \hline & 2AgNO_3 \end{array} = .05 \text{ mols } AgCl$$

$$\begin{array}{l|l} .05 \text{ mols } AgCl & 143.32g \\ \hline & 1 \text{ mol} \end{array} = 7.17 \text{ g } AgCl$$

0.01 mol of Ag^+ 0.01 mol of NO_3^- 0.01 mol of Mg^{2+} 0.01 mol of Cl^-

b. 100. mL of 1.00 M $AgNO_3(aq)$ reacts with 50. mL of 1.00 M $MgCl_2$.

.1 mol .05 mols



$2AgNO_{3(aq)} + MgCl_{2(aq)} \rightarrow 2AgCl_{(s)} + Mg(NO_3)_{2(aq)}$

.1 mol .05 mol

$\frac{.1\text{ mol } AgNO_3 \mid 2AgCl}{\mid 2AgNO_3} = \frac{.1\text{ mol } AgCl \mid 143.32\text{g}}{\mid 1\text{mol}} = 14.3\text{g } AgCl_{(s)}$

$\frac{.05\text{ mol } MgCl_2 \mid 2AgCl}{\mid 1MgCl_2} = .1\text{ mol } AgCl$ ← perfect stoich ratio

Final beaker: 0 mols Ag^+

.1 mol NO_3^-

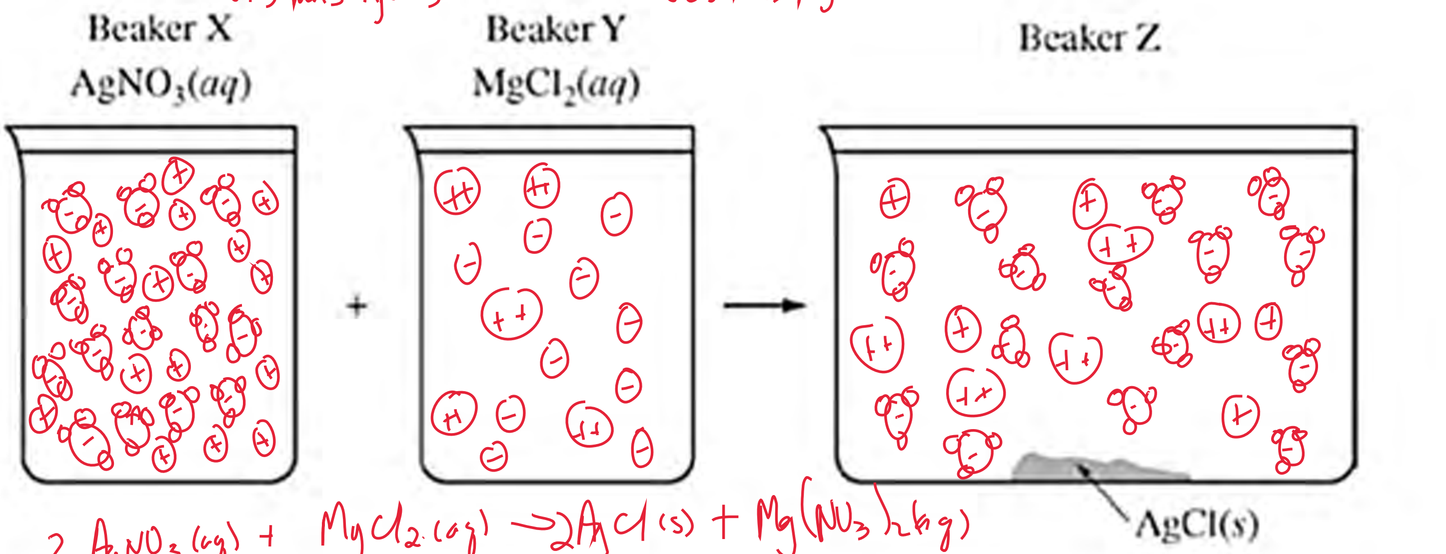
.05 mol Mg^{2+}

0 mol Cl^-

0.01 mol of Ag^+ 0.01 mol of NO_3^- 0.01 mol of Mg^{2+} 0.01 mol of Cl^-

c. 150. mL of 1.00 M $AgNO_3(aq)$ reacts with 50. mL of 1.00 M $MgCl_2$.

.15 mols $AgNO_3$ .05 mols $MgCl_2$



$2\,AgNO_3(aq) + MgCl_2(aq) \rightarrow 2\,AgCl(s) + Mg(NO_3)_2(aq)$

.15 mls .05 mols

$MgCl_2$ is L.R.

$.15 \text{ mols } AgNO_3 \left| \dfrac{2\,AgCl}{2\,AgNO_3} \right. = .15 \text{ mls } AgCl$

$.05 \text{ mols } MgCl_2 \left| \dfrac{2\,AgCl}{1\,MgCl_2} \right. = \boxed{.1 \text{ mols } AgCl} \left| \dfrac{143.32g}{1\,mol} \right. = 14.3g\ AgCl(s)$

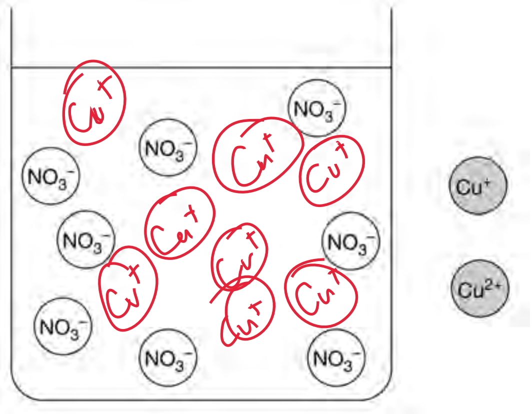
$.05 \text{ mols } MgCl_2 \left| \dfrac{2\,AgNO_3}{1\,MgCl_2} \right. = .1 \text{ mol } AgNO_3 \text{ used}$

.15 mols $AgNO_3$ start
.1 mol $AgNO_3$ used
.05 mol $AgNO_3$ end

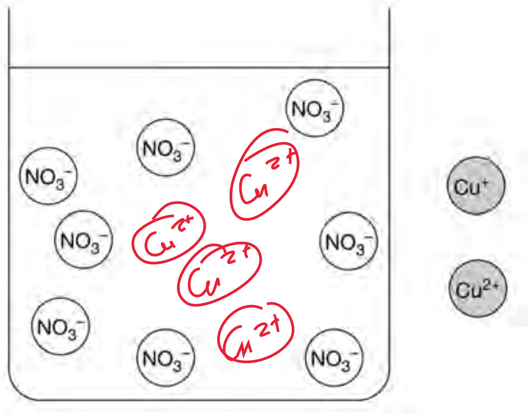
Final beaker
.05 mols Ag^+
.15 mols NO_3^-
.05 mols Mg^{2+}
0 mols Cl^-

243. Use the beakers below to draw an accurate representation of …

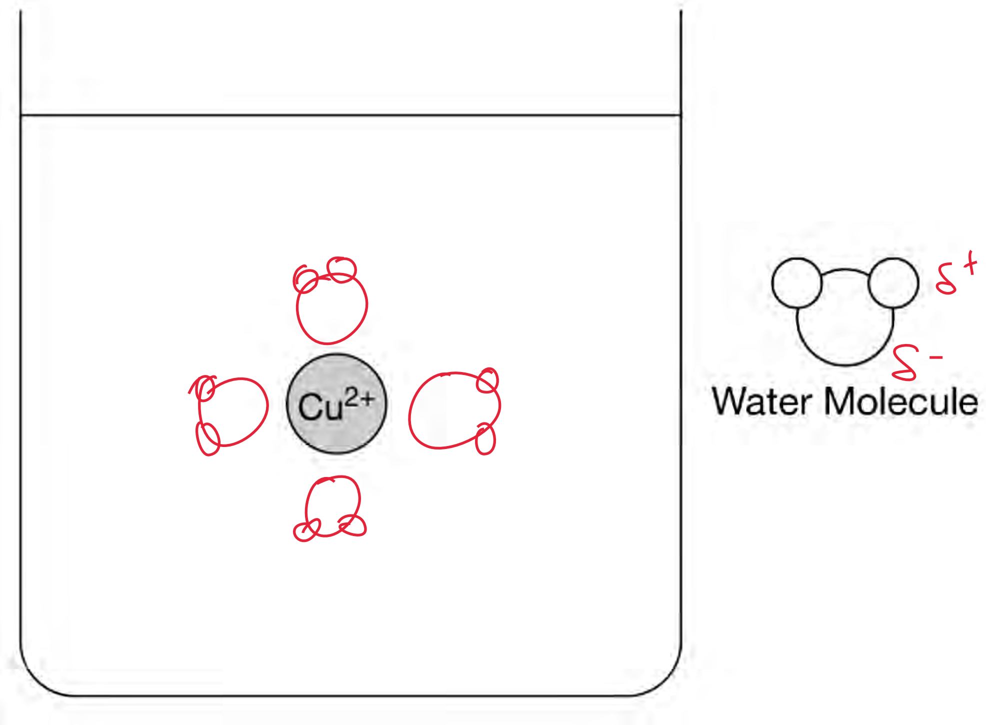
$CuNO_3(aq)$



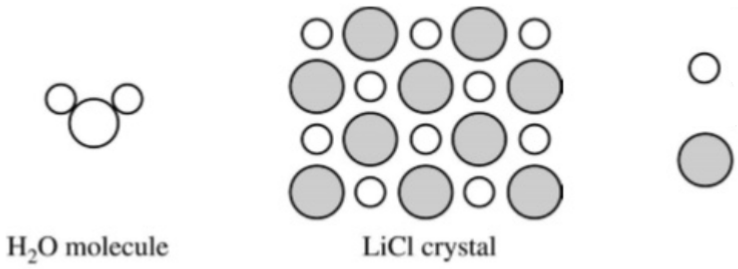
$Cu(NO_3)_2(aq)$



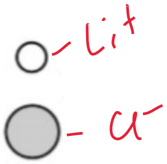
244. In the box below draw the most likely orientation of four $H_2O(l)$ molecules around the Cu^{2+} ion.



245. Shown below is a representation of water and a crystal of LiCl. Answer the questions that follow about a student making a solution of LiCl(aq).



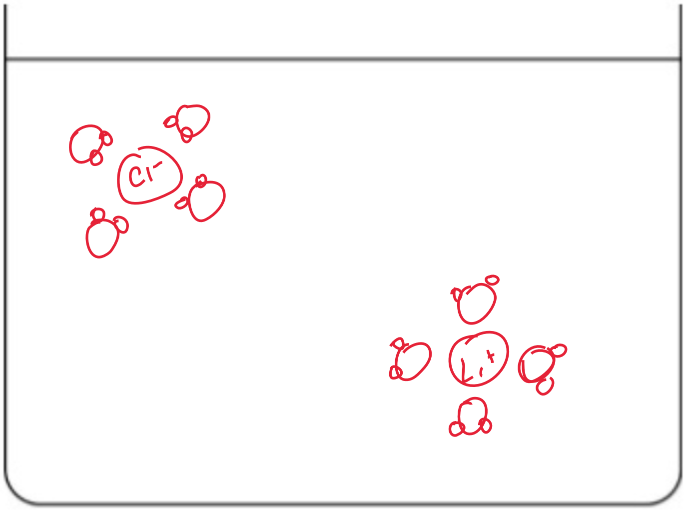
a. Determine the identity of each particle below. Explain your reasoning.



The anion of Cl^- should be larger than Li^+ because it has additional shells of electrons.

b. In the space provided below, show the interactions of the components of LiCl(aq) by making a drawing that represents the different particles present in the solution. Base the particles in your drawing on the particles shown in the representation above. Include only one formula unit of LiCl and no more than ten molecules of water. Your drawing must include the following details:

- Identify of ions (symbol and charge)
- The arrangement and proper orientation of the particles in the solution



LiCl (*aq*)

Topic 3.9 First Year Review

246. Determine if the following would form a homogeneous or heterogenous mixture. To do so, consider the intermolecular forces of the solute and solvent or the solubility of a solid. For the solubility of solids in AP Chemistry, sodium ions, nitrate ions, ammonium ions, and potassium ions are always soluble. All other ions are considered insoluble unless explicitly stated. In each case, explain your reasoning.
    a. A solute of liquid ethanol (CH_3CH_2OH) and a solvent of water (H_2O).

Homogeneous mixture. The solute and solvent have similar IMF and are therefore miscible. Both have London dispersion, dipole-dipole, and hydrogen bonding.

    b. A solute of liquid hexane (C_6H_{14}) and a solvent of water (H_2O).

Heterogeneous mixture. The solute and solvent have different IMF and are therefore not miscible. Hexane only has London dispersion forces while water has London dispersion, dipole-dipole, and hydrogen bonding.

    c. A solute of liquid hexane (C_6H_{14}) and a solvent of liquid octane (C_8H_{18}).

Homogeneous mixture. Both solute and solvent have similar IMF and are therefore miscible. They both only have London dispersion forces.

    d. A solute of ammonium chloride (NH_4Cl) in the solvent water (H_2O).

Homogeneous mixture. The solute has NH_4^+ which is always soluble in water. Therefore NH_4Cl will dissolve in water.

    e. A solute of copper(II) sulfate ($CuSO_4$) in the solvent water (H_2O).

Heterogeneous mixture. The solute copper(II) sulfate is not soluble in water as neither Cu^{2+} nor SO_4^{2-} are considered soluble.

247. When ethanol and methyl iodide are mixed the resulting solution has a boiling point different from ethanol and methyl iodide. However, when ethanol and water are mixed the boiling point of the resulting solution remains the same as ethanol and water, 78 °C and 100 °C respectively.

a. In which situation was a mixture created? Explain how the data supports your conclusion.

Mixing water and ethanol created a mixture. A mixture still retains the properties of the substances that make up that mixture. Since the mixture boils at 78^{o}C (same as ethanol) and 100^{o}C (same as water) there is not a change in properties. No change in properties signifies a mixture.

b. Explain why the boiling point does not change when ethanol and water are mixed.

Since ethanol and water create a mixture and not a chemical reaction there are no new substances formed. As a result, both ethanol and water retain their properties, including their respective boiling point.

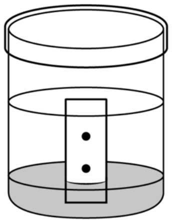
Topic 3.9 Worksheet

248. Briefly describe each separation technique and explain when that technique should be used.

a. Filtration

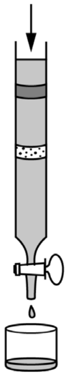
The filter paper traps solid and allows the solution to flow through. Should be used to separate a solid from a solution.

b. Paper Chromatography



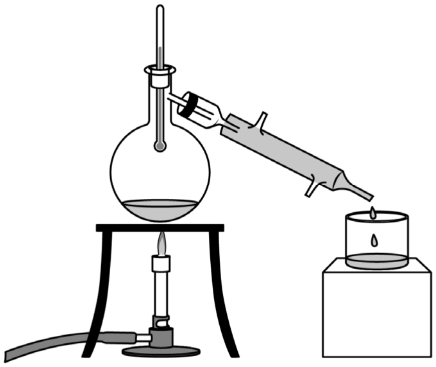
A piece of chromatography paper (stationary phase) has a mixture of liquids placed on it. A solvent (mobile phase) then travels up the paper separating the mixture by polarity and mass. Used to separate a mixture of liquids.

c. Column Chromatography



Same as paper chromatography but with this method you can collect the separated liquids. Used to separate and collect a mixture of liquids.

d. Distillation

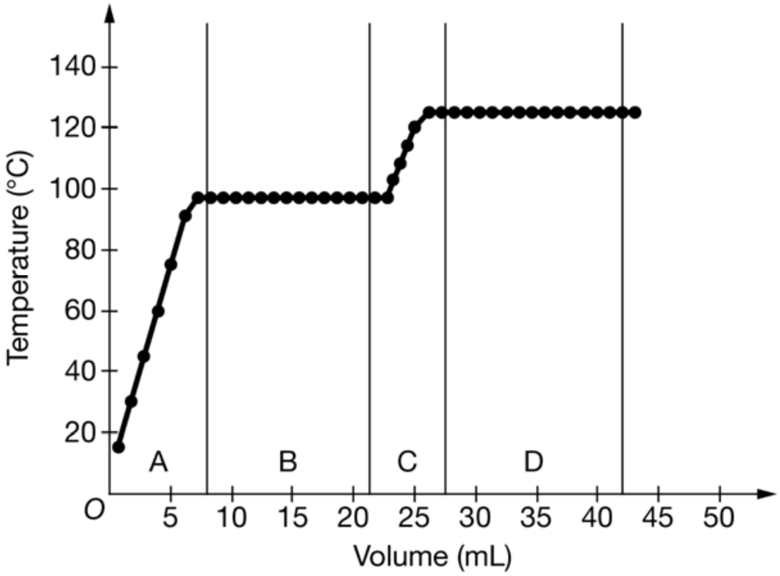


A mixture is heated until one component of that mixture is separate from the other based on a difference in boiling point. If a mixture of liquids, the boiling point is used. Used to separate and collect a mixture of liquids based on boiling point.

249.

A mistake that I'm not willing to go back and change.

250. A student performs a fractional distillation to separate a mixture of two hydrocarbons, C_7H_{16} and C_8H_{18}. Four ranges are shown for which the student collected the distillate, A, B, C, and D.



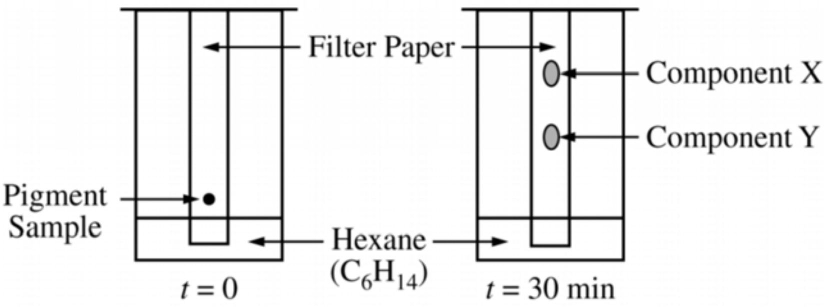
a. Why is there a plateau at both B and D?

Those are the boiling points of the two liquids.

b. Which range corresponds to C_7H_{16}? Explain your reasoning. In your explanation, you must discuss both hydrocarbons.

Range B. C_7H_{16} has weaker IMF than C_8H_{18} therefore C_7H_{16} has a lower boiling point than C_8H_{18}.

251. Consider the paper chromatography experiment shown below. A student is trying to separate a mixture of components X and Y. The paper is slightly polar.



a. What is the mobile phase in this experiment?

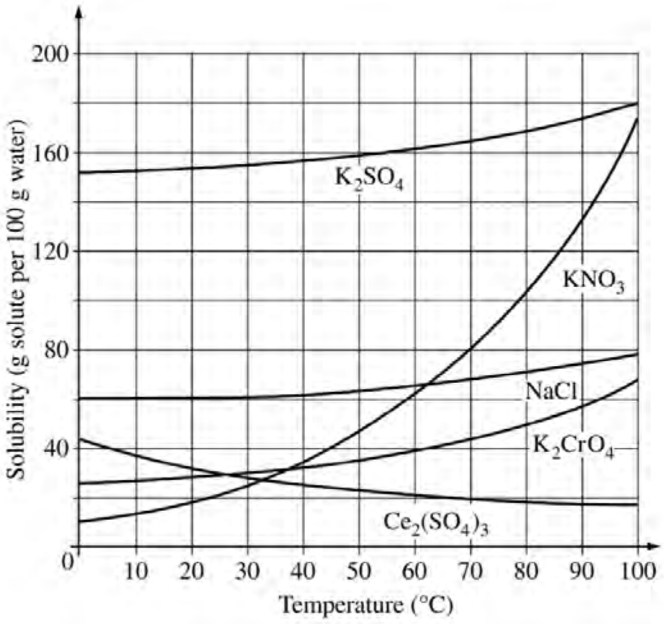
The nonpolar hexane.

b. What is the stationery phase in this experiment?

The filter/chromatography paper.

c. Is Component Y more or less polar than Component X? Explain your reasoning.

Component Y is more polar. The mobile phase is nonpolar therefore the more nonpolar substance will move more.



252. Use the solubility curve above to answer the question that follows. You have a mixture of saturated KNO_3 and saturated NaCl at 90 °C. Which solute will precipitate the most when the temperature is dropped to 70 °C?

The KNO_3 will precipitate the most. At 90^{o}C there is about 130 g of KNO_3 dissolved and about 75 g of NaCl dissolved. At 70^{o}C there are 80 g of KNO_3 dissolved and about 70 g of NaCl dissolved. Therefore 50 g of KNO_3 precipitated and 5 g of NaCl precipitated.

Topic 3.11 First Year Review

253. In the electromagnetic spectrum, what is the relationship between wavelength and energy?

As wavelength increases energy decreases.

254. In the electromagnetic spectrum, what is the relationship between frequency and energy?

As frequency increases energy increases.

255. Which is higher in energy, radio waves or X rays?

X-rays

256. Which is higher in energy, visible light or ultraviolet light?

Ultraviolet

257. Which is higher in energy, infrared or gamma rays?

Gamma rays

258. Describe the molecular motion or electronic transition for each of the energies listed:

a. Microwave radiation

Rotates molecule

b. Infrared radiation

Vibrates molecule/bonds

c. Ultraviolet/visible radiation

Kicks off electron

Section 3.12

Topic 3.12 and 3.13 First Year Review

259. Determine if the following solutions are dilute or concentrated when compared to a 1.00 M sample of the same solution.

a. 245 g of sucrose, $C_{12}H_{22}O_{11}$, in 1.20 L of total solution.

$$\frac{245\text{g sucrose}}{} \Bigg| \frac{1\text{mol}}{342.296\text{g}} = 0.716 \text{mols}$$

$M = \frac{n}{L}$

$= \frac{.716}{1.20}$

$M = 0.596M$

The solution is dilute Compared to the 1.00M solution

MM

$\frac{342.296\text{g}}{\text{mol}}$

b. 85 g of NaOH in 895 mL of total solution.

$$\frac{85\text{g NaOH}}{} \Bigg| \frac{1\text{mol}}{39.998\text{g}} = 2.13 \text{moles}$$

$M = \frac{n}{L}$

$= \frac{2.13}{.895}$

$M = 2.37M$

The solution is more concentrated than the 1.00M Solution

MM

$\frac{39.998\text{g}}{\text{mole}}$

c. 120 g of $Cu(NO_3)_2$ in 450 mL of total solution

$$\frac{120\text{g } Cu(NO_3)_2}{} \Bigg| \frac{1\text{mol}}{187.57\text{g}} = 0.640 \text{mols}$$

$M = \frac{n}{L}$

$= \frac{.640}{.450}$

$M = 1.42M$

The solution is more Concentrated than the 1.00M solution.

MM

$\frac{187.57\text{g}}{\text{mol}}$

260. Determine if there are fewer or more moles of the solute specified for the solutions in 259 above.

a. Which has a greater amount of moles of sucrose, 1.00 L of the solution made in 259.a above or 800 mL of a 1.00 M sucrose solution? Justify your answer with a calculation.

259a

$M = \frac{n}{L}$

$M \times L = n$

$.596 \times 1 = .596$ moles

$M = \frac{n}{L}$

$M \times L = n$

$(1)(.800) = .800$ moles

The 1.00M sucrose solution has more moles of solute.

b. Which has a greater amount of moles of hydroxide ion, OH^-, 500 mL of the solution made in 259.b or 650 mL of 1.0 M $Ca(OH)_2$? Justify your answer with a calculation.

259b

$NaOH = 1\ OH^-/mol$

$M = \frac{n}{L}$

$M \times L = n$

$(2.37)(.5) = 1.185$ moles

$Ca(OH)_2 = 2\ OH^-/mol$

$M = \frac{n}{L}$

$M \times L = n$

$(1)(.65) = .65 \times 2$

1.30 mols OH^-

The 1.0M $Ca(OH)_2$ has more moles of OH^-

c. Which has a greater amount of moles of nitrate, NO_3^-, 50 mL of the solution made in 259.c or 130 mL of a 0.185 M $LiNO_3$ solution? Justify your answer with a calculation.

$Cu(NO_3)_2 = 2\ NO_3^-/mol$

$M = \frac{n}{L}$

$M \times L = n$

$(1.42)(50) = n$

71 mmols $= n \times 2$

142 mmols $NO_3^- = n$

$LiNO_3 = 1\ NO_3^-/mol$

$M = \frac{n}{L}$

$M \times L = n$

$(.185)(130) = n$

24.05 mmols $NO_3^- = n$

The solution made in 259c ($Cu(NO_3)_2$) has more moles of NO_3^-

Topic 3.12 and 3.13 Worksheet

261. Identify all variables in the equations below and give the magnitude and units of any constants.

a. $E = h\nu$

energy (J)

Planck's constant $(6.626 \times 10^{-34}$ J·s)

frequency (Hz)

b. $c = \lambda\nu$

Speed of light 3×10^{8} m/s or 3×10^{17} nm/s

wavelength m/s or nm/s

frequency (Hz)

262. Perform the following calculations.

a. Determine the amount of energy, in J, of a photon with a frequency of 4 x 10^{14} Hz.

$E = h\nu$

$E = (6.626 \times 10^{-34})(4 \times 10^{14})$

$E = 2.65 \times 10^{-19}$ J

b. A photon has 3.3 x 10^{-19} J of energy.

i. What is the frequency of light in Hz?

$E = h\nu$

3.3×10^{-19} J $= (6.626 \times 10^{-34})\nu$

4.98×10^{14} Hz

ii. What is the wavelength of light in nm?

$c = \lambda\nu$

3×10^{8} m/s $= \lambda(4.98 \times 10^{14}$ Hz$)$

602 nm $= \lambda$

263. Identify the variables in the equation below.

path length

Absorbance

$$A = \varepsilon bc$$

Concentration

molar absorptivity

264. Answer the following questions about creating a Beer-Lambert Law plot.

a. What is a cuvette?

A small container that holds the solute. The cuvette is placed into the spectrometer/colorimeter.

b. Why do you need to create a “blank”?

To calibrate the spectrometer to the solvent. It "zeros" out the solvent so that any absorbance from the solvent is eliminated.

c. How do you select a wavelength to measure a colored solution?

You select a wavelength opposite of the colored solution. The wavelength selected should give maximum absorbance.

d. How would you create a standard curve in a Beer’s Law plot?

Determine the absorbance of known molarity solutions. Use at least 5 different molarities.

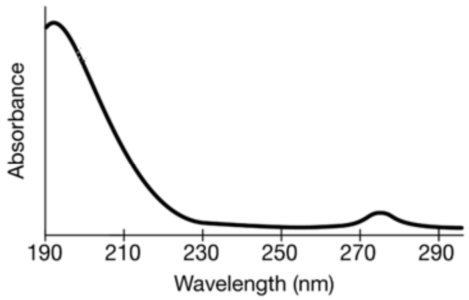


Diagram 1

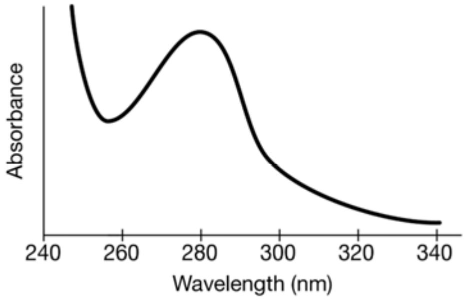


Diagram 2

265. Diagram 1 shows the absorbance of pure acetone, a common solvent. Diagram 2 is the absorbance of a solution of a solute dissolved in acetone.

a. Using diagram 1, explain why it's important to make a "blank" of acetone in the experiment.

The absorbance at 280 nm would be zero if a blank was made and the spectrometer calibrated.
Without a blank the acetone would contribute to the absorbance of the solution.

b. When the student creates Diagram 2 she finds that the absorbance at 280 nm is slightly more than it should be. What is one mistake that could cause the wavelength being greater than it should.

Acetone wasn't used for a blank.

266. Determine if the errors below would increase, decrease, or have no effect on the measured absorbance while creating the standard curve of a Beer-Lambert Law plot. Explain your reasoning.

a. There is some distilled water left in the cuvette when the sample is poured into it.

The absorbance would decrease. The water would dilute the sample in the cuvette. A more dilute solution absorbs less light.

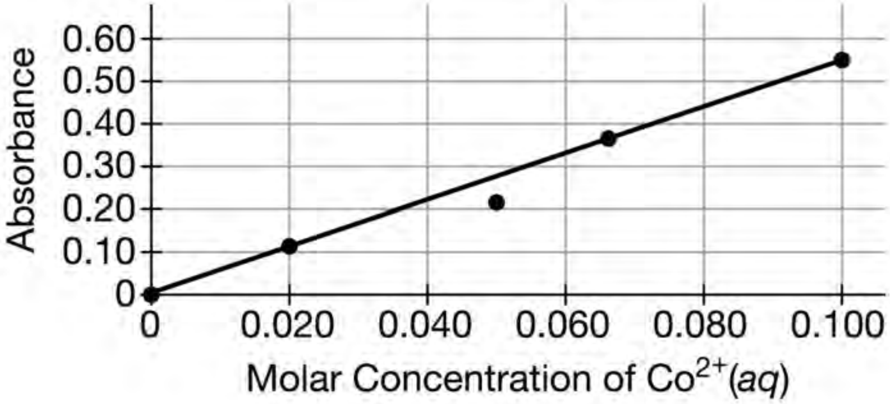
b. The cuvette isn't properly wiped before being placed in the spectrometer.

The absorbance would increase. The dirty cuvette would scatter the light more causing less light to be detected by the spectrometer. Less light detected means greater absorbance.

c. The molarity of the solution is recorded incorrectly.

No change in absorbance. It would however change the standard curve as the molarity would be recorded wrong for that absorbance.

267. A student uses visible spectrophotometry to determine the concentration of $CoCl_2(aq)$ in a sample solution. First the student prepares a set of $CoCl_2(aq)$ solutions of known concentration. Then the student uses a spectrophotometer to determine the absorbance of each of the standard solutions at a wavelength of 510 nm and constructs a standard curve. Finally, the student determines the absorbance of the sample of unknown concentration.



a. The absorbance at 0.050 M is lower than it should be.

i. Is the solution more concentrated or less concentrated than it should be at that point?

Less concentrated.

ii. What could cause this error? Explain your reasoning.

Water left in the cuvette. The water would dilute the solution.

b. The absorbance of the unknown is found to be 0.45.

iii. What is the molarity of $Co^{2+}(aq)$ in the solution?

0.080 M

iv. How many moles of $CoCl_2$ are in 150. mL of the solution?

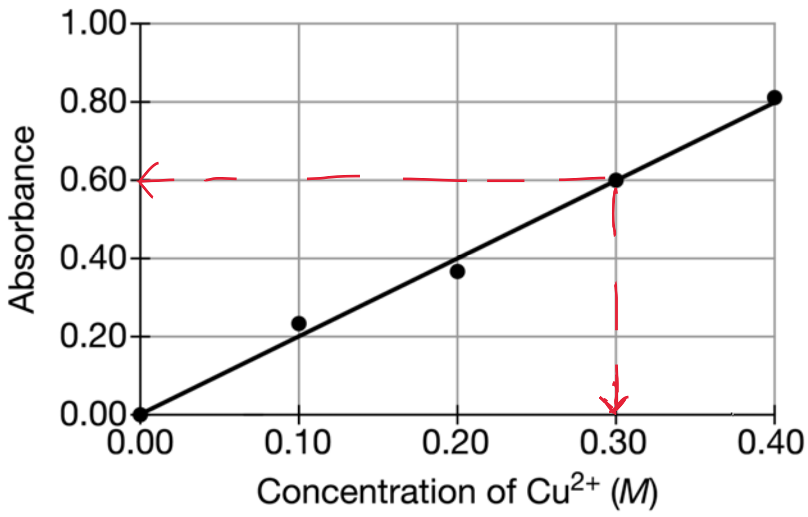
$M = \frac{n}{L}$

$(.080 M)(.150 L) = n$

$M \times L = n$

.012 mols $CoCl_2$

268. A 1.00 g mixture of sodium sulfate, Na_2SO_4, and copper(II) sulfate, $CuSO_4$ is to be analyzed to determine the percent by mass of Na_2SO_4. The 1.00 g sample is dissolved into 10.0 mL of total solution and then analyzed via spectrometry. First, the student prepares a calibration graph by measuring the absorbances of $CuSO_4(aq)$ solutions of known concentrations. The graph is shown below. The impure solid has an absorbance of 0.60.



a. Determine the molarity of Cu^{2+} for the solution.

0.30 M

b. Determine the number of moles of Cu^{2+} in the 10.0 mL sample.

$M = \frac{n}{L}$ $(.30)(.01) = n_{Cu^{2+}}$

$M \times L = n$ $.003 \text{ mols} = n_{Cu^{2+}}$

c. Determine the percent composition by mass of sodium sulfate in the mixture.

$.003 \text{ mols } Cu^{2+} = .003 \text{ mols } CuSO_4$

$.003 \text{ mols } CuSO_4 \times \frac{159.62g}{1 \text{ mol}} = 0.479g\ CuSO_4$

1.00g mixture

− .479g $CuSO_4$

.521g Na_2SO_4

$\frac{.521g}{1.00g} \times 100 = \%$

$52.1\% = \%\ Na_2SO_4$

d. If the original solid mixture was dissolved in 100. mL of total solution instead of 10. mL, would you expect the absorbance to increase, decrease, or remain the same? Explain your reasoning.

I would expect the absorbance to decrease because the solution is more dilute. A more dilute solution will absorb less light according to Beer's Law.

A = abc

"a" and "b" remain the same while "c" decreases.

Unit 4

269. Determine either the phase change or the term for the phase change. Then determine if the intermolecular forces are weakening, breaking, or forming.

	Phase Change			Term	Are the IMF weakening, breaking, or forming?		
(A)	Solid	→	[✓ liquid]	Melting	[✓ Weakening]	Breaking	Forming
(B)	liquid	→	[✓ gas]	Boiling	Weakening	[✓ Breaking]	Forming
(C)	[✓ Gas]	→	Liquid	Condensing	Weakening	Breaking	[✓ Forming]
(D)	solid	→	[✓ Gas]	Sublimation	Weakening	[✓ Breaking]	Forming
(E)	[✓ Gas]	→	Solid	Deposition	Weakening	Breaking	[✓ Forming]
(F)	[✓ liquid]	→	Solid	Freezing	Weakening	Breaking	[✓ Forming]

270. Determine if energy is going into the system or if energy is being released from the system for each of the phase changes from #269. Explain how the energy is being used in each case.

	Energy absorbed or released	Use of energy
(A)	Absorbed	Energy is used to weaken IMF
(B)	Absorbed	Energy is used to break IMF
(C)	Released	Energy is released as IMF are formed
(D)	Absorbed	Energy is used to break IMF
(E)	Released	Energy is released as IMF are formed
(F)	Released	Energy is released as IMF are formed

271. Go back to #269 and circle the phase that has more energy.

Note: This has been done in blue ink.

272.Complete the following reactions. **Balance each reaction**.

Single Replacement Reaction
Metal + Acid

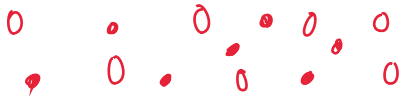
Metal		Acid	→	Ionic Compound	+	Gas
Magnesium +		Hydrochloric acid		Magnesium chloride		Hydrogen
Mg	+	$2HCl$	→	$MgCl_2$	+	H_2
Zinc +		Hydrochloric acid		Zinc chloride		Hydrogen
Zn	+	$2HCl$	→	$ZnCl_2$	+	H_2
Magnesium +		Sulfuric acid		Magnesium sulfate		Hydrogen
Mg	+	H_2SO_4	→	$MgSO_4$	+	H_2

Double Replacement Reaction
Carbonate + Acid

Carbonate		Acid	→	Ionic Compound	+	Water		Gas
Magnesium carbonate		Hydrochloric acid		Magnesium chloride		Water		Carbon dioxide
$MgCO_3$	+	$2HCl$	→	$MgCl_2$	+	H_2O	+	CO_2
Zinc carbonate		Hydrochloric acid		Zinc chloride		Water		Carbon dioxide
$ZnCO_3$	+	$2HCl$	→	$ZnCl_2$	+	H_2O	+	CO_2
Magnesium carbonate		Sulfuric acid		Magnesium sulfate		Water		Carbon dioxide
$MgCO_3$	+	H_2SO_4	→	$MgSO_4$	+	H_2O	+	CO_2

NOTE: A carbonate will react with an acid to form carbonic acid, H_2CO_3(aq). However, carbonic acid is only soluble when under increased pressure of CO_2(g). Without that increased pressure the carbonic acid will break down to form water and carbon dioxide.

Review Topic 1.3 and 1.4

273. Draw a picture of a monatomic element.	274. Draw a picture of a diatomic element.
	
275. Draw a picture of a compound.	276. Draw a picture of a mixture.
	

277. Consider the compound copper(II) acetate, $Cu(C_2H_3O_2)_2$.

a. Determine the percent composition by mass of C.

$$\frac{4\times C}{Cu(C_2H_3O_2)_2} \times 100 = \frac{4\times 12.01}{181.638} \times 100 = 26.45\% \text{ C}$$

b. How many grams of carbon would be in 1.85 g of copper(II) acetate.

$1.85\text{ g} \times 26.45\% = 0.489\text{ g C}$

278. Use the information below to determine the molecular formula of a compound.

a. Determine the empirical formula of the compound that is 62.01% carbon, 13.88% hydrogen, and 24.11% nitrogen.

$$\begin{array}{c|c} 62.01\text{g C} & 1\text{mol} \\ \hline & 12.01 \end{array} = 5.16\text{mol C}/1.72 = 3\text{C}$$

$$\begin{array}{c|c} 13.88\text{g H} & 1\text{mol} \\ \hline & 1.008\text{g} \end{array} = 13.77\text{mol H}/1.72 = 8\text{H}$$

C_3H_8N

$$\begin{array}{c|c} 24.11\text{g N} & 1\text{mol} \\ \hline & 14.01\text{g} \end{array} = 1.72\text{mol N}/1.72 = 1\text{N}$$

b. Determine the molecular formula of the compound if it has a molar mass of 174.3 g mol^{-1}.

$$\frac{MM_{C_3H_8N}}{58.104\,\frac{\text{g}}{\text{mol}}}$$

$$\frac{174.3}{58.104} = 3$$

$$\begin{array}{r} C_3H_8N \\ \times\ 3 \\ \hline C_9H_{24}N_3 \end{array}$$

Section 4.02

Topic 4.2 Worksheet

279. First, balance the equation. Then, give the net-ionic equation for each reaction below. Finally, circle the driving force for each reaction in the net-ionic equation.

a. ______$Co(NO_3)_{2(aq)}$ + __2__$NaOH_{(aq)}$ → __2__$NaNO_{3(aq)}$ + ______$Co(OH)_{2(s)}$

$$Co^{2+}(aq) + 2\,OH^-(aq) \rightarrow \boxed{Co(OH)_2\,(s)}$$

b. ______$H_2SO_4(aq)$ + ______$CaCO_3(aq)$ → ______$CaSO_4(aq)$ + ______$H_2O(l)$ + ______$CO_2(g)$

$$2\,H^+(aq) + CO_3^{2-}(aq) \rightarrow H_2O(l) + \boxed{CO_2\,(g)}$$

c. ______$HC_2H_3O_{2(aq)}$ + ______$NH_{3(aq)}$ → ______$NH_4C_2H_3O_{2(aq)}$ [NOTE: $HC_2H_3O_{2(aq)}$ is a weak acid!]

$$HC_2H_3O_2(aq) + NH_3(aq) \rightarrow NH_4^+(aq) + C_2H_3O_2^-(aq)$$

(no real "driving force" b/c of reasons from unit 8.)

280. Define the following:

a. Non electrolyte

A soluble solute that does not break into ions and does not conduct electricity.

b. Weak electrolyte

A soluble solute that does break into ions but not completely. Appears as mainly molecules. Will conduct electricity slightly producing a dim bulb.

c. Strong Electrolyte

A soluble solute that breaks up into ions 100%. Appears only as ions. Will conduct electricity producing a bright bulb.

281.What types of compounds are nonelectrolytes? Weak electrolytes? Strong electrolytes?

Nonelectrolytes = polar covalent compounds

Weak electrolytes = weak acids

Strong electrolytes = ionic compounds

282.Complete the following equations for an ionic compound being placed in water.

a. $Sr(NO_3)_{2(s)} \rightarrow Sr^{2+}(aq) + 2NO_3^-(aq)$

b. $LiCl_{(s)} \rightarrow Li^+(aq) + Cl^-(aq)$

c. $(NH_4)_2SO_{4(s)} \rightarrow 2NH_4^+(aq) + SO_4^{2-}(aq)$

Review Topic 3.1

283. Describe the formation of each of the following intermolecular forces.
   a. London dispersion forces

Temporary dipoles that result from the random movement of electrons. For any particle, if more electrons are on one side than the other then the particle becomes polar with a positive end and a negative end. The positive end of one particle becomes attracted to the negative end of a different particle.

   b. Dipole-dipole forces

A particle with a permanent dipole is attracted to another particle with a permanent dipole.

   c. Hydrogen bonding

A H bound to a N, O, or F is attracted to a different particle's N, O, or F. It can also form within macromolecules.

   d. Dipole-induced dipole forces

A particle with a permanent dipole causes a particle to temporarily form a dipole (induced dipole). A force of attraction forms between the two particles.

   e. Ion-dipole forces

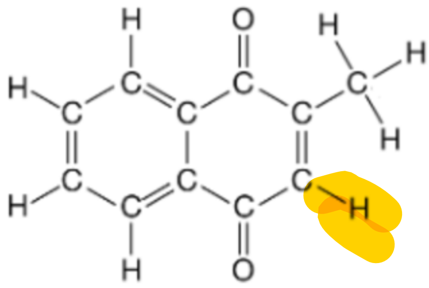
A force of attraction between an ion (a particle with a + or - charge) and a particle with a permanent dipole. The + ion (cation) will be attracted to the negative portion of the dipole while the - ion (anion) will be attracted to the positive portion of the dipole.

284. Two molecules have the same intermolecular forces present. Molecule A has a higher boiling point than Molecule B.
   a. Which molecule has stronger intermolecular forces? Justify your answer.

Molecule A. If Molecule A has a higher boiling point than Molecule B then more energy is required to break the IMF of Molecule A than Molecule B. If more energy is required then the forces of attraction for Molecule A are greater than the forces of attraction for Molecule B.

   b. What is one possible cause for the difference in boiling point?

With same IMF:
1. More areas with hydrogen bonding
2. Stronger dipole
3. More polarizable by either having greater surface area or having more electrons



285. Use the structure of Molecule A above to answer the questions that follow.

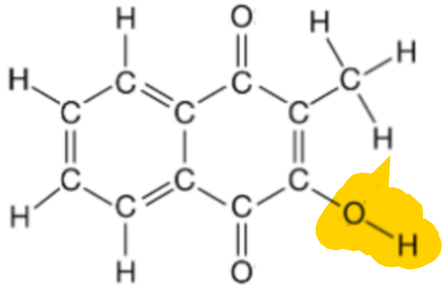
   a. A student claims that the strongest intermolecular force present is hydrogen bonding. Do you agree with the student? Support your answer by referring to the molecule.

I disagree with the student. This molecule does not have any hydrogen bonding. In order to have hydrogen bonding a H must be bonded to a N, O, or F. In this molecule there are no H bonded to a N, O, or F.

   b. Give all intermolecular forces present in the molecule.

The molecule has London dispersion forces and dipole-dipole forces.

   c. Would you expect the boiling point of Molecule B below to be greater than, less than, or equal to the boiling point of the molecule above? Justify your answer.

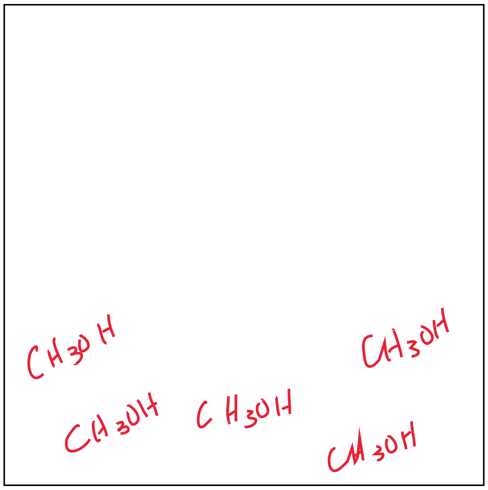


I would expect the boiling point of Molecule B to be greater than the boiling point of Molecule A above. Both molecules are similar but Molecule B has an OH where Molecule A has just an H. (Note: The areas in question have been highlighted yellow on this answer key.) The OH allows Molecule B to exhibit hydrogen bonding which is a stronger IMF. The stronger IMF would lead to greater attraction and more energy required to break the IMF apart.

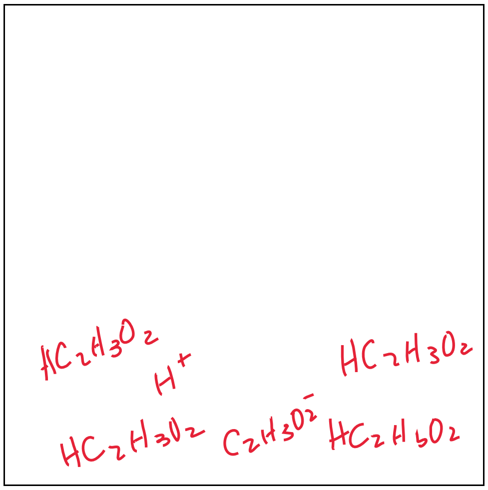
Section 4.03

Topic 4.3 Worksheet

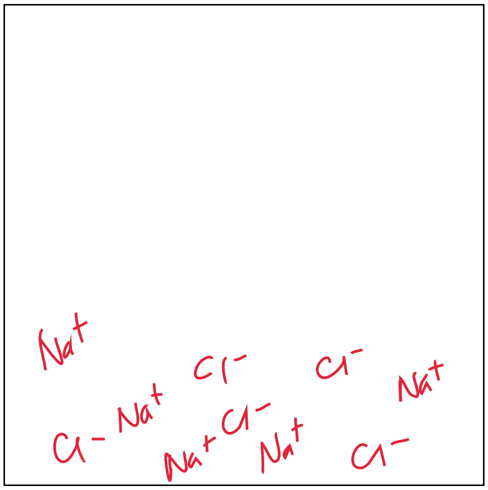
286. Draw a picture of a nonelectrolyte in water (CH_3OH), a weak electrolyte in water ($HC_2H_3O_2$), and a strong electrolyte in water (NaCl). In your drawing include 5 particles of CH_3OH, $HC_2H_3O_2$, and NaCl.



Nonelectrolyte in water

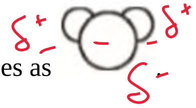


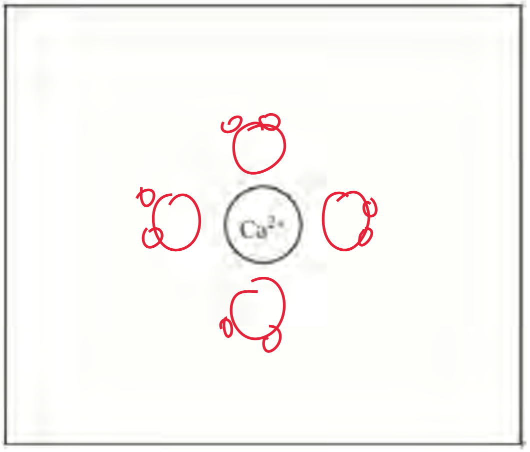
Weak electrolyte in water



Strong electrolyte in water.

287. In the box below, complete a particle representation diagram that includes four water molecules with proper orientation around the Ca^{2+} ion. Represent water molecules as



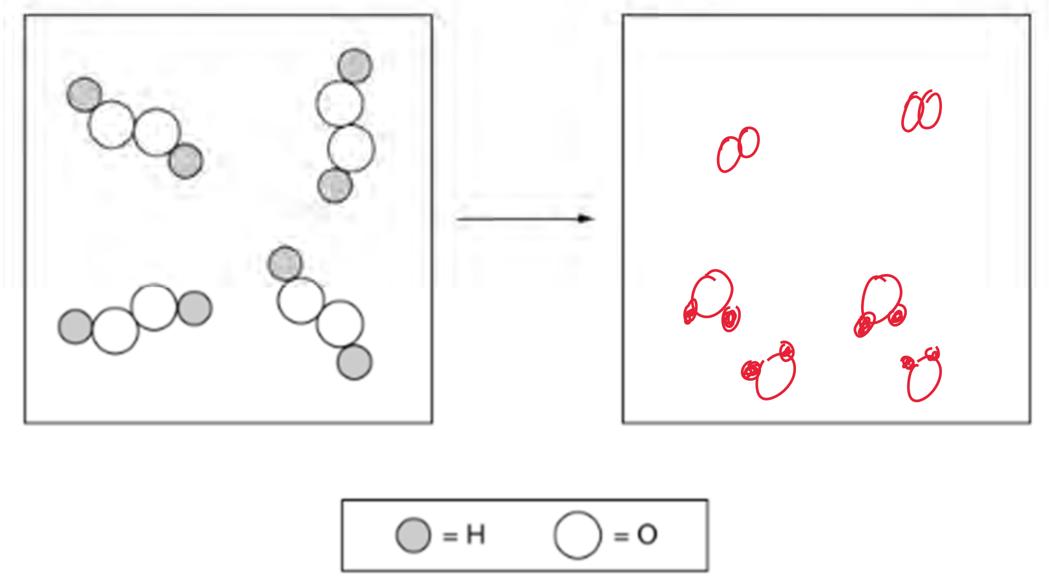


288. Consider the decomposition reaction of hydrogen peroxide, $H_2O_2(aq)$, breaking down to water and oxygen gas.

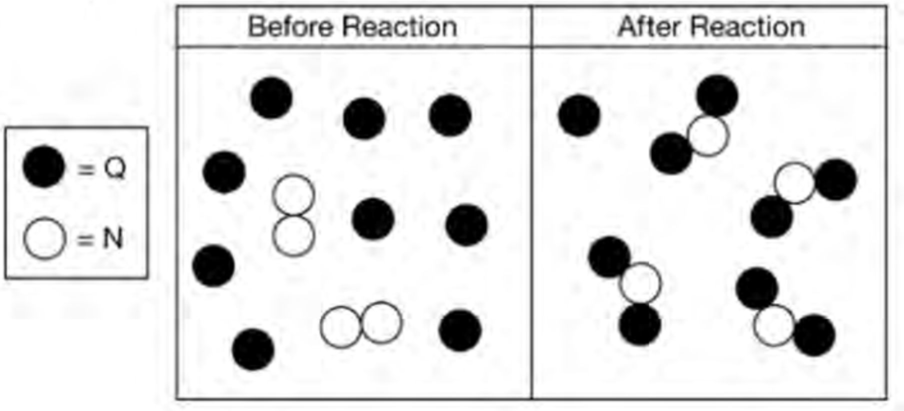
a. Write the balanced chemical reaction that occurs.

$2\ H_2O_2(aq) \longrightarrow 2\ H_2O(l) + O_2(g)$

b. Draw a particle diagram of H_2O_2 breaking down to water and oxygen. The reactants are given in the box below.



289. Consider the particle diagram shown below:



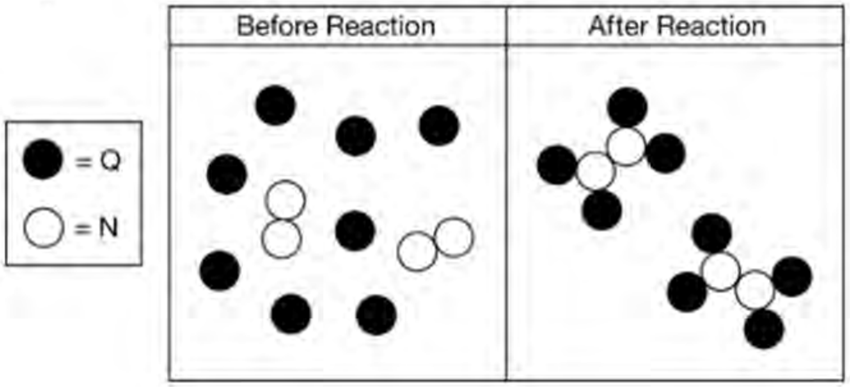
a. Write the balanced chemical equation that takes place between Q and N.

$4Q + N_2 \longrightarrow 2\ Q_2N$

b. Which is the limiting reactant, Q or N? Justify your answer by referring to the particle diagram

N_2 is the limiting reactant. In the "After Reaction" box there is still an unreacted Q particle. Also, there are no N_2 particles remaining.

290.Consider the particle diagram shown below:



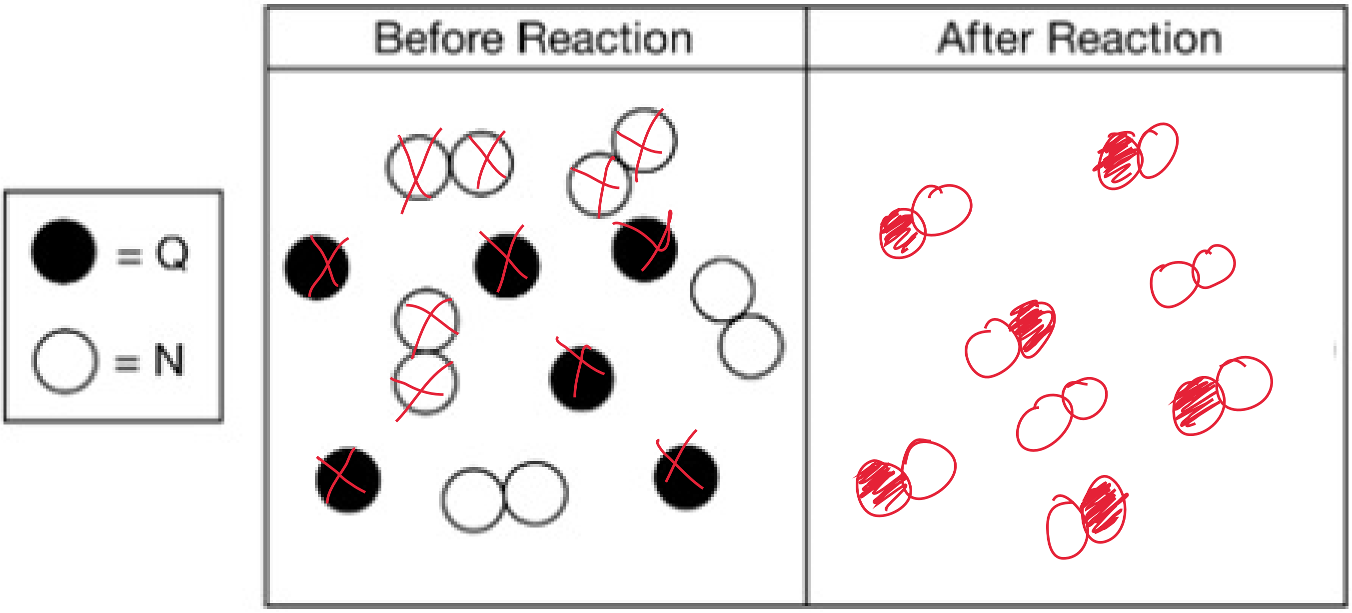
a. Write the balanced chemical equation that takes place between Q and N.

$4Q + N_2 \longrightarrow Q_4N_2$

b. A student claims that the particle diagram represents a perfect stoichiometric ratio. Do you agree or disagree with the student. Support your answer by referring to the particle diagram.

This does represent a perfect stoichiometric ratio of reactants. There are no unreacted particles remaining in the "After Reaction" container.

291. Shown below are the reactants for the unbalanced chemical reaction $Q + N_2 \rightarrow NQ$. Draw the particles in the "After Reaction" box.



Review Topic 2.5 & 2.7

Bonding Domains around central atom	Nonbonding Domains around central atom	Total # of Electron Domains around central atom	Electron Domain Geometry	Molecular Geometry	Example of a Molecule	Example of a Polyatomic Ion	Hybridization around central atom
2	0	2	Linear	linear	CO_2	NO_2^+	sp
3	0	3	Trigonal planar	trigonal planar	SO_3	NO_3^-	sp^2
2	1	3	Trigonal planar	bent	SO_2	NO_2^-	sp^2
4	0	4	Tetrahedral	tetrahedral	CF_4	SO_4^{2-}	sp^3
3	1	4	Tetrahedral	trigonal pyramid	PF_3	ClO_3^-	sp^3
2	2	4	Tetrahedral	bent	SF_2	ClO_2^-	sp^3
5	0	5	Trigonal bipyramidal	trigonal bipyramid	PF_5	SF_5^+	sp^3d
4	1	5	Trigonal bipyramidal	See-saw	SF_4	PF_4^-	sp^3d
3	2	5	Trigonal bipyramidal	t-shaped	ClF_3	N/A	sp^3d
2	3	5	Trigonal bipyramidal	linear	XeF_2	Br_3^-	sp^3d
6	0	6	Octahedral	octahedral	SF_6	PF_6^-	sp^3d^2
5	1	6	Octahedral	square pyramid	ClF_5	N/A	sp^3d^2
4	2	6	Octahedral	square planar	XeF_4	ClF_4^-	sp^3d^2

In the “Molecular Geometry” column, write one of the following molecular shapes in the appropriate spot in the table. Note that some terms may be used more than once.

bent	seesaw	T-shaped	trigonal bipyramidal
linear	square planar	tetrahedral	trigonal pyramidal
octahedral	square pyramidal	trigonal planar	

In the “Example of a Molecule” column, write one of the following chemical formulas in the appropriate spot in the table.

CO_2	ClF_3	PF_3	SF_2	SO_2	XeF_2
CF_4	ClF_5	PF_5	SF_4	SO_3	XeF_4
			SF_6		

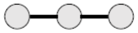
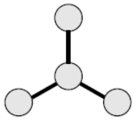
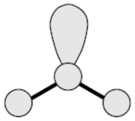
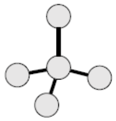
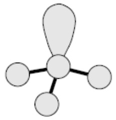
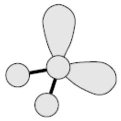
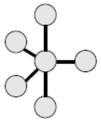
In the “Example of a Polyatomic Ion” column, write one of the following chemical formulas in the appropriate spot in the table.

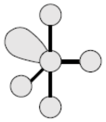
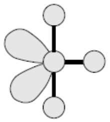
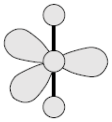
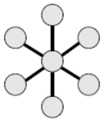
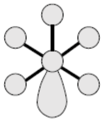
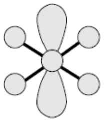
Br_3^-	ClO_2^-	NO_2^+	PF_4^-	SO_4^{2-}
	ClO_3^-	NO_2^-	PF_6^-	SF_5^+
	ClF_4^-	NO_3^-		

In the “Hybridization around central atom” column, write one of the following in the appropriate spot in the table. Note that some terms may be used more than once.

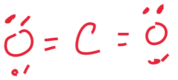
sp	sp^2	sp^3	sp^3d	sp^3d^2
			(Not tested on exam)	(Not tested on exam)

Complete the following table.

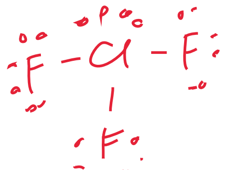
Molecular Geometry	# of Bonding Domains around central atom	# of Nonbonding Domains around central atom	Name of Molecular Shape	All Bond Angles	Hybridization around central atom
	2	0	linear	180°	sp
	3	0	trigonal planar	120°	sp^2
	2	1	bent	< 120°	sp^2
	4	0	tetrahedral	109.5°	sp^3
	3	1	trigonal pyramid	< 109.5°	sp^3
	2	2	bent	<< 109.5°	sp^3
	5	0	trigonal bipyramid	120° 90°	sp^3d

	4	1	See-saw	∠120° 90°	Sp^3d
	3	2	T-shaped	90°	Sp^3d
	2	3	linear	180°	Sp^3d
	6	0	octahedral	90°	Sp^3d^2
	5	1	Square pyramid	90°	Sp^3d^2
	4	2	Square planar	90°	Sp^3d^2

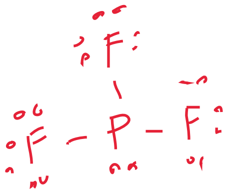
CO_2



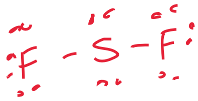
ClF_3



PF_3

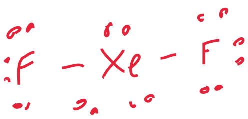


SF_2

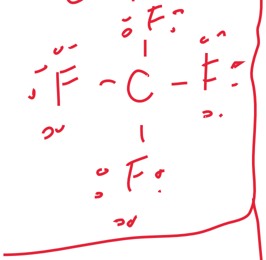


SO_2

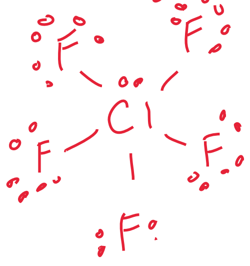
XeF_2

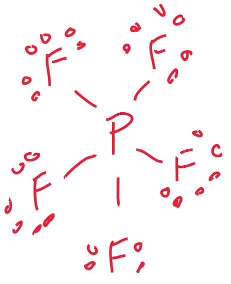


CF_4

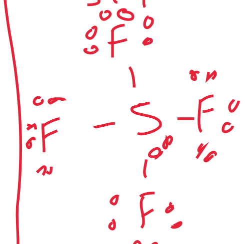


ClF_5

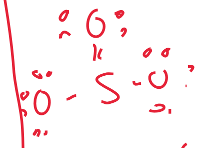




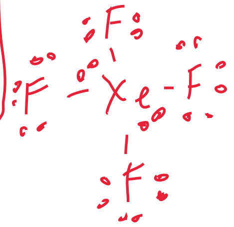
SF_4



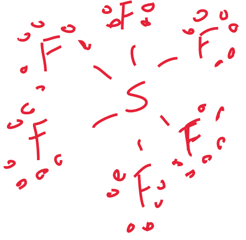
SO_3



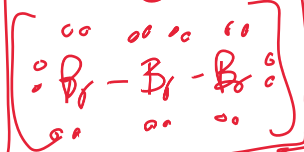
XeF_4



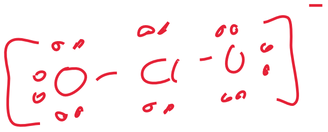
SF_6



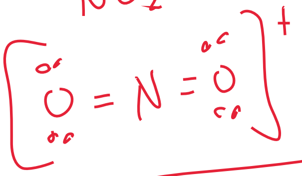
Br_3^-



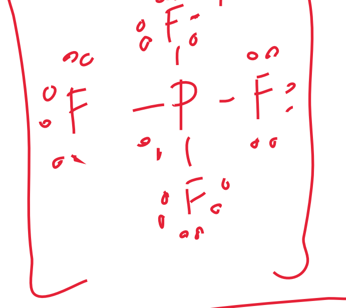
ClO_2^-



NO_2^+

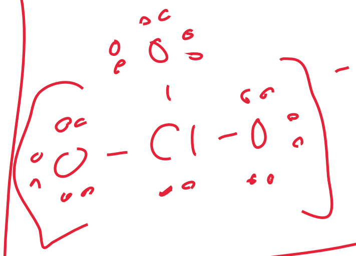


PF_4^-

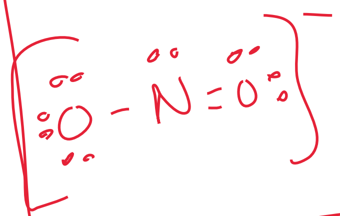


SO_4^{2-}

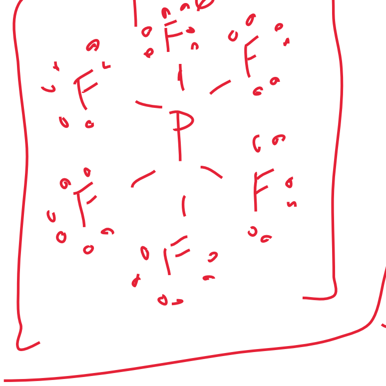
ClO_3^-



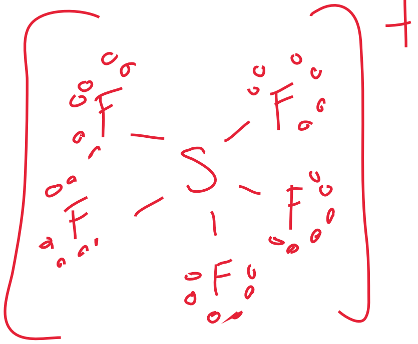
NO_2^-



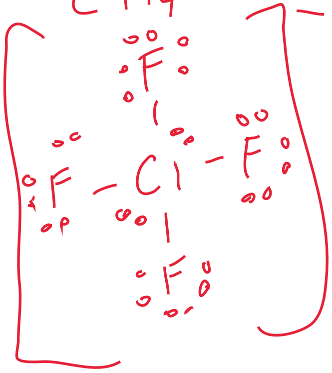
PF_6^-



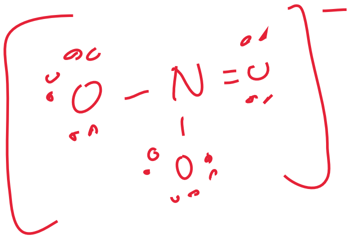
SF_5^+



ClF_4^-



NO_3^-



Topic 4.4 Worksheet

292. Discuss the changes in forces of attraction that occur during a physical change.

The IMF are either weakened (solid to liquid), broken (liquid to gas), or formed (gas to liquid or liquid to solid).

293. Discuss the changes in forces of attraction that occur during a chemical change.

The bonds, either covalent or ionic, are broken or formed.

Review Topic 2.4

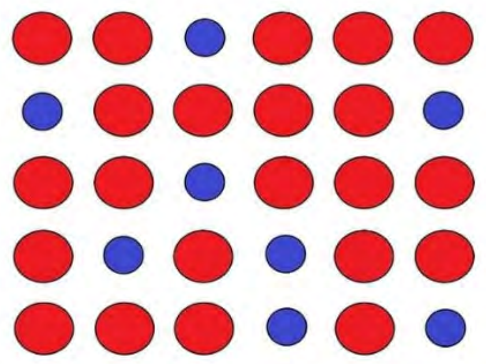
294. What are the relative sizes of atoms in an interstitial alloy?

Large atoms and smalls atoms in between.

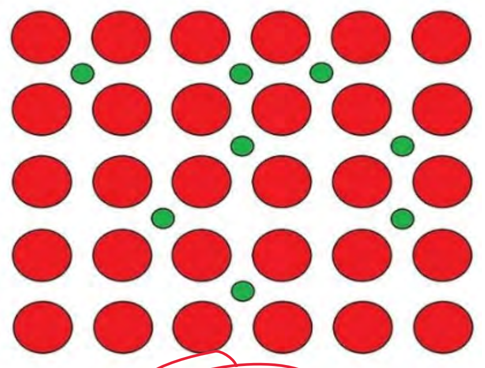
295. What are the relative sizes of atoms in a substitutional alloy?

About the same size.

296. In the pictures below, circle the interstitial alloy and the substitutional alloy.



Interstitial Alloy
or
[✓ Substitutional Alloy]



[✓ Interstitial Alloy]
or
Substitutional Alloy

297. Would you expect the density of an interstitial alloy to be greater than, less than, or equal to the density of the original metal? Explain your reasoning.

The density should be greater. There are more particles in the same amount of space since there are particles between the original metals particles.

298. Which type of alloy is more malleable and why?

Substitutional alloy is more malleable than the interstitial. The particles between the original metal in the interstitial makes it more resistant to being bent.

Section 4.05

Topic 4.5 Worksheet

299. According to the balanced equation below, how many moles of HI would be necessary to produce 2.5 mol of I_2, starting with 4.0 mol of $KMnO_4$ and 3.0 mol of H_2SO_4?

$$10\ HI + 2\ KMnO_4 + 3\ H_2SO_4 \rightarrow 5\ I_2 + 2\ MnSO_4 + K_2SO_4 + 8\ H_2O$$

$$\begin{array}{c|c} 2.5\text{mol } I_2 & 10HI \\ \hline & 5I_2 \end{array} = 5\text{ mol } HI$$

300. According to the reaction represented below, how many grams of aluminum (atomic mass 27 g) are necessary to produce 0.50 mol of hydrogen gas at 25 °C and 1.00 atm?

$$2\ Al_{(s)} + 6\ HCl_{(aq)} \rightarrow 2\ AlCl_{3(aq)} + 3\ H_{2(g)}$$

$$\begin{array}{c|c} .5\text{mol } H_2 & 2Al \\ \hline & 3H_2 \end{array} = .333\text{mol } Al$$

$$\begin{array}{c|c} .333\text{mol } Al & 26.98g \\ \hline & 1\text{mol} \end{array} = 8.98g\ Al$$

301. According to the balanced equation below, how many moles of $ClO_2^-{}_{(aq)}$ are needed to react completely with 20. mL of 0.20 M $KMnO_4$ solution?

$$2\ H_2O(l) + 4\ MnO_4^-(aq) + 3\ ClO_2^-{}_{(aq)} \rightarrow 4\ MnO_{2(s)} + 3\ ClO_4^-{}_{(aq)} + 4\ OH^-{}_{(aq)}$$

MnO_4^-

$M = \frac{n}{L}$

$M \times L = n$

$$(.20)(.02) = \begin{array}{c|c} .004\text{moles } MnO_4^- & 3ClO_2^- \\ \hline & 4MnO_4^- \end{array} = .003\text{mols } ClO_2^-$$

$$2\ KClO_{3(s)} \xrightarrow[MnO_2]{\Delta} 2\ KCl_{(s)} + 3\ O_{2(g)}$$

302. According to the equation above, how many moles of potassium chlorate, $KClO_3$, must be decomposed to generate 1.0 L of O_2 gas at standard temperature and pressure? (This multiple choice gives you an idea of the types of answers that show up from time to time on the AP exam.)

A $\frac{1}{3}\left(\frac{1}{22.4}\right)$ mol

B $\frac{1}{2}\left(\frac{1}{22.4}\right)$ mol

[✓ C $\frac{2}{3}\left(\frac{1}{22.4}\right)$ mol]

D $\frac{3}{2}\left(\frac{1}{22.4}\right)$ mol

E $2\left(\frac{1}{22.4}\right)$ mol

$\frac{1L\ O_2 \mid 1 mol}{\mid 22.4L} = \frac{1}{22.4}$ mols O_2 $\frac{\mid 2\ KClO_3}{\mid 3O_2}$

303. Acetic acid and sodium bicarbonate are reacted and the gas collected.

$$HC_2H_3O_{2(aq)} + NaHCO_{3(s)} \rightarrow NaC_2H_3O_{2(aq)} + H_2O_{(l)} + CO_{2(g)}$$

L.R.

actually formed

a. Determine the volume of CO_2 produced when 2.50 g of $NaHCO_3$ reacts with 55.0 mL of 0.875 M acetic acid at STP.

$\frac{2.50g\ NaHCO_3 \mid 1 mol}{\mid 84.008g} = \frac{0.0298 mols\ NaHCO_3 \mid 1\ CO_2}{\mid 1 NaHCO_3} = 0.0298$ mols CO_2

$\frac{.0298 mols\ CO_2 \mid 22.4L}{\mid 1 mol} =$

$M \times L = n$

$(.875)(.055) = \frac{.0487 mols\ HC_2H_3O_2 \mid 1\ CO_2}{\mid 1\ HC_2H_3O_2} = .0487$ mols CO_2

$\boxed{0.667\ L}$

b. What mass of sodium bicarbonate is required to produce 19.0 L of carbon dioxide gas at 20 °C and 1.2 atm of pressure?

$\frac{PV}{RT} = n$

$\frac{(1.2)(19.0)}{(.08206)(293)} = n = \frac{.948 mls\ CO_2 \mid 1 NaHCO_3}{\mid 1\ CO_2} = .948$ mols $NaHCO_3$

$\frac{.948 mols\ NaHCO_3 \mid 84.008g}{\mid 1 mol} = 79.7g\ NaHCO_3$

304. A 5.000 g sample of an organic hydrocarbon is combusted and the products measured. In the reaction, 15.37 g of carbon dioxide and 7.186 g of water are produced. Assuming the oxygen used for the combustion was in excess, determine the empirical formula of the hydrocarbon.

$$\left(\frac{12}{44}\right) 15.37 = 4.19 \text{ g C}$$

$$\left(\frac{2}{18}\right) 7.186 = .798 \text{ g H}$$

$$\frac{4.19 \text{ g C} \mid 1 \text{ mol}}{\mid 12.01 \text{ g}} = .349 \text{ mol C} / .349 = 1 \text{ C}$$

$$\frac{.798 \text{ g H} \mid 1 \text{ mol}}{\mid 1.008 \text{ g}} = .792 \text{ mol H} / .349 = 2.25 \text{ H}$$

$$CH_{2.25}$$

$$\frac{\times 4}{C_4H_9}$$

305. An organic compound, containing only C, H, and O, is analyzed via combustion analysis. A 1.875 g sample of the compound is combusted and 3.834 g of CO_2(g) and 1.177 g of H_2O(l) is collected. Determine the empirical formula of the compound.

$C_xH_yO_z + O_2 \longrightarrow CO_2 + H_2O$

$\underline{CO_2}$

$3.834\text{g } CO_2 \times \frac{12.01}{44.01} = 1.05\text{g C}$

$\underline{H_2O}$

$1.177\text{g } H_2O \times \frac{2.016}{18.016} = .132\text{g H}$

1.875g
-1.05g C
$-.132\text{g H}$
$.693\text{g O}$

$\frac{1.05\text{g C}}{} \Big| \frac{1\text{mol}}{12.01\text{g}} = .0874\text{mol C} / .0433 = 2C$

$\frac{.132\text{g H}}{} \Big| \frac{1\text{mol}}{1.008\text{g}} = .131\text{mol H} / .0433 = 3H$

C_2H_3O

$\frac{.693\text{g O}}{} \Big| \frac{1\text{mol}}{16\text{g}} = .0433\text{mol O} / .0433 = 1O$

306. A hydrocarbon undergoes combustion analysis to determine the empirical formula of the compound. After complete combustion it is determined that there are 66 g of CO_2 and 36 g of H_2O. Determine the empirical formula of the hydrocarbon.

$$66g\ CO_2 \times \frac{1mol}{44g} = 1.5mol\ CO_2 \times \frac{1C}{1CO_2} = 1.5mol\ C$$

$$36g\ H_2O \times \frac{1mol}{18.01g} = 2mol\ H_2O \times \frac{2H}{1H_2O} = 4mol\ H$$

$C_{1.5}H_4$

×2

C_3H_8

example: using moles instead of % by mass

OR

$$66g\ CO_2 \times \frac{12.01}{44.01} = 18.01g\ C \times \frac{1mol}{12.01g} = 1.50mol\ C$$

$$36g\ H_2O \times \frac{2.016}{18.016} = 4.02g\ H \times \frac{1mol}{1.008g} = 4.00mol\ H$$

$C_{1.5}H_4$

×2

C_3H_8

307. A 3.00 g sample of $MgSO_4 \cdot xH_2O$ hydrate is thoroughly heated. The data below is collected.

Mass of test tube	24.310 g
Mass of test tube + hydrate	27.330
Mass of test tube + hydrate after 1st heating	26.320
Mass of test tube + hydrate after 2nd heating	25.852
Mass of test tube + hydrate after 3rd heating	25.850

a. Explain why the test tube was heated three times.

To make sure all of teh water is driven off. Once the mass of the substance no longer changes, or changes by a negligible amount, all water has been driven off.

b. Determine the mass of water in the hydrate.

$$\begin{array}{r} 27.330 \\ -\ 25.850 \\ \hline 1.480\text{g } H_2O \end{array}$$

c. Determine the ratio of moles of water to moles of anhydrate.

$$\begin{array}{l|l} 1.480\text{g } H_2O & 1\text{mol} \\ \hline & 18.016\text{g} \end{array} = 0.0821 \text{ mols } H_2O$$

$$\frac{.0821}{.0128} = 6.4 \qquad 1\ MgSO_4 : 6.4\ H_2O$$

$$\begin{array}{r} 25.850 \\ -\ 24.310 \\ \hline \end{array}$$

$$\begin{array}{l|l} 1.54\text{g } MgSO_4 & 1\text{mol} \\ \hline & 120.38\text{g} \end{array} = .0128 \text{ mols } MgSO_4$$

d. Determine the formula of the hydrate.

$MgSO_4 \cdot 6.4\ H_2O$

e. The hydrate is $MgSO_4 \cdot 7H_2O$.

i. Determine the percent yield of water driven off.

$$\frac{7H_2O}{MgSO_4 \cdot 7H_2O} \times 100 = 51.15\%\ H_2O$$

mass of water driven off in expt.

↓

$$\frac{1.480}{1.545} \times 100 = 95.79\%\ \text{yield}$$

$3.02g \times .5115 = 1.545g\ H_2O$

↑ mass of hydrate

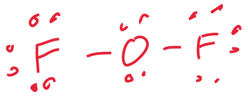
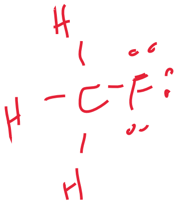
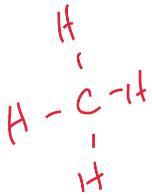
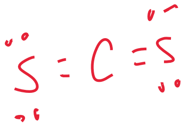
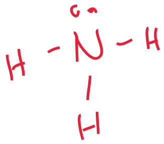
↑ amount of water that should have been driven off

ii. Determine the percent error of the molar mass of the hydrate.

$$\frac{(MgSO_4 \cdot 6.4H_2O) - (MgSO_4 \cdot 7H_2O)}{MgSO_4 \cdot 7H_2O} \times 100$$

$$\frac{10.8096}{246.492} \times 100 = 4.38\%\ \text{error}$$

308. Predict whether the following molecules are polar or nonpolar. Justify your answer using VSEPR models.

Oxygen difluoride, OF_2  polar dipoles don't cancel	Fluoromethane, CH_3F  polar dipoles do not cancel
Methane, CH_4  nonpolar no dipoles present	Hydrogen peroxide, H_2O_2  polar dipoles do not cancel
Carbon disulfide, CS_2  nonpolar dipoles cancel	Ammonia, NH_3  polar dipoles do not cancel

309. Place an "X" in the box for the kinds of attractive forces that are expected in the molecules

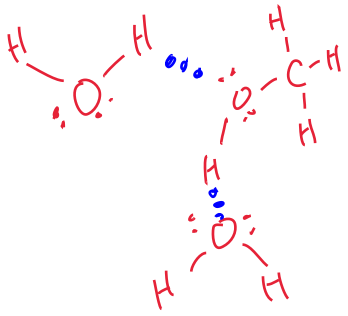
Molecule	LDF	DD	HB
Oxygen difluoride	X	X	
Methane	X		
Carbon disulfide	X		
Fluoromethane	X	X	
Hydrogen peroxide	X	X	X
Ammonia	X	X	X

310. What is required for a molecule to be polar?

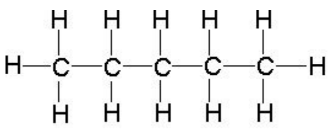
Unequal distribution of electrons as a result of dipoles not cancelling.

311. When will hydrogen bonding occur? Draw a picture of two different hydrogen bonds between water and CH_3OH.

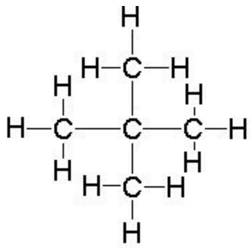
H bound to a N, O, or F



312. The boiling point of neopentane is much lower than the boiling point of pentane. Explain why.

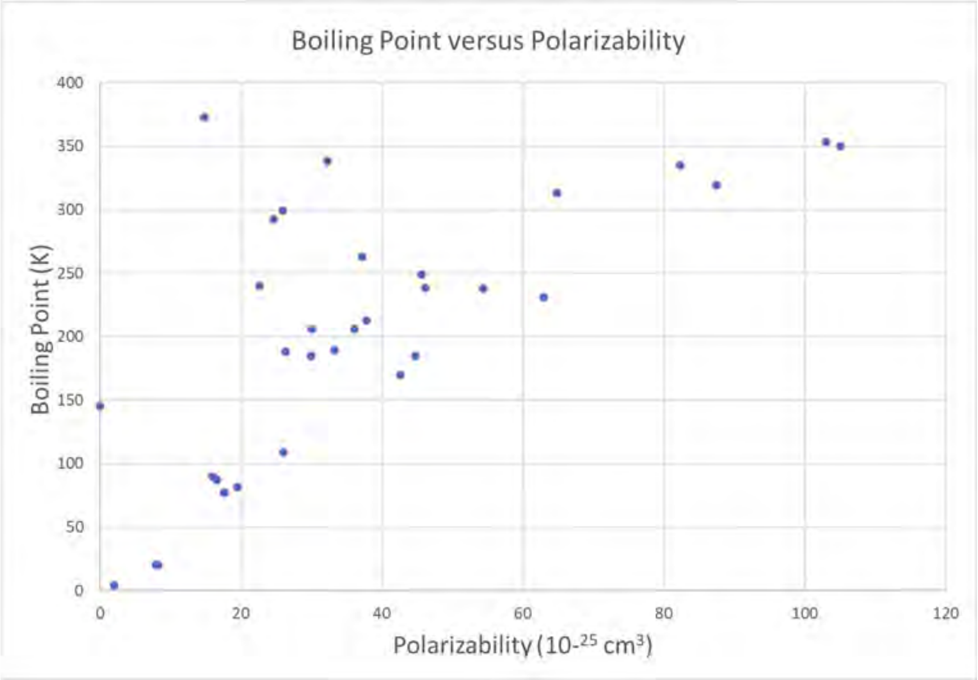


Pentane



Neopentane

Both molecules are nonpolar and only have London dispersion forces present. Pentane has a greater surface area and as a result is more polarizable. A more polarizable molecule will have greater London dispersion forces present.



313. What is polarizability?

Likelihood of being polar. Ability to be made polar.

314. Which types of compounds exhibit higher polarizability?

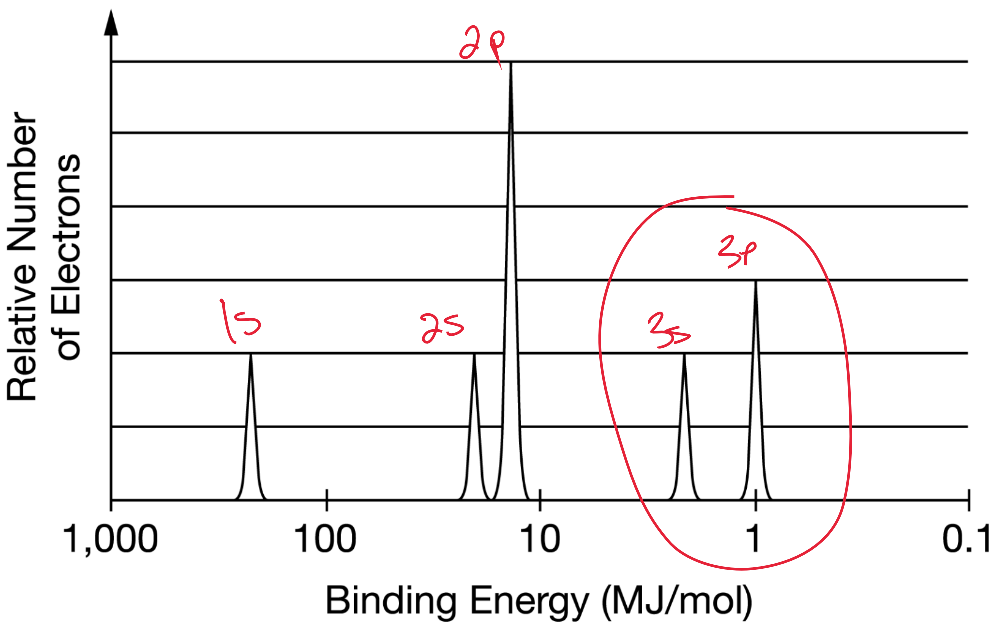
Molecules with more electrons.

315. What does the graph above tell you about the general trend for polarizability and boiling point?

As polarizability increases boiling point increases.

316. Give two reasons why the highest boiling point of about 375 K has among the lowest polarizability of about 18 x 10^{-25} cm^3 .

1. Very polar
2. Hydrogen bonding



The photoelectron spectrum of an unknown element is shown above.

317. Based on the photoelectron spectrum, identify the unknown element and write its electron configuration.

The unknown element is Phosphorus.

$1s^2 2s^2 2p^6 3s^2 3p^3$

318. Label each peak with its shell and subshell designation.

319. Circle the valence electrons on the PES graph.

320. Consider the element on the periodic table that is directly to the right of the element identified in #1. Would the 1s peak of this element appear to the left of, the right of, or in the same position as the 1s peak of the element in part 1? Explain your reasoning.

To the left. There is greater attraction for the 1s electrons because the element to the right has more protons and a greater nuclear charge.

Topic 4.6 Worksheet

321. What is a titration?

An experimental method of analysis that utilizes concentrations of solutions.

322. What is a titrant?

The solution with the known concentration.

323. What is an analyte?

The solution with an unknown concentration that is being analyzed.

324. What is the equivalence point?

When all of the titrant has reacted with all of the analyte.

325. What is the end point?

The point where the indicator changes color.

326. A 0.350 M solution of sodium hydroxide is reacted with 25.00 mL of an unknown molarity of sulfuric acid. It takes 17.3 mL of sodium hydroxide to react completely with the sulfuric acid. Determine the molarity of the sulfuric acid.

$$\underline{\ 2\ }NaOH_{(aq)} + \underline{\quad\ }H_2SO_{4(aq)} \rightarrow \underline{\quad\ }Na_2SO_{4(aq)} + \underline{\ 2\ }H_2O_{(l)}$$

.350 M
17.3 mL

25.00 mL
? M

$M = \frac{n}{L}$

$M \times V = n$

$.350 \times 17.3 = 6.055 \text{ mmols NaOH} \left| \frac{H_2SO_4}{2\,NaOH} \right. = 3.0275 \text{ mmols}$

$= \frac{3.0275 \text{ mmols}}{25 \text{ mL}}$

$M = .121\ M$

327. Determine the molarity of an unknown HCl solution if it takes 20. mL of 0.75 M sodium hydroxide to react completely with 10. mL of the HCl solution.

$$\underline{\quad\ }HCl_{(aq)} + \underline{\quad\ }NaOH_{(aq)} \rightarrow \underline{\quad\ }NaCl_{(aq)} + \underline{\quad\ }H_2O_{(aq)}$$

10 mL
? M

20 mL
.75 M

$M = \frac{n}{L}$

$M \times L = n$

$(20)(.75) = 15 \text{ mmols NaOH} \left| \frac{1\,HCl}{1\,NaOH} \right. = 15 \text{ mmols HCl}$

$= \frac{15 \text{ mmols}}{10 \text{ mL}}$

$M = 1.5\ M$

328. A colorless solution of hydrogen peroxide, H_2O_2, is titrated with an acidified, dark purple solution of potassium permanganate, $KMnO_4$, as shown by the net-ionic reaction below. The student places 5.00 mL of hydrogen peroxide in a 125-mL Erlenmeyer flask and properly fills a buret with the potassium permanganate. Note: At the end point of the titration the solution appears pale pink in color.

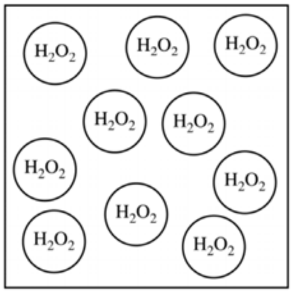
$$5\ H_2O_{2(aq)} + 2\ MnO_4{}^-{}_{(aq)} + 6\ H^+{}_{(aq)} \rightarrow 2\ Mn^{2+}{}_{(aq)} + 8\ H_2O_{(l)} + 5\ O_{2(g)}$$

a. Describe the appearance of the solution as potassium permanganate is added. Explain your reasoning.

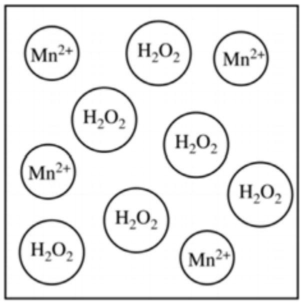
The solution turns purple as MnO_4^- is added but then clear as H_2O_2 reacts with it. Once all of the H_2O_2 is reacted the solution turns pale pink as Mn^{2+} remains.

b. What in the net-ionic reaction causes the pale pink color at the end point?

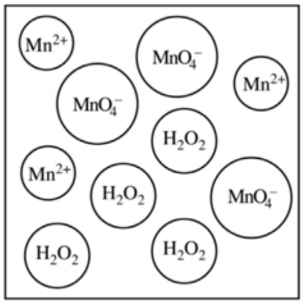
Mn^{2+}



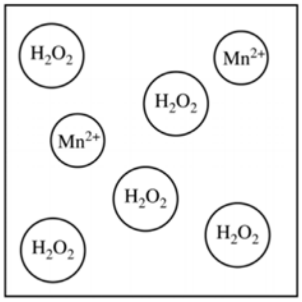
c. A particle view of a sample of $H_2O_{2(aq)}$ is shown above. The $H_2O_{2(aq)}$ is titrated with $KMnO_4$(aq), as represented in the equation above. Which of the following particle views best represents the mixture when the titration is halfway to the equivalence point? (H_2O molecules and H^+ ions are not shown.)



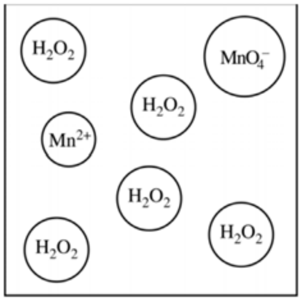
A



B



C



D

Answer C. Halfway to equivalence 1/2 of the H_2O_2 should be removed from the solution. Aswers A and B do not have 1/2 of the H_2O_2 removed. Halfway to equivalenec the MnO_4^- would be the limiting reactant and would not show up in the solution as the excess H_2O_2 completely reacts with it. Therefore D can not be the correct answer as well.

d. Determine the molarity of the hydrogen peroxide if a 5.00 mL sample requires 7.98 mL of 0.15 M $KMnO_4$ to reach equivalence.

$M \times L = n$

$(.015)(.00798) = n$

$.001197 \text{ mols} = n$

$$\frac{.001197 \text{ mols } MnO_4^- \;|\; 5 H_2O_2}{\;|\; 2 MnO_4^-} = .00299 \text{ mols } H_2O_2$$

$$M = \frac{n}{L} = \frac{.00299 \text{ mols}}{.005 \text{ L}} = .599 \text{ M}$$

assume 1.00L of solution

e. Determine the percentage by volume of hydrogen peroxide in the aqueous solution. The density of hydrogen peroxide is 1.02 g/mL. Assume the solution of hydrogen peroxide is made up of just water (d = 1.00 g/mL) and hydrogen peroxide and that the volumes are additive.

$$\frac{.599 \text{ mols } H_2O_2 \;|\; 34.016 \text{ g}}{\;|\; 1 \text{ mol}} = 20.38 \text{ g } H_2O_2$$

$$\frac{\text{vol. } H_2O_2}{\text{total vol}} \times 100 = \%$$

$$\frac{20.38 \text{ g } H_2O_2 \;|\; 1 \text{ mL}}{\;|\; 1.02 \text{ g}} = 19.98 \text{ mL } H_2O_2$$

$$\frac{19.98 \text{ mL}}{1000 \text{ mL}} \times 100 = \%$$

2.00% H_2O_2 by volume

Review Topic 1.5

329. Which has more energy, red light or blue light?

Blue

330. Give the electron configuration of the following elements:

a. Mg $1s^2 2s^2 2p^6 3s^2$

b. O $1s^2 2s^2 2p^4$

c. Br $1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^{10} 4p^5$

d. Ni $1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^8$

e. Zr $1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^{10} 4p^6 5s^2 4d^2$

331. Give the orbital notation of each of the elements from #2. You can use the abbreviated electron configuration.

a. Mg [Ne] $\underset{3s}{\underline{\uparrow\downarrow}}$

b. O [He] $\underset{2s}{\underline{\uparrow\downarrow}}\ \underbrace{\underline{\uparrow\downarrow}\ \underline{\uparrow}\ \underline{\uparrow}}_{2p}$

c. Br [Ar] $\underset{4s}{\underline{\uparrow\downarrow}}\ \underbrace{\underline{\uparrow\downarrow}\ \underline{\uparrow\downarrow}\ \underline{\uparrow\downarrow}\ \underline{\uparrow\downarrow}\ \underline{\uparrow\downarrow}}_{3d}\ \underbrace{\underline{\uparrow\downarrow}\ \underline{\uparrow\downarrow}\ \underline{\uparrow}}_{4p}$

d. Ni [Ar] $\underset{4s}{\underline{\uparrow\downarrow}}\ \underbrace{\underline{\uparrow\downarrow}\ \underline{\uparrow\downarrow}\ \underline{\uparrow\downarrow}\ \underline{\uparrow}\ \underline{\uparrow}}_{3d}$

e. Zr [Kr] $\underset{5s}{\underline{\uparrow\downarrow}}\ \underbrace{\underline{\uparrow}\ \underline{\uparrow}\ \underline{\ \ }\ \underline{\ \ }\ \underline{\ \ }}_{4d}$

332. Give the electron configuration of the most common ion for the following elements:

a. Ca^{2+} [Ne] $3s^2 3p^6$

b. Fe^{2+} [Ar] $3d^6$

c. Fe^{3+} [Ar] $3d^5$

d. O^{2-} [He] $2s^2 2p^6$

Section 4.07

Topic 4.7, 4.8, & 4.9 Worksheet

Acid and Base Reactions.

333. Define an acid.

Donates an H^+

334. Define a base.

Accepts and H^+

335. Define amphoteric.

Can both accept and donate an H^+

336. What is an acid-base conjugate pair? How do you identify an acid-base conjugate pair?

An acid and the base that it becomes. Acid has an H^+ and base is missing the H^+

337. Identify the acid, base, conjugate acid, and conjugate base in the following reactions:

$HCl\,(aq) + H_2O(l) \rightarrow H_3O^+(aq) + Cl^-(aq)$

$HCl\,(aq)$	$H_2O(l)$	$H_3O^+(aq)$	$Cl^-(aq)$
☒ Acid	☐ Acid	☐ Acid	☐ Acid
☐ Base	☒ Base	☐ Base	☐ Base
☐ Conjugate Acid	☐ Conjugate Acid	☒ Conjugate Acid	☐ Conjugate Acid
☐ Conjugate Base	☐ Conjugate Base	☐ Conjugate Base	☒ Conjugate Base

$NH_3\,(aq) + H_2O(l) \rightleftharpoons NH_4^+(aq) + OH^-(aq)$

$NH_3\,(aq)$	$H_2O(l)$	$NH_4^+(aq)$	$OH^-(aq)$
☐ Acid	☒ Acid	☐ Acid	☐ Acid
☒ Base	☐ Base	☐ Base	☐ Base
☐ Conjugate Acid	☐ Conjugate Acid	☒ Conjugate Acid	☐ Conjugate Acid
☐ Conjugate Base	☐ Conjugate Base	☐ Conjugate Base	☒ Conjugate Base

CH_3NH^- (aq)	+	HCl(*l*)	→	CH_3NH_2(aq)	+	Cl^-(aq)
☐ Acid		☒ Acid		☐ Acid		☐ Acid
☒ Base		☐ Base		☐ Base		☐ Base
☐ Conjugate Acid		☐ Conjugate Acid		☒ Conjugate Acid		☐ Conjugate Acid
☐ Conjugate Base		☐ Conjugate Base		☐ Conjugate Base		☒ Conjugate Base

338. What is a sign that the acid/base is stronger than the conjugate acid/conjugate base in a chemical reaction?

A one-way yield symbol showing that only products are formed.

→

339. If the acid/base is stronger than the conjugate acid/conjugate base, to which direction does the reaction mostly proceed?

To the right.

340. What is a sign that the acid/base is weaker than the conjugate acid/conjugate base in a chemical reaction?

A two-way yield symbol (equilibrium) showing that both products and reactants are formed.

⇌

341. If the conjugate acid/conjugate base is stronger than the acid/base, to which direction does the reaction mostly proceed?

To the left.

342. What is a neutralization reaction? What are the products of a neutralization reaction?

Acid reacting with a base. Salt and water

343. What is the net-ionic reaction for any strong acid/strong base reaction?

$$\underset{\substack{\uparrow \\ \text{strong acid}}}{H^+} + \underset{\substack{\uparrow \\ \text{strong base}}}{OH^-} \longrightarrow H_2O\,(l)$$

344. What are the signs that a chemical reaction is an acid-base reaction?

One species donates an H^+ and the other species accepts an H^+.

345. Write neutralization reactions for the following reactions. Then give the net-ionic equation. Finally, identify the acid, base, salt, and water. Note: NH_3 is a weak base and $HC_2H_3O_2$ is a weak acid. Since they are weak they do not break up completely and some molecules are left in solution.

a. $HCl + NaOH \rightarrow NaCl + H_2O$
acid base salt water

$H^+ + OH^- \rightarrow H_2O(l)$

b. $H_2SO_4 + 2NaOH \rightarrow Na_2SO_4 + 2H_2O$
acid base salt water

$H^+ + OH^- \rightarrow H_2O(l)$

c. $NH_3 + HCl \rightarrow NH_4Cl(aq)$
base acid salt

$NH_3(aq) + H^+ \rightarrow NH_4^+(aq)$

d. $HC_2H_3O_2 + NaOH \rightarrow NaC_2H_3O_2 + H_2O(l)$
acid base salt water

$HC_2H_3O_2(aq) + OH^-(aq) \rightarrow C_2H_3O_2^-(aq) + H_2O(l)$

Reduction and Oxidation (Redox) Reactions.

346. Define oxidation.

Loss of electrons

347. Define reduction.

Gain of electrons

348. Define oxidizing agent. What types of elements tend to be good oxidizing agents?

Causes oxidation. The species being reduced. Nonmetals.

Note: This term is no longer on the exam but it can still be helpful when discussing what is being oxidized and what is being reduced.

349. Define reducing agent. What types of elements tend to be good reducing agents?

Causes reduction. The species being oxidized. Metals.

Note: This term is no longer on the exam but it can still be helpful when discussing what is being oxidized and what is being reduced.

350. How can you identify that a substance has been oxidized?

The oxidation number becomes more positive.

351. How can you identify that a substance has been reduced?

The oxidation number becomes more negative.

352. How can you identify a reduction-oxidation reaction?

Change in oxidation number. If an element becomes bonded in a compound.

353. Where are the electrons located in a balanced oxidation half-reaction? A reduction half-reaction?

Oxidation = products
Reduction = reactants

354. What are the rules for applying oxidation numbers?

1. Elements have an oxidation number of zero.
2. The sum of the oxidation numbers in a neutral substance is zero.
3. The sum of the oxidation numbers in an ion is equal to the ion charge.
4. Common oxidation numbers: Group 1 = +1, Group 2 = +2, oxygen sometimes is -2, hydrogen is almost always +1 except when bonded to metals.
5. In binary compounds Group 17 = -1, Group 16 = -2, Group 15 = -3

355. Determine the oxidation number of the atom listed below.

a. C in CO_2

CO_2

+4 -2

b. S in S^{2-}

S^{2-}

-2

c. S in $S_{(s)}$

0

d. S in $SO_4{}^{2-}$

$SO_4{}^{2-}$

+6 -2

e. S in $SO_3{}^{2-}$

$SO_3{}^{2-}$

+4 -2

f. Mn in $MnO_4{}^{-}$

$MnO_4{}^{-}$

+7 -2

356. For the following reactions,

1. Balance the equation.
2. Write a net-ionic equation.
3. Determine the oxidation states of all species in the net-ionic equation.
4. Determine if the reaction is a redox reaction.
   - (A) If the reaction is not a redox reaction then write **NONE**.
   - (B) If the reaction is a redox reaction determine what is being oxidized and what is being reduced

a. ______$Zn_{(s)}$ + __2__$HCl_{(aq)}$ → ______$ZnCl_{2(aq)}$ + ______$H_{2(g)}$

$Zn_{(s)} + 2H^{+}_{(aq)} \longrightarrow Zn^{2+}_{(aq)} + H_{2(g)}$ Redox Rxn

0 +1 +2 0

Zn is oxidized

H^+ is reduced

b. ______$Co(NO_3)_{2(aq)}$ + ______$Na_2HPO_{4(aq)}$ → ______$CoHPO_{4(s)}$ + __2__$NaNO_{3(aq)}$

$Co^{2+}_{(aq)} + HPO_4^{-2}{}_{(aq)} \longrightarrow CoHPO_{4(s)}$ None

+2 +1 +5 -2 +2 +1 +5 -2

c. ______$CH_{4(g)}$ + __2__$O_{2(g)}$ → __2__$H_2O_{(l)}$ + ______$CO_{2(g)}$

$CH_{4(g)} + 2O_{2(g)} \longrightarrow 2H_2O_{(l)} + CO_{2(g)}$ Redox

-4 +1 0 +1 -2 +4 -2

C is oxidized (CH_4)

O_2 is reduced

d. ______$H_2SO_{4(aq)}$ + ______$Ca(OH)_{2(aq)}$ → ______$CaSO_{4(s)}$ + __2__$H_2O_{(l)}$

$2H^{+}_{(aq)} + SO_4^{2-}{}_{(aq)} + Ca^{2+}_{(aq)} + 2OH^{-}_{(aq)} \longrightarrow CaSO_{4(s)} + 2H_2O_{(l)}$

+1 +6 -2 +2 -2 +1 +2 +6 -2 +1 -2

None

Note: These answers will vary from student to student.

357. Write spontaneous redox reactions and determine the overall voltage for the following reactions. DO NOT REPEAT PAIRS OF REACTIONS.

1	2	3	4	5	6	7	8	9	10	11	12	13	14	15	16	17	18	19	20	21	22	23	24	25	26
A	B	C	D	E	F	G	H	I	J	K	L	M	N	O	P	Q	R	S	T	U	V	W	X	Y	Z

a. For reaction 1 use the number that corresponds to your first name initial and for reaction 2 use the number that corresponds to your last name initial. If the second reaction is the same letter as your first reaction then add one to the number.

J

D

$Au^{3+} + 3e^- \rightarrow Au \quad +1.40V$ ← Flip

$3\times \quad 3PbO_2 + \overset{12}{\cancel{4}}H^+ + 3SO_4^{2-} + \overset{6}{\cancel{2}}e^- \rightarrow 3PbSO_4 + \overset{6}{\cancel{2}}H_2O \quad +1.69V$

$2\times \quad 2Au \rightarrow 2Au^{3+} + \overset{6}{\cancel{3}}e^- \quad -1.40V$

$3PbO_2 + 12H^+ + 3SO_4^{2-} + 2Au \rightarrow 3PbSO_4 + 6H_2O + 2Au^{3+} \quad +0.29V$

b. For reaction 1 use the second letter of your first name and for reaction 2 use the second letter of your last name. If the second reaction is the same letter as your first reaction then add one to the number.

$Au^+ + e^- \rightarrow Au \quad +1.69V$ Flip

$F_2 + 2e^- \rightarrow 2F^- \quad +2.87V$

$2Au \rightarrow 2Au^{2+} + 2e^- \quad -1.69V$

$2Au + F_2 \rightarrow 2Au^{2+} + 2F^- \quad 1.18V$

c. For reaction 1 use the first letter of your friends first name and for reaction 2 use the second letter of your friends first name. If the second reaction is the same letter as your first reaction then add one to the number.

$Ag^+ + e^- \longrightarrow Ag$ $+0.80V$ (flip)

$Ce^{4+} + e^- \longrightarrow Ce^{3+}$ $+1.61V$

$Ag \longrightarrow Ag^+ + e^-$ $-0.80V$

$Ce^{4+} + Ag \longrightarrow Ag^+ + Ce^{3+}$ $+0.81V$

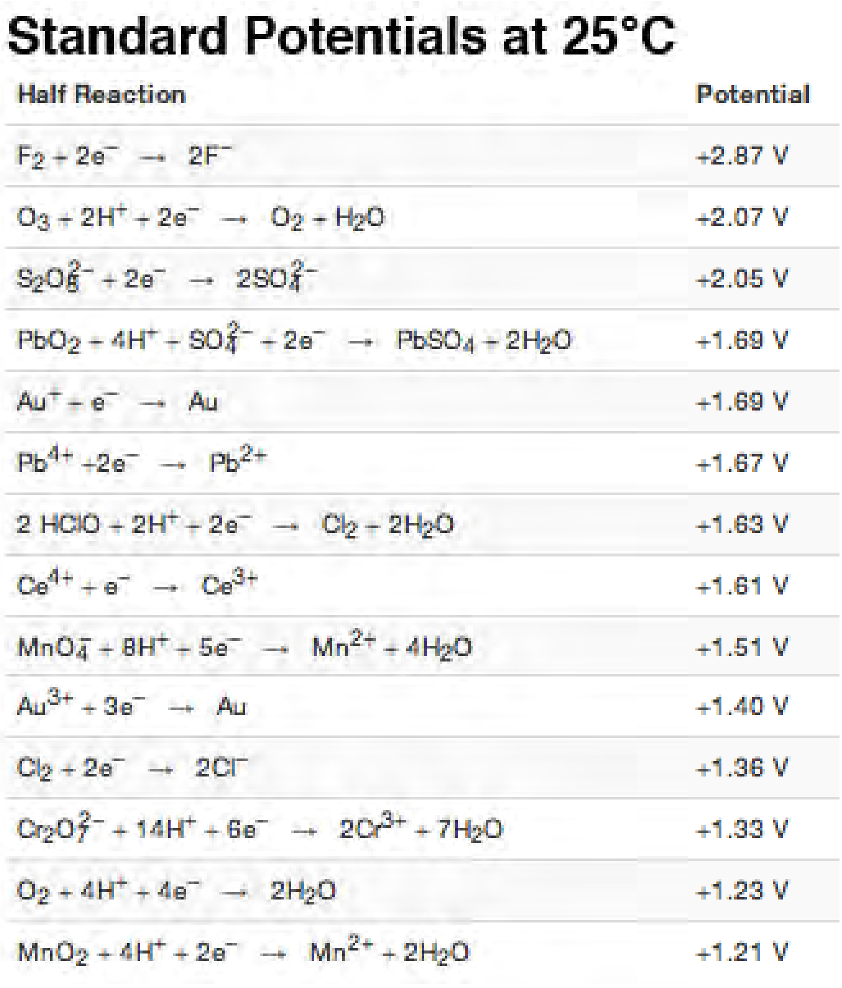
d. For reaction 1 use the fifth letter (or the last letter if there are not five) of your favorite food and for reaction 2 usse the third letter of your middle name. If you don't have a middle name then use the third letter of your favorite tv show. If the second reaction is the same letter as your first reaction then add one to the number.

$Au^+ + e^- \longrightarrow Au$ $+1.69V$ (flip)

$S_2O_8^{2-} + 2e^- \longrightarrow 2SO_4^{2-}$ $+2.05V$

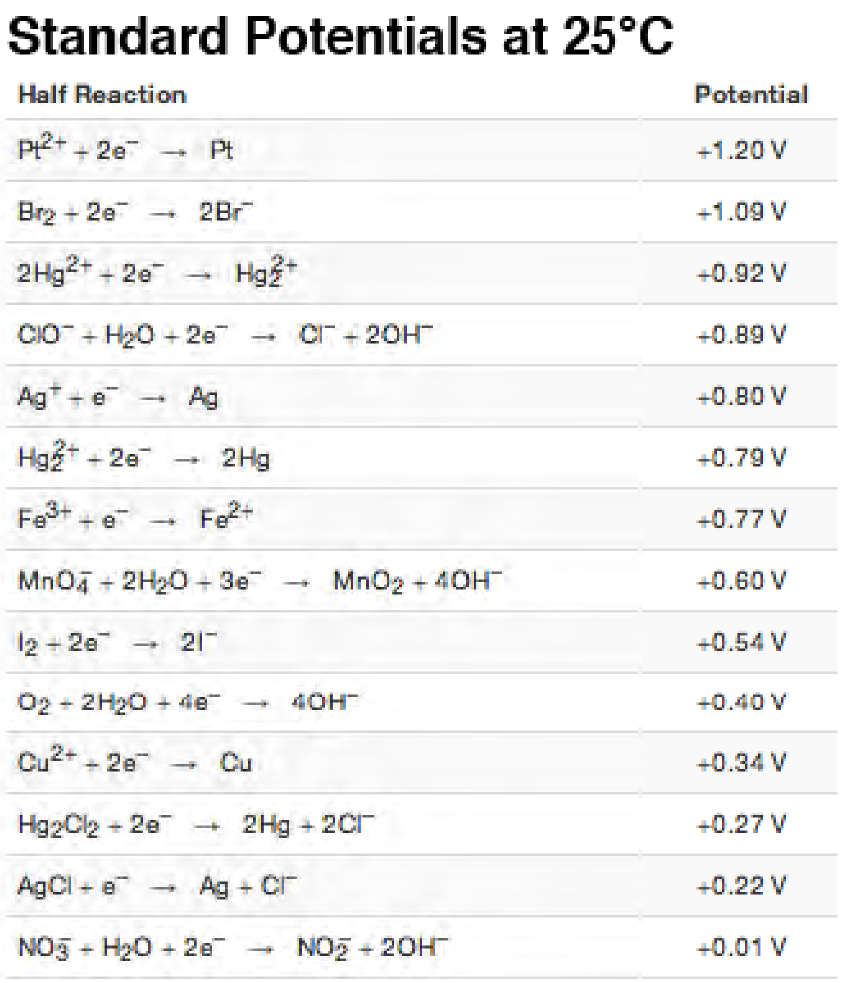
$2Au \longrightarrow 2Au^+ + 2e^-$ $-1.69V$

$2Au + S_2O_8^{2-} \longrightarrow 2Au^+ + 2SO_4^{2-}$ $0.36V$



Standard Potentials at 25°C

Half Reaction	Potential
$F_2 + 2e^- \rightarrow 2F^-$	+2.87 V
$O_3 + 2H^+ + 2e^- \rightarrow O_2 + H_2O$	+2.07 V
$S_2O_8^{2-} + 2e^- \rightarrow 2SO_4^{2-}$	+2.05 V
$PbO_2 + 4H^+ + SO_4^{2-} + 2e^- \rightarrow PbSO_4 + 2H_2O$	+1.69 V
$Au^+ + e^- \rightarrow Au$	+1.69 V
$Pb^{4+} + 2e^- \rightarrow Pb^{2+}$	+1.67 V
$2\ HClO + 2H^+ + 2e^- \rightarrow Cl_2 + 2H_2O$	+1.63 V
$Ce^{4+} + e^- \rightarrow Ce^{3+}$	+1.61 V
$MnO_4^- + 8H^+ + 5e^- \rightarrow Mn^{2+} + 4H_2O$	+1.51 V
$Au^{3+} + 3e^- \rightarrow Au$	+1.40 V
$Cl_2 + 2e^- \rightarrow 2Cl^-$	+1.36 V
$Cr_2O_7^{2-} + 14H^+ + 6e^- \rightarrow 2Cr^{3+} + 7H_2O$	+1.33 V
$O_2 + 4H^+ + 4e^- \rightarrow 2H_2O$	+1.23 V
$MnO_2 + 4H^+ + 2e^- \rightarrow Mn^{2+} + 2H_2O$	+1.21 V



Standard Potentials at 25°C

Half Reaction	Potential
$Pt^{2+} + 2e^- \rightarrow Pt$	+1.20 V
$Br_2 + 2e^- \rightarrow 2Br^-$	+1.09 V
$2Hg^{2+} + 2e^- \rightarrow Hg_2^{2+}$	+0.92 V
$ClO^- + H_2O + 2e^- \rightarrow Cl^- + 2OH^-$	+0.89 V
$Ag^+ + e^- \rightarrow Ag$	+0.80 V
$Hg_2^{2+} + 2e^- \rightarrow 2Hg$	+0.79 V
$Fe^{3+} + e^- \rightarrow Fe^{2+}$	+0.77 V
$MnO_4^- + 2H_2O + 3e^- \rightarrow MnO_2 + 4OH^-$	+0.60 V
$I_2 + 2e^- \rightarrow 2I^-$	+0.54 V
$O_2 + 2H_2O + 4e^- \rightarrow 4OH^-$	+0.40 V
$Cu^{2+} + 2e^- \rightarrow Cu$	+0.34 V
$Hg_2Cl_2 + 2e^- \rightarrow 2Hg + 2Cl^-$	+0.27 V
$AgCl + e^- \rightarrow Ag + Cl^-$	+0.22 V
$NO_3^- + H_2O + 2e^- \rightarrow NO_2^- + 2OH^-$	+0.01 V

Precipitation Reactions

358. For the reactions below:

(A) Complete the reaction.

(B) Determine if a precipitate or gas forms. Include phase symbols.

(C) If a precipitate or gas forms write a net-ionic reaction.

a. $NaCl_{(aq)} + NH_4NO_{3(aq)} \rightarrow$ $NaNO_3\,(aq) + NH_4Cl\,(aq)$

no ppt. or gas

ppt.

b. $2\,Na_3PO_{4(aq)} + 3\,Ba(NO_3)_{2(aq)} \rightarrow 6\,NaNO_3\,(aq) + Ba_3(PO_4)_2\,(s)$

$2\,PO_4^{3-} + 3Ba^{2+} \rightarrow Ba_3(PO_4)_2\,(s)$

c. $K_2S_{(aq)} + (NH_4)_2CO_{3(aq)} \rightarrow$ $K_2CO_3\,(aq) + (NH_4)_2S\,(aq)$

no ppt.

d. $HCl_{(aq)} + NaHCO_{3(aq)} \rightarrow$ $NaCl\,(aq) + H_2O\,(l) + CO_2\,(g)$

$H^+ + HCO_3^- \rightarrow H_2O\,(l) + CO_2\,(g)$

Back titration. (https://cnx.org/contents/r1hm-4a2@3/Back-titration)
A back titration is sometimes used to determine the amount of a substance if the reactant is volatile, if the reactant is contaminated, if the reactant is an insoluble salt, the reaction is too slow, or the reaction involves weak acids and bases. In a back titration, reactant A of unknown concentration is reacted with excess reactant B of known concentration. Then, a normal titration is performed to determine the amount of reactant B in excess. Once you know the excess amount of reactant B you can determine the amount of reactant B used and thus the amount of reactant A.

359. A student was asked to determine the mass, in grams, of calcium carbonate present in a 0.125 g sample of chalk. The student placed the chalk sample in a 250 mL Erlenmeyer flask and added 50.00 mL of 0.200 mol L $^{-1}$ HCl using a pipette. The excess HCl was then titrated with 0.250 mol L $^{-1}$ NaOH. The average amount of NaOH required to reach equivalence was 32.12 mL. Calculate the mass of calcium carbonate, in grams, present in the chalk sample.

2nd Rxn w/ excess HCl

$HCl + NaOH \rightarrow NaCl + H_2O$

.25M

32.12mL

$M \times V = mols$

$(.25)(32.12) = mols$

.00803 mols NaOH (1:1 mol ratio)

.00803 mols HCl

↑ This is the amount of HCl left over after reacting w/ the chalk

1st Rxn in excess

$2HCl + CaCO_3 \rightarrow CaCl_2 + H_2O + CO_2$

50.00mL

.2 M

$M \times V = n$

$(.2)(.05) = n$

.01 mols HCl in excess

− .00803 mols from rxn 2

.00197 mols HCl used in Rxn 1

$$\frac{.00197 \text{ mols HCl}}{} \Big| \frac{1\ CaCO_3}{2\ HCl} = 9.85 \times 10^{-4} \text{ mols } CaCO_3$$

$$\frac{9.85 \times 10^{-4} \text{ mols } CaCO_3}{} \Big| \frac{100.09\ g}{1\ mol} = .0986\ g\ CaCO_3$$

% of $CaCO_3$ in chalk

$$\frac{.0986g}{.125g} \times 100 = 78.9\% \text{ Chalk}$$

(This wasn't asked but it's a good review of mixtures)

Gravimetric analysis.

360. A bottle of magnesium chloride, $MgCl_2$, has been contaminated with an unknown amount of sodium nitrate, $NaNO_3$. In order to determine the percent by mass of magnesium chloride in the mixture a student conducts a gravimetric filtration. The student takes a known mass of the mixture and dissolves it in 50.0 mL of distilled water to make an aqueous solution. That solution is then reacted with an excess of silver nitrate, $AgNO_3$, to precipitate the chloride as silver chloride, $AgCl_{(s)}$. The precipitate is filtered, dried and weighed. The data from the experiment is shown below.

Mass of $MgCl_2$ & $NaNO_3$ mixture	0.7209 g
Mass of filter paper	4.450 g
Mass of filter paper + precipitate after drying	5.482 g

a. Determine the mass of precipitate formed in the reaction.

$$\begin{array}{r} 5.482g \\ -\ 4.450g \\ \hline 1.032g \text{ ppt.} \end{array}$$

b. Determine the mass of magnesium chloride in the original sample assuming the silver nitrate was in excess and the reaction below went to completion.

$$\underline{\qquad}MgCl_{2(aq)} + \underline{\ 2\ }AgNO_{3(aq)} \rightarrow \underline{\ 2\ }AgCl_{(s)} + \underline{\qquad}Mg(NO_3)_{2(aq)}$$

$\underline{\ ?\ }g$

$$\frac{1.032g\ AgCl}{} \left| \frac{1mol}{143.35g} \right. = \frac{.007199 \text{ mols } AgCl}{} \left| \frac{1MgCl_2}{2AgCl} \right. = .003599 \text{ mols } MgCl_2$$

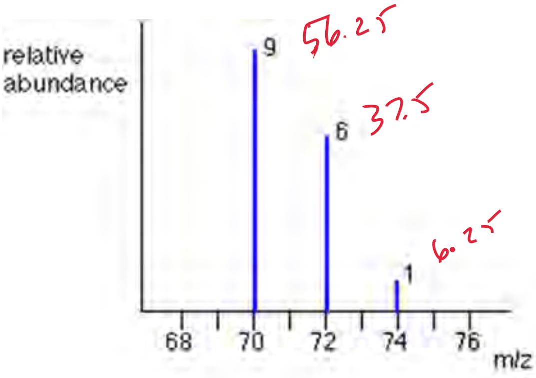
$$\frac{.003599 \text{ mols } MgCl_2}{} \left| \frac{95.21g}{1mol} \right. = 0.343g\ MgCl_2$$

c. Determine the percent by mass of magnesium chloride in the original sample.

$$\frac{.343g\ MgCl_2}{.7209\ g \text{ mixture}} \times 100 = 47.5\%\ MgCl_2$$

Review Topic 1.2

361. What is the relative atomic mass of this monatomic element?



56.25

37.5

6.25

$$\frac{(70 \times 9) + (72 \times 6) + (74 \times 1)}{16} = 71 \text{ amu}$$

Unit 5

Topic 5.5 Worksheet

362. What are the conditions for the collision theory?

1. Particles must collide.
2. Particles must have correct orientation when they collide.
3. Particles must have enough energy to overcome activation energy when they collide.

363. Define activation energy.

The minimum amount of energy required to start a reaction.

364. Explain the various ways a reaction could be unsuccessful according to the collision model.

1. Particles do not collide.
2. Specific atoms do not collide.
3. Not enough energy.

365. Explain how increasing temperature can increase the rate of a reaction in terms of…

a. the collisions.

A higher temperature means more energy for the particles. The particles move around faster which means more possible collisions.

b. the activation energy.

A higher temperature means more energy for the particles. Therefore more particles will have higher energy which means that more particles will have an amount of energy greater than the activation energy.

366. Explain why a termolecular elementary step is rare in terms of the collision model.

Three molecules colliding in the correct orientation is rare.

367. Explain why an increase in concentration of a reactant may lead to an increase in reaction rate according to the collision model.

As concentration increases more molecules are available for collisions. More collisions means a greater chance of a reaction which leads to a higher reaction rate.

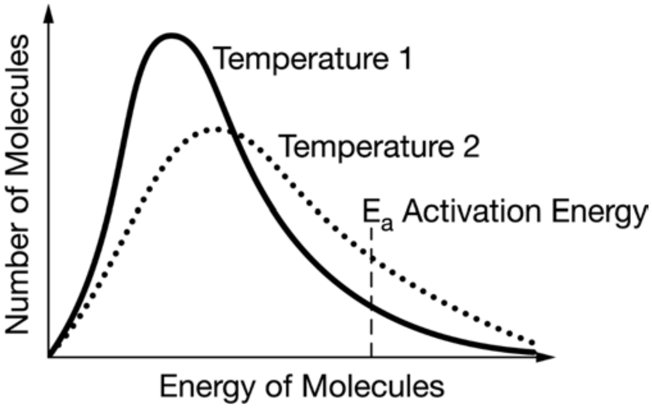
368. Explain why, at constant temperature, an increase in pressure of a gas causes an increase in the reaction rate in terms of the collision model.

An increase in pressure at constant temperature means more gas particles present which would cause an increase in the number of collisions. (Assumes an increase in the moles of gas which causes an increase in pressure.)

OR

An increase in pressure means the particles are closer to each other than at low pressures. Particles closer to each other will lead to more collisions and more collisions leads to an increase in reaction rate. (Assumes a decrease in container volume which causes an increase in pressure.)

369. Use the figure below to answer the questions that follow.



a. Which is at a higher temperature, T_1 or T_2? Justify your answer by referring to the figure.

T_2 is at a higher temperature. More particles have higher energy because the curve is shifted to the right.

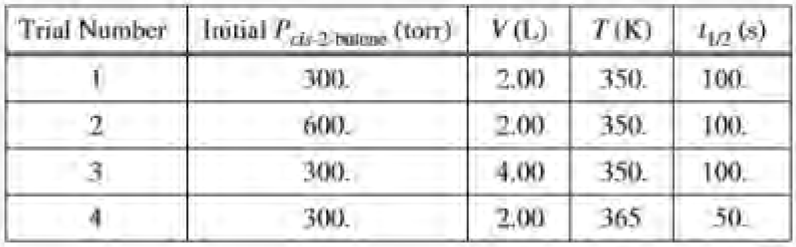
b. Which would have a greater rate of disappearance of reactants, T_1 or T_2. Explain your answer by referring to the collision model.

T_2 would have a greater rate of disappearance of reactants. A higher temperature means greater energy which means more collisions. Also, at a higher energy more particles have enough energy to overcome the activation energy required for the reaction.

370. The two elementary steps below are carried out at the same temperature. Reaction 1 takes longer than reaction 2. Assuming no difference in correctly oriented collision, explain why reaction 1 would take longer than reaction 2.

Reaction 1	$H_2(g) + ICl(g) \rightarrow HI(g) + HCl(g)$	Slow
Reaction 2	$HI(g) + ICl(g) \rightarrow HCl(g) + I_2(g)$	Fast

Reaction 1 must have a larger activation energy. A larger activation energy will be a slower reaction since fewer particles have enough energy to overcome the activation energy.



Trial Number	Initial $P_{cis\text{-}2\text{-}butene}$ (torr)	V (L)	T (K)	$t_{1/2}$ (s)
1	300.	2.00	350.	100.
2	600.	2.00	350.	100.
3	300.	4.00	350.	100.
4	300.	2.00	365	50.

371. Refer to the data table above. The half-life of a reaction was measured under various conditions. The half-life of the reaction in trial 4 is less than the half-life in trial 1. Explain why in terms of activation energy.

Trial 4 is at a higher temperature. A higher temperature means more particles have the required energy to overcome the activation energy.

Topic 5.4 Worksheet

372. Determine the molecularity of the reaction.

	Elementary Step	Molecularity		
(A)	A + B → X + Y	Unimolecular	[✓ Bimolecular]	Termolecular
(B)	2A + B → A_2B	Unimolecular	Bimolecular	[✓ Termolecular]
(C)	NO + O_3 → NO_2 + O_2	Unimolecular	[✓ Bimolecular]	Termolecular
(D)	Cl + CH_4 → HCl + CH_3	Unimolecular	[✓ Bimolecular]	Termolecular
(E)	Ar + O_3 → Ar + O_3^*	Unimolecular	[✓ Bimolecular]	Termolecular
(F)	A + A → B + C	Unimolecular	[✓ Bimolecular]	Termolecular
(G)	O_3 → O_2 + O	[✓ Unimolecular]	Bimolecular	Termolecular
(H)	O + O_2 + N_2 → O_3 + N_2	Unimolecular	Bimolecular	[✓ Termolecular]
(I)	A → B + C + D	[✓ Unimolecular]	Bimolecular	Termolecular

373. Describe the *specific* collision that must occur for the elementary steps from 372.

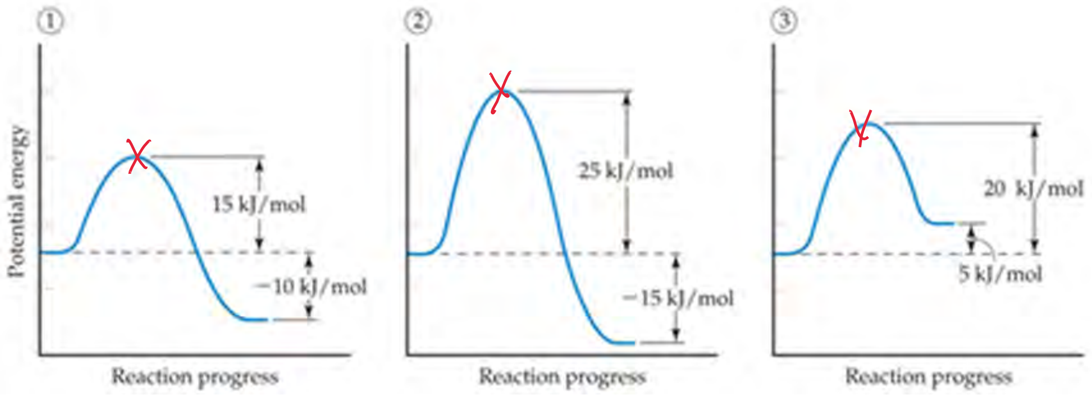
a. Reaction C

The O from NO must collide with an O from O_3.

b. Reaction D

The H from CH_4 must collide with the Cl from HCl.

374. Consider a series of reactions having these reaction coordinate energy profiles



a. Rank the reactions from slowest to fastest. Explain your reasoning.

2 < 3 < 1 (from slowest to fastest). These are ranked from highest activation energy to smallest activation energy. A higher activation energy will take longer because fewer particles will have enough energy to overcome the activation energy.

b. Rank the reactions from slowest to fastest in the reverse direction. Explain your reasoning.

2 < 1 < 3 (from slowest to fastest). Same reasoning as #374a.

c. Give the activation energy of each reaction in the forward direction.

Rxn 1: 15 kJ/mol

Rxn 2: 25 kJ/mol

Rxn 3: 20 kJ/mol

d. Give the activation energy of each reaction in the reverse direction.

Rxn 1: 25 kJ/mol

Rxn 2: 40 kJ/mol

Rxn 3: 15 kJ/mol

e. Is reaction 1 endothermic or exothermic? Justify your answer. What about the reverse direction?

Exothermic. The products end with less energy than the reactants therefore energy must be released. The reverse reaction is endothermic.

f. Is reaction 3 endothermic or exothermic? Justify your answer. What about the reverse direction?

Endothermic. The products end with more energy than the reactants therefore energy must be absorbed. The reverse reaction is exothermic.

g. Give the overall energy change for each reaction in the forward direction.

Rxn 1: -10 kJ/mol

Rxn 2: -15 kJ/mol

Rxn 3: +5 kJ/mol

h. Give the overall energy change for each reaction in the reverse direction.

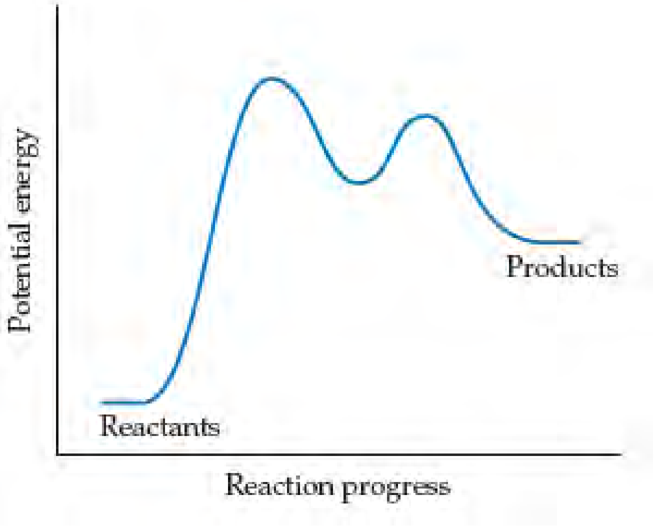
Rxn 1: +10 kJ/mol

Rxn 2: +15 kJ/mol

Rxn 3: -5 kJ/mol

i. Place an “x” on each diagram where the transition state occurs.

375. Use the reaction coordinate energy profile given below to answer the questions that follow.



a. How many elementary steps are involved in the reaction? Explain your reasoning.

There are 2 elementary steps. Since there are 2 energy "hills" (activation energy) there are 2 elementary steps.

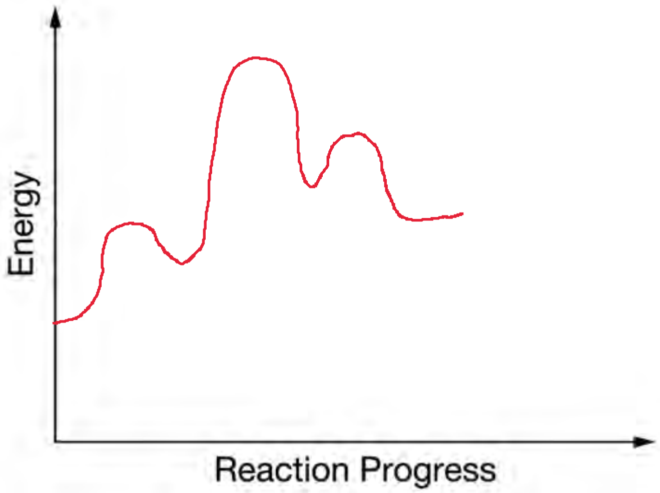
b. Which step would be faster? Explain your reasoning.

Step 2. Step 2 has a lower activation energy than step 1. A lower activation energy means that more particles would have the required energy to overcome the activation energy.

376. The following sequence of elementary steps is a proposed mechanism for a given reaction.

Step 1:	$NO + NO \rightleftharpoons N_2O_2$	Fast
Step 2:	$N_2O_2 + H_2 \rightleftharpoons H_2O + N_2O$	Slow
Step 3:	$N_2O + H_2 \rightleftharpoons N_2 + H_2O$	Fast

a. On the incomplete reaction energy diagram below, draw a curve that shows the following two details:
   i. The relative activation energies of the three elementary steps
   ii. The enthalpy change of the overall reaction if the reaction is endothermic



Review Topic 3.13

377. Explain the process of creating a blank and calibrating a spectrophotometer with acetone as the solvent in the experiment.

Fill the cuvette with acetone. Wipe cuvette with kimwipe. Place cuvette in spectrophotometer and zero out the spectrophotometer.

378. Explain why a blank is needed to calibrate the spectrophotometer.

It "zeros out" the absorbance of the solvent.

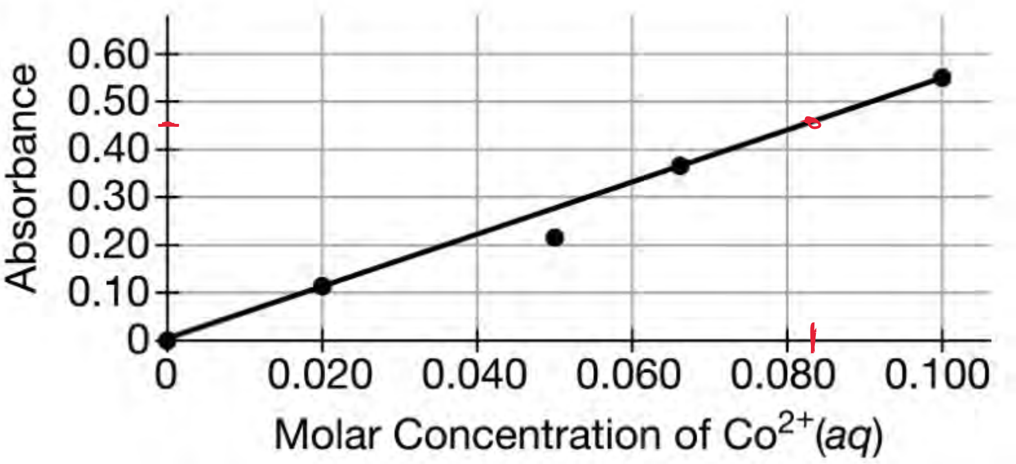
379. Explain the effect of a fingerprint on the cuvette for the recorded absorbance.

Absorbance would be greater as light is scattered because of the fingerprint. If light is scattered then it is not absorbed by the sensor.

380. Explain why the concentration of a colorless solution of sodium chloride can not be measured via Beer’s Law.

A colorless solution will not absorb light at any molarity. Beer's Law assumes light is being absorbed by a colored solution. The more concentrated the colored solution the darker the color. The darker the color the more light will be absorbed.

381. Use the graph below to answer the questions that follow.



a. Determine the molarity of a solution that has an absorbance of 0.45.

About 0.085 M

b. What error could have possibly caused the point the student plotted at 0.050 M Co^{2+}*(aq)*?

Wrong molarity recorded.
Water left in the cuvette.

c. How many moles of Co^{2+}(aq) would be present in 100. mL of a solution with an absorbance of 0.15 ?

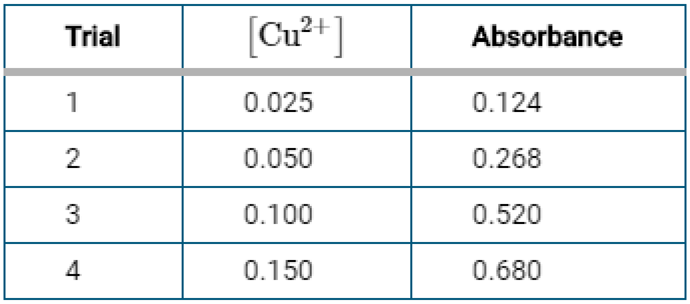
$M = .028$ $M = \frac{n}{L}$ $.028 \times .1 = n$

$M \times L = n$ $.0028 \text{ moles} = n$

382. A student conducts an experiment to determine the absorbance of red food coloring at various concentrations. The student wants to do the same experiment with blue food coloring. How would the student need to modify the original experimental procedure to determine the absorbance of blue food coloring at various concentrations? Why is the modification required?

The wavelength of the spectrophotometer would need changed. Different colors of solutions absorb different wavelengths of light. The spectrophotometer's wavelength should be set at a color opposite the color of the solution.

383. A student prepares a calibration curve for [Cu^{2+}] by making a serial dilution and recording the absorbance for each concentration as shown in the data table below. Use the data table below to answer the questions that follow.



Trial	$[Cu^{2+}]$	Absorbance
1	0.025	0.124
2	0.050	0.268
3	0.100	0.520
4	0.150	0.680

a. Which trial has discrepant data compared to the others? Justify your answer by referring to the data.

Trial 4. The absorbance does not increase enough.

b. Would the actual concentration of the trial identified in #383a be greater than, less than, or equal to the concentration expected in the trial?

Less than. The absorbance isn't high enough for the molarity that is recorded. For the absorbance the molarity should be lower than recorded.

Section 5.02

Topic 5.1 Worksheet

384. Consider the combustion of ethylene, $C_2H_4(g) + 3\ O_2(g) \rightarrow 2\ CO_2(g) + 2\ H_2O(g)$. If the concentration of C_2H_4 is decreasing at the rate of 0.036 M/s, what are the rates of change in the concentrations of CO_2 and H_2O?

$\frac{1}{1} Rate_{C_2H_4} = \frac{1}{2} Rate_{CO_2}$

$1(.036) = \frac{1}{2} Rate_{CO_2}$

$.072\ M/s = Rate_{CO_2}$

$\frac{1}{1} Rate_{C_2H_4} = \frac{1}{2} Rate_{H_2O}$

$1(.036) = \frac{1}{2} Rate_{H_2O}$

$.072\ M/s = Rate_{H_2O}$

385. The rate of decrease in N_2H_4 partial pressure in a closed reaction vessel from the reaction $N_2H_4(g) + H_2(g) \rightarrow 2\ NH_3(g)$ is 74 torr per hour. What are the rates of change of NH_3 partial pressure and total pressure in the vessel?

$\frac{1}{1} Rate_{N_2H_4} = \frac{1}{2} Rate_{NH_3}$

$74\ torr/hr = \frac{1}{2} Rate_{NH_3}$

$148\ torr/hr = Rate_{NH_3}$

There is no change in total pressure because 2 moles of gas reactants are converted to 2 moles of gas products.

$$5\ H_2O_2(aq) + 2\ MnO_4^-(aq) + 6\ H^+(aq) \rightarrow 2\ Mn^{2+}(aq) + 8\ H_2O(l) + 5\ O_2(g)$$

386. At a certain time during the titration, the rate of appearance of $O_2(g)$ was 2.5×10^{-3} mol/(L•s).

a. What was the rate of disappearance of MnO_4^- at the same time?

$\frac{1}{5} Rate_{O_2} = \frac{1}{2} Rate_{MnO_4^-}$

$\frac{1}{5}(2.5\times10^{-3}) = \frac{1}{2} Rate_{MnO_4^-}$

$.0010\ \frac{M}{s} = Rate_{MnO_4^-}$

b. What was the rate of appearance of Mn^{2+} at the same time?

$\frac{1}{5} Rate_{O_2} = \frac{1}{2} Rate_{Mn^{2+}}$

$Rate_{Mn^{2+}} = .0010\ \frac{M}{s}$

c. As time progresses, would you expect the rate of disappearance of MnO_4^- to increase, decrease, or remain the same? Explain your answer.

I would expect the rate of disappearance of MnO_4^- to decrease. As more MnO_4^- is used there are fewer MnO_4^- to react, causing the rate to decrease.

387. A kinetics experiment is set up to collect the gas that is generated when a sample of solid $CaCO_3$ is added to a solution of ethanoic acid, CH_3COOH. The rate of reaction between $CaCO_3$ and CH_3COOH is determined by measuring the volume of gas generated at 25 °C and 1 atm as a function of time. Describe how each of the following would effect the rate of the reaction. Explain your answer in terms of the collision theory.

a. Decreasing the volume of ethanoic acid used in the experiment.

Reaction rate would decrease. Fewer collisions.

b. Decreasing the molarity of the ethanoic acid used in the experiment.

Reaction rate would decrease. Fewer collisions.

c. Increasing the temperature at which the experiment is performed.

Reaction rate would increase. More collisions and more particles have the required energy to overcome the activation energy.

d. Decreasing the temperature at which the experiment is performed.

Reaction rate would decrease. Fewer collisions and fewer particles have the required energy to overcome the activation energy.

e. Decreasing the particle size of the $CaCO_3$ by grinding it into a fine powder.

Reaction rate would increase. More surface area means more opportunities for collisions.

388. Explain how catalysts …

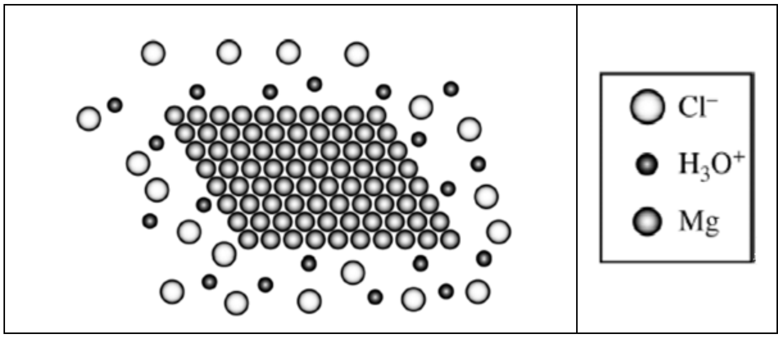
a. lower the required activation energy.

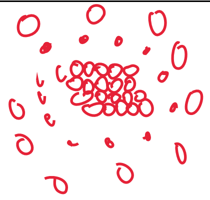
Catalysts provide an alternative reaction pathway/mechanism for the reaction to happen. This different reaction mechanism has a lower activation energy than the mechanism without the catalyst.

b. correct for orientation.

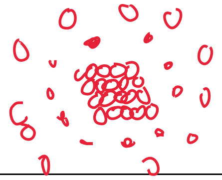
Catalysts can hold the reactant in an optimal position for collisions, i.e. adsorption.

389. In the box below on the left is a picture of a solid sample of Mg and the strong acid HCl*(aq)* reacting. In the box below on the right draw a picture of Mg(s) with a smaller particle size reacting with HCl*(aq)*. Use your picture to explain why the Mg with smaller particle size reacts at a faster rate.





More area for H_3O^+ to react



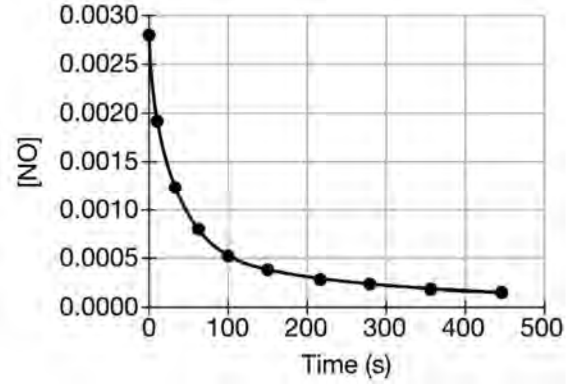
390. In a lab to determine the percent of copper in brass, long ribbons of brass were reacted with nitric acid. The teacher explains that you could have also used small shavings instead of the long ribbons. However, the shavings don't work as well because the nitric acid "bubbles" over. Explain why using the small shavings bubbles over.

The small shavings have a greater surface area which means a greater reaction rate. That increased reaction rate makes it "bubble" over.

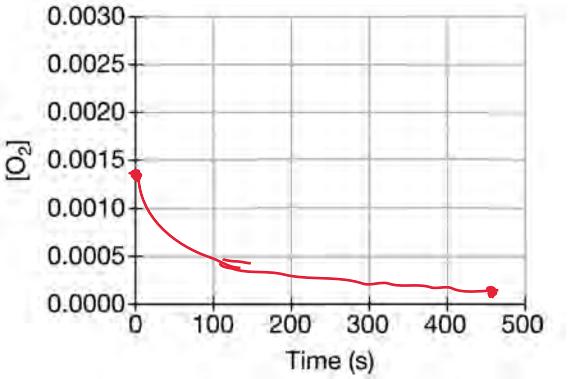
$2\ NO(g) + O_2(g) \rightarrow 2\ NO_2(g)$

Reactant	Initial Concentration
NO	0.0028 M
O_2	0.0014 M

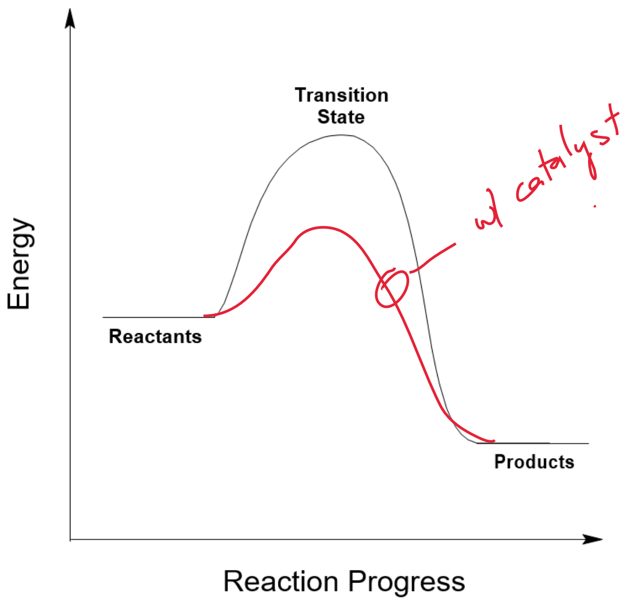
391. The oxidation of NO(g) producing NO_2(g) is represented by the chemical equation shown above. The initial concentration of NO and O_2 are given in the table above. The changes in concentration of NO(g) as a function of time are shown in the following graph.



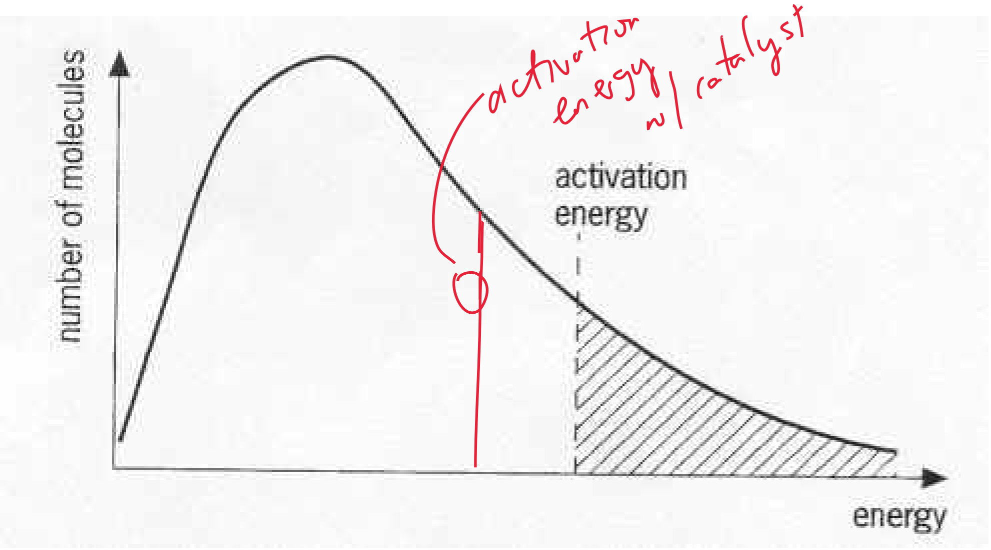
On the graph below, draw the curve for the rate of disappearance of O_2.



392.On the reaction energy diagram below, draw and label the reaction pathway in the presence of a catalyst.



393.On the Boltzmann diagram below draw and label the location of the activation energy in the presence of a catalyst.



394.How does the rate of disappearance of a reactant change with increasing strength of covalent bonds?

As a covalent bond strengthens the rate decreases. A stronger bond requires more energy for the reaction to occur.

395. Complete the following table by giving the element name, element symbol, the complete electron configuration, or Noble gas configuration of the following atoms:

	Element Name	Element Symbol	Complete Electron Configuration	Noble Gas Electron Configuration
1	Fluorine	F	$1s^2 2s^2 2p^5$	$[He] 2s^2 2p^5$
2	Calcium	Ca	$1s^2 2s^2 2p^6 3s^2 3p^6 4s^2$	$[Ar] 4s^2$
3	Cobalt	Co	$1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^7$	$[Ar]\ 4s^2\ 3d^7$
4	Arsenic	As	$1s^2\ 2s^2\ 2p^6\ 3s^2\ 3p^6\ 4s^2\ 3d^{10}\ 4p^3$	$[Ar] 4s^2 3d^{10} 4p^3$
5	Sulfur	S	$1s^2 2s^2 2p^4 3s^2 3p^4$	$[Ne] 3s^2 3p^4$
6	Bromine	Br	$1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^{10} 4p^5$	$[Ar] 4s^2 3d^{10} 4p^5$
7	Iron	Fe	$1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^6$	$[Ar] 4s^2 3d^6$

396. Complete the following table by giving the ion name, ion symbol, the complete electron configuration, or Noble gas configuration of the following ions:

	Element Name	Element Symbol	Complete Electron Configuration	Noble Gas Electron Configuration
1	Fluoride	F^-	$1s^2 2s^2 2p^6$	$[He] 2s^2 2p^6$
2	Calcium ion	Ca^{2+}	$1s^2 2s^2 2p^6 3s^2 3p^6$	$[Ne] 3s^2 3p^6$
3	Iron (II)	Fe^{2+}	$1s^2 2s^2 2p^6 3s^2 3p^6 3d^6$	$[Ar] 3d^6$
4	Phosphide	P^{3-}	$1s^2 2s^2 2p^6 3s^2 3p^6$	$[Ne] 3s^2 3p^6$
5	Zinc	Zn^{2+}	$1s^2 2s^2 2p^6 3s^2 3p^6 3d^{10}$	$[Ar] 3d^{10}$
6	Bromide	Br^-	$1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d^{10} 4p^6$	$[Ar] 4s^2 3d^{10} 4p^6$
7	Iron (III)	Fe^{3+}	$1s^2 2s^2 2p^6 3s^2 3p^6 3d^5$	$[Ar] 3d^5$

397.Complete the orbital notation for each of the elements listed in #395 above.

	1s	2s	2p	3s	3p	3d	4s	4p	
1	↑↓	↑↓	↑↓ ↑↓ ↑						F
2	↑↓	↑↓	↑↓ ↑↓ ↑↓	↑↓	↑↓ ↑↓ ↑↓	↑↓ ↑↓ ↑↓ ↑↓ ↑↓	↑↓		Ca
3	↑↓	↑↓	↑↓ ↑↓ ↑↓	↑↓	↑↓ ↑↓ ↑↓	↑↓ ↑↓ ↑ ↑ ↑			Co
4	↑↓	↑↓	↑↓ ↑↓ ↑↓	↑↓	↑↓ ↑↓ ↑↓	↑↓ ↑↓ ↑↓ ↑↓ ↑↓	↑↓	↑ ↑ ↑	As
5	↑↓	↑↓	↑↓ ↑↓ ↑↓	↑↓	↑↓ ↑ ↑				
6	↑↓	↑↓	↑↓ ↑↓ ↑↓	↑↓	↑↓ ↑↓ ↑↓	↑↓ ↑↓ ↑↓ ↑↓ ↑↓	↑↓	↑↓ ↑↓ ↑	Br
7	↑↓	↑↓	↑↓ ↑↓ ↑↓	↑↓	↑↓ ↑↓ ↑↓	↑↓ ↑ ↑ ↑ ↑			Fe

$$\text{Rate} = k[H_3AsO_4]^2\,[I^-]\,[H_3O^+]^0$$

398. Use the rate law given above to answer the questions that follow.

a. What is the order with respect to H_3AsO_4 ?

2nd

b. What happens to the rate of the reaction if the concentration of H_3AsO_4 is doubled while the others remain the same?

The rate increases by a factor of 4.

c. What is the order with respect to I^- ?

1st

d. What happens to the rate of the reaction if the concentration of I^- is doubled while the others remain the same?

The rate increases by a factor of 2.

e. What is the order with respect to H_3O^+ ?

0th

f. What happens to the rate of the reaction if the concentration of H_3O^+ is doubled while the others remain the same?

Nothing

g. What is the overall order of the reaction?

3rd

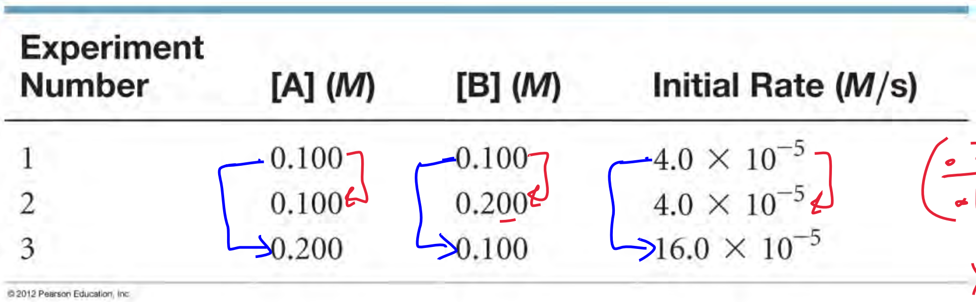
h. What would be the units of k, the rate constant, for this reaction?

$M^{-2}s^{-1}$

399.Complete the table below about the units for the rate constant, k.

Overall Order	Units of Rate Constant
1	s^{-1}
2	$M^{-1}s^{-1}$
3	$M^{-2}s^{-1}$

400.Use the data to answer the questions that follow.



Experiment Number	[A] (*M*)	[B] (*M*)	Initial Rate (*M*/s)
1	0.100	0.100	4.0×10^{-5}
2	0.100	0.200	4.0×10^{-5}
3	0.200	0.100	16.0×10^{-5}

© 2012 Pearson Education, Inc.

B

$\left(\frac{.2}{.1}\right)^x = \left(\frac{4\times10^{-5}}{4\times10^{-5}}\right)$

$x = 0$

a. Determine the rate law.

Rate = $k\,[A]^2[B]^0$

A

$\left(\frac{.2}{.1}\right)^x = \frac{16\times10^{-5}}{4\times10^{-5}}$

$2^x = 4$

$x = 2$

B

$\left(\frac{.2}{.1}\right)^x = \left(\frac{4\times10^{-5}}{4\times10^{-5}}\right)$

$2^x = 1$

$x = 0$

b. Determine the overall order of the reaction.

2nd

c. Determine the value of k, along with units. (Pick any experiment)

Rate = $k[A]^2$

$16\times10^{-5} = k(.2)^2$

$k = .004\ M^{-1}s^{-1}$

d. Determine the rate of the reaction when [A] is 0.0500 M and the [B] is 0.0750 M.

Rate = k [A]^2

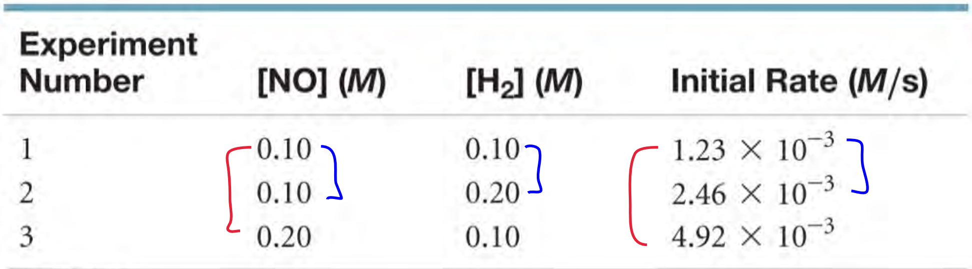
Rate = $(.004)(.05)^2$

Rate = 1×10^{-5} M/s

e. When does the value of k change?

With a change in temperature

401. Use the data to answer the questions that follow.



Experiment Number	[NO] (*M*)	[H_2] (*M*)	Initial Rate (*M*/s)
1	0.10	0.10	1.23×10^{-3}
2	0.10	0.20	2.46×10^{-3}
3	0.20	0.10	4.92×10^{-3}

a. Determine the rate law.

Rate = k [NO]^2 [H_2]^1

NO

$$\left(\frac{.2}{.1}\right)^x = \left(\frac{4.92}{1.23}\right)$$

$$2^x = 4$$

$$x = 2$$

H_2

$$\left(\frac{.2}{.1}\right)^x = \left(\frac{2.46}{1.23}\right)$$

$$2^x = 2$$

$$x = 1$$

b. Determine the overall order of the reaction.

3rd

c. Determine the value of k, along with units.

$Rate = k[NO]^2[H_2]^1$

$1.23 \times 10^{-3} = k(.1)^2(.1)$

$k = 1.23\ M^{-2}s^{-1}$

d. Determine the rate of the reaction when [NO] is 0.100 M and the [H_2] is 0.150 M.

$Rate = k[NO]^2[H_2]$

$Rate = 1.23(.100)(.150)$

$Rate = .00185\ M/s$

e. When does the value of k change?

with a change in temperature

402. Use the data to answer the questions that follow.

Experiment	$[A]_0$	$[B]_0$	$[C]_0$	**Initial rate, v_0 (mol L^{-1} s^{-1})**
1	1.25×10^{-3} M	1.25×10^{-3} M	1.25×10^{-3} M	0.0087
2	2.50×10^{-3} M	1.25×10^{-3} M	1.25×10^{-3} M	0.0174
3	1.25×10^{-3} M	3.02×10^{-3} M	1.25×10^{-3} M	0.0508
4	1.25×10^{-3} M	3.02×10^{-3} M	3.75×10^{-3} M	0.457

a. Determine the rate law.

$Rate = k[A]^1[B]^2[C]^2$

A

$\left(\frac{2.5}{1.25}\right)^x = \left(\frac{.0174}{.0087}\right)$

$2^x = 2$

$x = 1$

B

$\left(\frac{3.02}{1.25}\right)^x = \left(\frac{.0508}{.0087}\right)$

$2.42^x = 5.84$

$x = 2$

C

$\left(\frac{3.75}{1.25}\right)^x = \left(\frac{.457}{.0508}\right)$

$3^x = 9$

$x = 2$

b. Determine the overall order of the reaction.

5th

c. Determine the value of k, along with units.

$Rate = k[A]^1[B]^2[C]^2$

$.457 = k(1.25 \times 10^{-3})^1(3.02 \times 10^{-3})^2(3.75 \times 10^{-3})^2$

$k = 2.85 \times 10^{-12}\ M^{-4}\ s^{-1}$

403. Use the data to answer the questions that follow.

The rate changes because of $[H_2O_2]$ AND $[H^+]$. So 1st figure out H_2O_2

	$[H_2O_2]$	$[I^-]$	$[H^+]$	Rate (M/s)
I	0.100	5.00×10^{-4}	1.00×10^{-2}	0.137
II	0.100	1.00×10^{-3}	1.00×10^{-2}	0.268
III	0.200	1.00×10^{-3}	1.00×10^{-2}	0.542
IV	0.400	1.00×10^{-3}	2.00×10^{-2}	1.084

1.084

a. Determine the rate law.

$Rate = k[H_2O_2]^1[I^-]^1[H^+]^0$

H_2O_2

$\left(\frac{.2}{.1}\right)^x = \left(\frac{.542}{.268}\right)$

$2^x = 2$

$x = 1$

I^-

$\left(\frac{1}{.5}\right)^x = \left(\frac{.268}{.137}\right)$

$2^x = 2$

$x = 1$

$[H^+]$

$\left(\frac{2}{1}\right)^x = \left(\frac{1.084}{.542}\right)$

$2^x = 1$

$x = 0$

Note: For the order of H^+ you have to use Exp. III and IV. But H_2O_2 also changes during exp. III and IV. But we already know how a change in $[H_2O_2]$ changes the reaction - it's 1st order. So doubling the $[H_2O_2]$ in Exp. III and IV also doubles the rate. So the rate should change from 0.542 to 1.084 because of the $[H_2O_2]$. Any further change in the rate will be because of the change in $[H^+]$. Problem #404 is a better example of this type of work.

b. Determine the overall order of the reaction.

2nd

c. Determine the value of k, along with units.

$k = 2740\ M^{-1}\ s^{-1}$

404. Use the data to answer the questions that follow.

	[A]	[B]	[C]	Rate (M/s)
I	0.1	0.05	0.02	0.2
II	0.2	0.05	0.02	0.8
III	0.2	0.15	0.02	2.4
IV	0.4	0.15	0.04	19.2

B/c of [A] the rate should be 9.6

a. Determine the rate law.

Rate = $k\ [A]^2 [B]^1 [C]^1$

A

$$\left(\frac{.2}{.1}\right)^x = \left(\frac{.8}{.2}\right)$$

$2^x = 4$

$x = 2$

B

$$\left(\frac{.15}{.05}\right)^x = \left(\frac{2.4}{0.8}\right)$$

$3^x = 3$

$x = 1$

C

$$\left(\frac{.04}{.02}\right)^x = \left(\frac{19.2}{9.6}\right)$$

$2^x = 2$

$x = 1$

Note on #404a: You have to use Exp III and IV to determine the order of C. B remains constant so we don't have to worry about B. But both [A] and [C] changes. We already know that A is 2nd order so when the [A] doubles from 0.2 to 0.4 the rate should quadruple. So because of A the rate should go from 2.4 to 9.6. Any change beyond 9.6 is as a result of the change in [C]. Since the final rate is 19.2 we can determine that the rate changed from 9.6 to 19.2 because of the change in [C].

b. Determine the overall order of the reaction.

4th

c. Determine the value of k, along with units.

$k = 20\ 000\ M^{-3} s^{-1}$

405. Write the rate law for each of the elementary steps.

	Elementary Step	Rate Law
(A)	$A + B \rightarrow X + Y$	Rate = $k[A][B]$
(B)	$2A + B \rightarrow A_2B$	Rate = $k[A]^2[B]$
(C)	$NO + O_3 \rightarrow NO_2 + O_2$	Rate = $k[NO][O_3]$
(D)	$Cl + CH_4 \rightarrow HCl + CH_3$	Rate = $k[Cl][CH_4]$
(E)	$Ar + O_3 \rightarrow Ar + O_3^*$	Rate = $k[Ar][O_3]$
(F)	$A + A \rightarrow B + C$	Rate = $k[A]^2$
(G)	$O_3 \rightarrow O_2 + O$	Rate = $k[O_3]$
(H)	$O + O_2 + N_2 \rightarrow O_3 + N_2$	Rate = $k[O][O_2][N_2]$
(I)	$A \rightarrow B + C + D$	Rate = $k[A]$

406. When given the rate law of the elementary step, determine the molecularity and give the reactant side of the elementary step.

	Rate Law	Molecularity			Elementary Step (Reactants only)
(A)	Rate = $k[A]$	[✓ Unimolecular]	Bimolecular	Termolecular	$A \rightarrow$
(B)	Rate = $k[NO]^2[O_2]$	Unimolecular	Bimolecular	[✓ Termolecular]	$2NO + O_2 \rightarrow$
(C)	Rate = $k[CO]^2$	Unimolecular	[✓ Bimolecular]	Termolecular	$2CO \rightarrow$
(D)	Rate = $k[A][B]$	Unimolecular	[✓ Bimolecular]	Termolecular	$A + B \rightarrow$

407. A possible reaction mechanism for the conversion of ozone to O_2 is shown below:

Step 1 $O_3(g) \rightarrow O_2(g) + \cancel{O(g)}$

Step 2 $O_3(g) + \cancel{O(g)} \rightarrow 2\,O_2(g)$

$2O_3(g) \rightarrow 3O_2(g)$

a. What is the molecularity of each step?

step 1 = unimolecular

step 2 = bimolecular

b. What is the overall equation?

$2\,O_3(g) \rightarrow 3O_2(g)$

c. What is the intermediate in the reaction? Explain your answer.

$O(g)$ Enters as a product and leaves as a reactant

408. Nitrous oxide, N_2O, is believed to decompose by a two-step reaction mechanism:

Step 1 $N_2O(g) \rightarrow N_2(g) + O(g)$ (slow)

Step 2 $N_2O(g) + O(g) \rightarrow N_2(g) + O_2(g)$ (fast)

a. Write the equation for the overall reaction.

$2N_2O(g) \rightarrow 2N_2(g) + O_2(g)$

b. Write the rate law for the overall reaction.

$Rate = k[N_2O]$

c. Which step has the highest activation energy?

Step 1

409. Consider the reaction mechanism shown below:

$\text{Rate}_F = \text{Rate}_R$ b/c equilibrium

$k[NO]^2 = k^{-1}[N_2O_2]$

Step 1 $\quad NO(g) + NO(g) \underset{k_{-1}}{\overset{k_1}{\rightleftharpoons}} N_2O_2(g)$ $\quad$ (fast, equilibrium)

$\frac{k}{k^{-1}}[NO]^2 = [N_2O_2]$

Step 2 $\quad N_2O_2(g) + Br_2(g) \xrightarrow{k_2} 2\ NOBr(g)$ $\quad$ (slow)

$\text{Rate} = k[N_2O_2][Br_2]$

Determine the rate law for the reaction using the steady-state approximation. $\text{Rate} = (k)\left(\frac{k}{k_{-1}}\right)[NO]^2[Br_2]$

$\text{Rate} = k[NO]^2[Br_2]$

410. Another mechanism for the conversion of ozone to O_2 shows the possible destruction in the upper atmosphere by NO(g), as shown below.

Step 1 $\quad O_3(g) + \cancel{NO}(g) \rightarrow \cancel{NO_2}(g) + O_2(g)$ $\quad$ (slow)

Step 2 $\quad \cancel{NO_2}(g) + O(g) \rightarrow \cancel{NO}(g) + O_2(g)$ $\quad$ (fast)

a. What is the overall equation?

$O_3(g) + O(g) \longrightarrow 2O_2(g)$

b. What is the intermediate in the equation?

$NO_2(g)$

c. What is the role of nitrogen monoxide, NO, in the equation? Explain your reasoning.

Catalyst. Enters as a reactant and leaves as a product

d. What is the molecularity of each step in the mechanism?

step1: bimolecular

step2: bimolecular

e. What is the rate law?

$\text{rate} = k[O_3]$

411. Experiments were conducted to study the rate of the reaction represented by the equation below. Initial concentrations and rates of reaction are given in the table.

$$2\ NO(g) + 2\ H_2(g) \rightarrow N_2(g) + 2\ H_2O(g)$$

Experiment	[NO]	[H_2]	Rate
1	0.0060	0.0010	1.80×10^{-4}
2	0.0060	0.0020	3.60×10^{-4}
3	0.0010	0.0060	0.30×10^{-4}
4	0.0020	0.0060	1.20×10^{-4}

a. Determine the order for each of the reactants, NO and H_2, from the data given and show your reasoning.

NO = 2nd order

H_2 = 1st order

NO

$$\left(\frac{.002}{.001}\right)^x = \left(\frac{1.2}{.3}\right)$$

$$2^x = 4$$

$$x = 2$$

H_2

$$\left(\frac{.002}{.001}\right)^x = \left(\frac{3.6}{1.8}\right)$$

$$2^x = 2$$

$$x = 1$$

b. Write the overall rate law of the reaction.

$$Rate = k[NO]^2[H_2]^1$$

c. Calculate the value of the rate constant, k, for the reaction. Include units.

$$k = 5000\ M^{-2}\ s^{-1}$$

d. The following sequence of elementary steps is a proposed mechanism for the reaction:

Step 1 $NO + NO \rightarrow \cancel{N_2O_2}$

Step 2 $\cancel{N_2O_2} + H_2 \rightarrow H_2O + \cancel{N_2O}$

Step 3 $\cancel{N_2O} + H_2 \rightarrow N_2 + H_2O$

i. Based on the data presented, which of the above steps is the rate-determining step? Explain your reasoning.

Step 2 is the rds. The rds must agree w/ the experimental rate law. Step 2 agrees w/ the rate law.

ii. Show that the mechanism is consistent with the observed rate law and the overall stoichiometry of the reaction.

↳ "balanced equation"

$2NO + 2H_2 \rightarrow N_2 + 2H_2O$

The rxn mechanism adds up to the overall balanced equation given.

412. For a hypothetical chemical reaction that has the stoichiometry 2X + Y → Z, the following initial rate data were obtained. All the measurements were made at the same temperature.

Rate (M/min)	[X] (M)	[Y] (M)
7.0×10^{-4}	0.2	0.2
1.4×10^{-3}	0.4	0.2
2.8×10^{-3}	0.4	0.4
4.2×10^{-3}	0.6	0.6

a. Give the rate law for this reaction from the data above.

$rate = k[X]^1[Y]^1$

b. Calculate the specific rate constant for this reaction and specify its units.

$k = .0175\ M^{-1}s^{-1}$

c. The reaction is carried out with [X] = 0.80 molar and [Y] = 0.60 molar.

i. Determine the rate of the reaction.

$Rate = k[X]^1[Y]^1$

$Rate = (.0175)(.8)(.6)$

$Rate = .0084\ M/min$

ii. A student claims that you can determine the [Z] after 20 minutes by using the reaction rate determined in 412ci. Explain why the student is wrong.

The rate determined in 412ci is the instantaneous rate. As the reaction proceeds the rate decreases as reactants are converted to products. As the [reactants] decreases there are fewer collisions and as a result the rate changes from that calculated in 412ci.

d. Select from the mechanisms below the one most consistent with the observed data. Explain your reasoning for each mechanism.

React ion Mechanism 1		Reaction Mechanism 2		Reaction Mechanism 3	
$X + Y \rightarrow M$	Slow	$X + X \rightleftharpoons M$	Fast	$Y \rightarrow M$	Slow
$X + M \rightarrow Z$	Fast	$Y + M \rightarrow Z$	Slow	$M + X \rightarrow N$	Fast
				$N + X \rightarrow Z$	Fast

Two things must be true:
1. Does the slow step agree with the experimental rate law?
2. Does it add up to the balanced equation?

Mech 1

① rate = $k[X][Y]$ good

② $2x + Y \rightarrow Z$ good

Mech 2

① rate = $k[X]^2[Y]$ bad

② $2x + Y \rightarrow Z$ good

Mech 3

① rate = $k[Y]$ bad

② $2x + Y \rightarrow Z$ good

The best mechanism is Mechanism 1 b/c it agrees w/ the rate law and agrees w/ the overall balanced equation.

e. Identify any intermediates and catalysts in the reaction mechanisms in 412.d above.

Mech 1

M = intermediate

Mech 2

M = intermediate

Mech 3

M = intermediate

N = intermediate

Review Topic 4.6

Perform the following calculations that involve titration to determine an unknown.

413. Hydrochloric acid and calcium hydroxide react as shown below.

__2__ HCl(aq) + _____ $Ca(OH)_2(aq) \rightarrow$ _____ $CaCl_2(aq)$ + __2__ H_2O

a. Balance the chemical reaction.
b. Determine the molarity of the hydrochloric acid if 25.2 mL is in equivalence with 13.9 mL of 0.23 M calcium hydroxide.

M x L = mols

(.23M)(13.9mL) = mmols

$$\frac{3.197 \text{ mmols } Ca(OH)_2}{} \Big| \frac{2\ HCl}{1\ Ca(OH)_2} = 6.394 \text{ mmols } HCl$$

$M = \frac{n}{L}$

$= \frac{6.394 \text{ mmols}}{25.2 \text{ mL}}$

M = 0.254 M

414. A titration is carried out in an acidic solution between potassium permanganate and sodium chloride. The balanced reaction is shown below:

$$11\ H^+ (aq) + 2\ MnO_4^- (aq) + 5\ Cl^- (aq) \rightarrow 2\ Mn^{2+} (aq) + 5\ HClO (aq) + 3\ H_2O (l)$$

a. Determine the molarity of the sodium chloride, NaCl(aq), if 22.0 mL reacts completely with 13.1 mL of 0.35 M potassium permanganate.

$M = \frac{n}{L}$

M x L = n

(.35)(13.1) = n

4.585 mmols = $n_{MnO_4^-}$

$$\frac{4.585 \text{ mols } MnO_4^-}{} \Big| \frac{5Cl^-}{2MnO_4^-} = 11.46 \text{ mmols } Cl^-$$

$M = \frac{n}{L}$

$M = \frac{11.46}{22}$

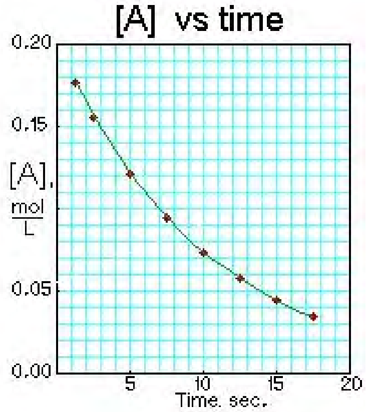
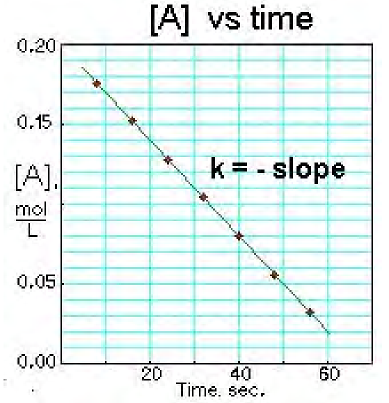
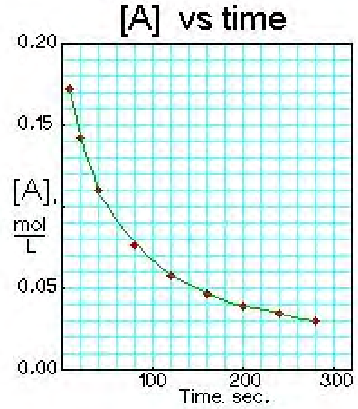
M = .521 M NaCl

b. Determine the mass of NaCl in the solution from 414a.

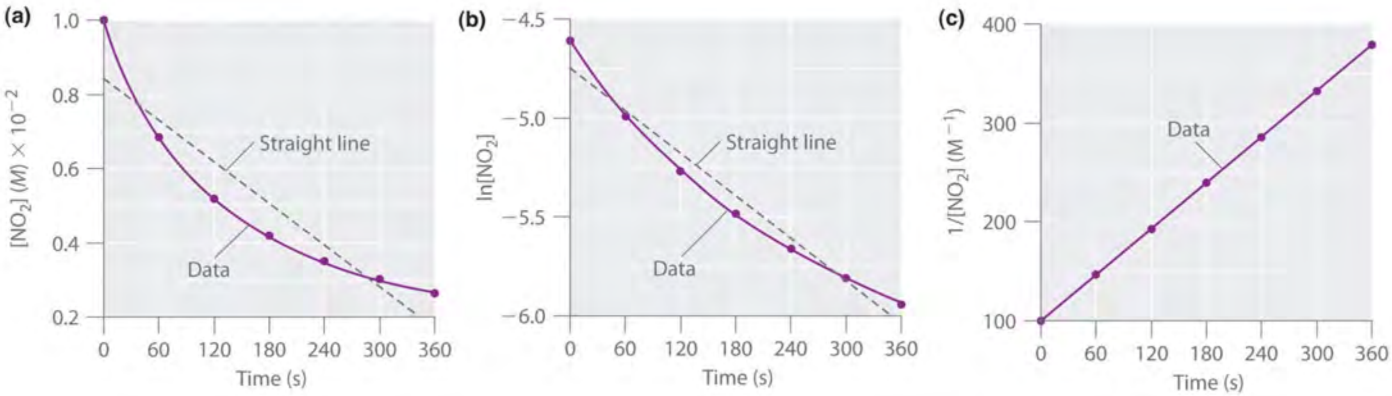
$$\frac{11.46 \text{ mmols NaCl}}{} \Big| \frac{1 \text{ mol}}{1000 \text{ mmols}} = 0.01146 \text{ mols NaCl}$$

$$\frac{.01146 \text{ mols NaCl}}{} \Big| \frac{58.44}{1 \text{ mol}} = 0.670 \text{ g NaCl}$$

415. Complete the table below.

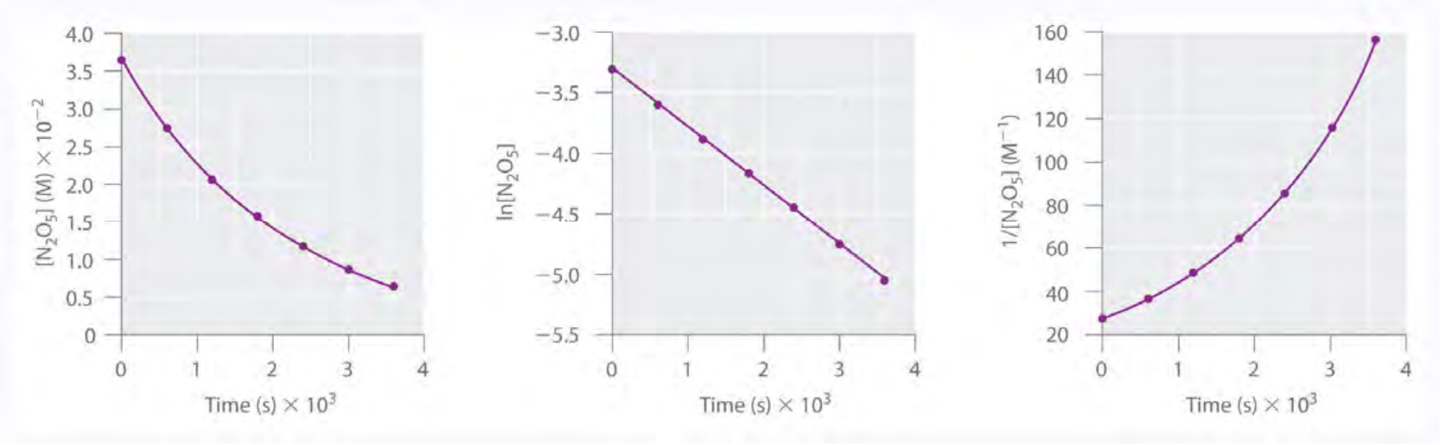
			
Order?	1st	0th	2nd
How do you know?	Smooth curve	Straight line	More of a "L" curve (exponential)
What must be plotted on the y-axis to give a straight line?	ln [A]	[A]	$\frac{1}{[A]}$

416. What is the order of the reaction with respect to NO_2 if the data from the lab is plotted as shown in the graphs below. Explain your reasoning.



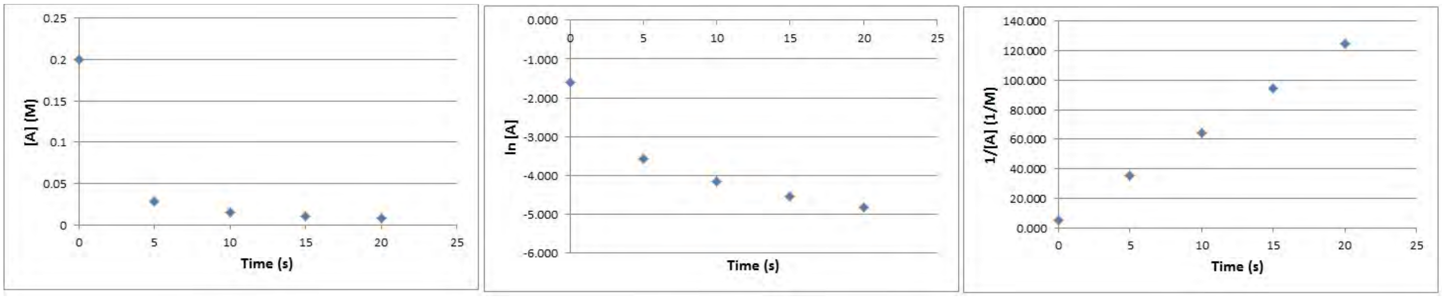
2nd order. A plot of $1/[NO_2]$ vs. time gives a straight line

417. What is the order of the reaction with respect to N_2O_5 if the data from the lab is plotted as shown in the graphs below. Explain your reasoning.



1st order. A plot of $\ln[N_2O_5]$ v. time gives a straight line.

418. Data from the lab is plotted below. Determine the order of the reaction with respect to A. Explain your reasoning.



2nd order. A plot of 1/[A] vs. time gives a straight line.

419. Complete the table using the AP Chemistry Equations and Constants Sheet

Order of Reaction	0	1	2
Integrated Rate Law	$[A]_t - [A]_0 = -kt$	$\ln[A]_t - \ln[A]_0 = -kt$	$(1/[A]_t) - (1/[A]_0) = kt$
Equation for Half-life		$t_{1/2}=0.693/k$	

420. Determine the rate constant or the half-life for the following first order reactions.

a. Half-life of 36 s.

$t_{1/2} = \frac{0.693}{k}$ $36 = \frac{.693}{k}$ $k = .01925\ s^{-1}$

b. Rate constant of 2.00 x 10 $^{-2}$ s $^{-1}$

$t_{1/2} = \frac{0.693}{k}$ $t_{1/2} = \frac{.693}{2\times10^{-2}}$ $t_{1/2} = 34.65\ s$

c. k = 25 s $^{-1}$

$t_{1/2} = \frac{.693}{k}$ $t_{1/2} = \frac{.693}{25}$ $.028\ s = t_{1/2}$

421. Use your answers to 420.b and 420.c to discuss the relationship between half-life and the rate constant. As the rate constant increases in magnitude, does the half-life increase, decrease, or remain the same?

As the rate constant increases the half-life decreases.

422. A certain reaction is first order. Determine the amount of time required for the reactant to go from 0.500 M to 0.250 M if the rate constant is 1200. s $^{-1}$.

$\ln[A]_t - \ln[A]_0 = -kt$ OK $t_{1/2} = \frac{0.693}{K}$

$\ln(.250) - \ln(.500) = -1200\,t$

$t_{1/2} = 5.78\times10^{-4}\ s$

$t = 5.78\times10^{-4}\ s$

423. Determine the order for each reaction given the data below. Explain your answer.

Time	Reaction 1			Reaction 2			Reaction 3.		
	[A]	ln[A]	1/[A]	[B]	ln[B]	1/[B]	[C]	ln[C]	1/[C]
0	150	5	0.0067	285.7	5.65	0.00350	0	0	0
15	75	4.3	0.013	200.0	5.30	0.00500	12.25	2.51	0.0816
30	38	3.6	0.027	153.8	5.04	0.00650	24.5	3.20	0.0408
45	19	2.9	0.053	125.0	4.83	0.00800	36.75	3.60	0.0272

1st order. ln[A] changes by the same amount over time which would give a straight line when you plot ln[A] vs. time.

2nd order. 1/[B] changes by the same amount over time which would give a straight line when you plot 1/[B] vs. time.

0th order. [C] changes by the same amount over time which would give a straight line when you plot [C] vs. time

424. Complete the table below for half-life if the reaction is first order. Assume the half-life is 15 seconds.

Percentage of parent nuclide remaining	Percentage decayed	Number of half-lives	Elapsed time (s)
50%	50%	1	15
25%	75%	2	30
12.5%	87.5%	3	45
6.25%	93.75%	4	60

425. A certain reactant is second order. Determine the final concentration if the initial concentration is 0.450 M and 15 seconds have passed. The rate constant, k, is 1.24 x 10^{-2} M^{-1}s^{-1}.

$$\frac{1}{[A]_t} - \frac{1}{[A]_0} = kt$$

$$\frac{1}{x} - \frac{1}{.450} = (1.24 \times 10^{-2})(15)$$

$$\frac{1}{x} - 2.22 = 0.186$$

$$\frac{1}{x} = 2.408$$

$$x = .415\,M = [A]_t$$

426. Determine the beginning concentration of a zeroth order reaction if the concentration after 25 seconds is 0.500 M and the rate constant is 1 x 10^{-3}.

$[A]_t - [A]_0 = -kt$

$(.500) - x = -(1 \times 10^{-3})(25)$

$-x = -.025 - .500$

$x = .525M = [A]_0$

427. A rigid 2.0 L container at 1.5 atm contains CO_2, O_2, and N_2. The partial pressure of CO_2 is 0.35 atm.

a. Determine the mole fraction of CO_2.

$P_{CO_2} = \chi_{CO_2} P_{tot}$

$0.233 = \chi_{CO_2}$

$(.35) = \chi_{CO_2}(1.5)$

$.35/1.5 = \chi_{CO_2}$

b. The mole fraction of O_2 is 0.467. Determine the partial pressure of O_2.

$P_{O_2} = \chi_{O_2} P_{tot}$

$P_{O_2} = (.467)(1.5)$

$P_{O_2} = .701 \text{ atm}$

c. Determine the mole fraction of N_2 …

i. By using the mole fraction of O_2 and CO_2.

$\chi_{N_2} = 1 - \chi_{O_2} - \chi_{CO_2}$

$\chi_{N_2} = 1 - .467 - .233$

$\chi_{N_2} = 0.3$

ii. By using the total pressure of the container and the partial pressure of N_2.

$P_{tot} = P_{N_2} + P_{O_2} + P_{CO_2}$

$1.5 = P_{N_2} + .701 + .35$

$P_{N_2} = .449 \text{ atm}$

$P_{N_2} = \chi_{N_2} P_{tot}$

$0.449 = \chi_{N_2}\, 1.5$

$0.300 = \chi_{N_2}$

428. Determine the moles of each gas if the container from #427 above is at 30 °C.

a. Moles of CO_2

$PV = nRT$

$(.35\,atm)(2.0\,L) = n(.08206)(303\,K)$

0.0282 mols CO_2

b. Moles of O_2

$PV = nRT$

$(.701\,atm)(2.0\,L) = n(.08206)(303\,K)$

0.0564 mols $= n$

c. Moles of N_2

$PV = nRT$

$(.449\,atm)(2.0\,L) = n(.08206)(303\,K)$

$.0361$ mols $= n$

Unit 6

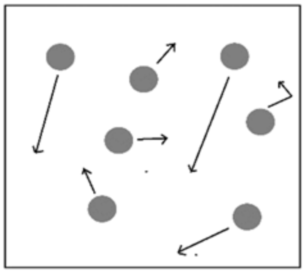
429. Place an X in the box of the substance that will have the greater average kinetic energy. If the average KE will be the same then place an X in the box for "Same average KE".

☐ 1 L of water at 25 °C	☒ 1 L of water at 35 °C	☐ Same average KE
[✓ ☐ 1 L of water at 25 °C]	☐ 5 L of C_6H_6 at 25 °C	☒ Same average KE
☒ 25 g of gold at 55 °C	☐ 25 g of lead at 35 °C	☐ Same average KE
☐ 1.0 M KNO_3 at 80 °C	[✓ ☐ 2.0 M NaOH at 80 °C]	☒ Same average KE
☐ 1.00 mole of Al at 30 °C	☒ 5.00 mole of Cu at 80 °C	☐ Same average KE

430. For the answers in 429 above that have the same average KE, **circle** the substance with the greater velocity.

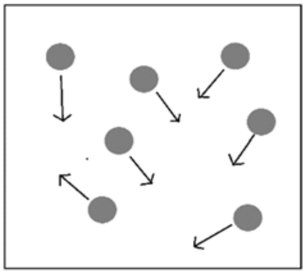
431. Answer the questions that follow about particle pictures of a gas and kinetic energy.

a. What do the arrows indicate? Why are they at different lengths?



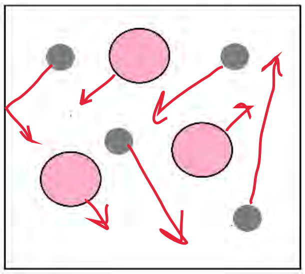
The arrows indicate velocity. They are different lengths because they have different velocities.

b. What is wrong with the arrows in the picture below?



They are all the same length. The particles would not have all the same velocities.

c. Draw arrows for the two gases below that are at the same temperature.



Note: The large atoms should have arrows shorter than the small atoms.

432. A 50 g sample of a metal is heated to 75.0 °C and placed into 50 g of water at 25.0 °C. The temperature of the water rose to reach a final temperature of 26.4 °C.

a. What is the final temperature of the metal? Justify your reasoning.

26.4°C. They are in contact with each other therefore they will reach thermal equilibrium

b. Which substance had the larger change in temperature, the metal or the water?

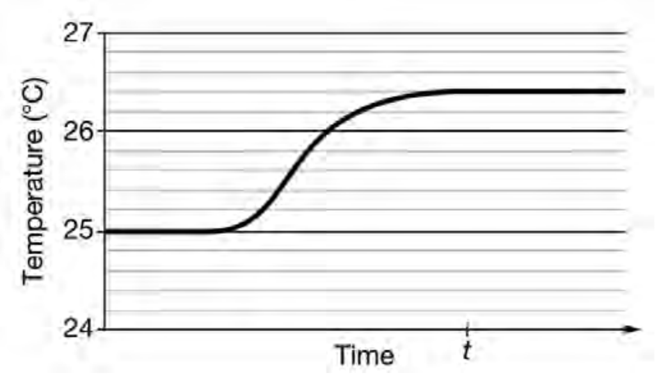
The metal.

c. Which substance had the largest change in average energy, the metal or the water? Justify your answer.

The metal. The larger change in temperature means a larger change in average energy. Temperature is the measure of the average KE.

d. Explain on the particulate level how the temperature of the water increases after addition of the metal. Your answer should explain the role of kinetic energy.

The metal is at a higher temperature therefore the KE of the metal atoms is greater than the KE of the water molecules. As the slower moving water comes into contact with the high KE metal atoms, energy is transferred from the metal atoms to the water molecules. As a result, the water molecules begin speeding up which translates into higher KE and higher temperature.



e. The student claims that thermal equilibrium is reached at time t.

i. What is thermal equilibrium?

The point where the average KE for both substances is equal.

ii. Is the student correct about time t? Justify your answer by referring to the graph.

Yes. At time "t" the temperature stops changing and has reached a maximum.

iii. What can be said about the average kinetic energy of the metal and the water at time t?

They are the same.

433. A sample of iron at 25 °C is placed into a sample of water at 75 °C.

a. Before the iron is placed into the water, how does the average kinetic energy of the iron compare to the average kinetic energy of the water?

The average KE of the iron is less than the average KE of the water.

b. What happens to the temperature of the iron once it is placed into the water?

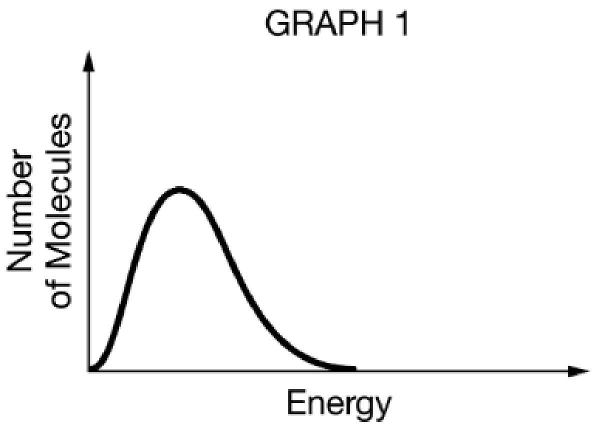
The temperature increases.

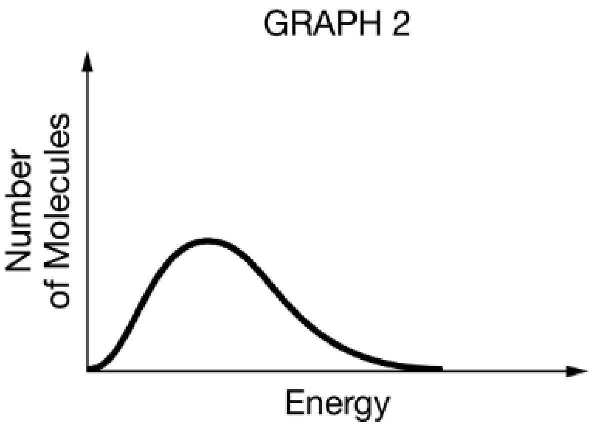
c. Explain on the particulate level how this system will reach thermal equilibrium.

The water particles will collide with the slower iron particles and transfer energy. This will continue to happen until the average KE of both are the same.

d. After the system has reached thermal equilibrium, how does the average kinetic energy of the iron compare to the average kinetic energy of the water?

They are the same.





434. The graphs above show Maxwell-Boltzmann distributions for one-mole samples of Ar*(g)*.

a. Which graph is at a higher temperature? Explain how you know this.

Graph 2. There are more particles at a higher energy. The graph is shifted to the right.

b. The two samples are mixed. What would happen to the average kinetic energy of …

i. Graph 1

Increase.

ii. Graph 2

Decrease.

Review Topic 1.3 and 1.4, 4.5

435. Determine the molar mass of the following compounds:
   a. C_2H_4

$28.052\ g/mol$

   b. $Ca(C_2H_3O_2)_2$

$158.168\ \frac{g}{mol}$

   c. C_5H_{12}

$72.146g\ \frac{g}{mol}$

436. Determine the number of moles for 5.00 grams of each of the following compounds:
   a. C_2H_4

$$\begin{array}{l|l} 5.00g\ C_2H_4 & 1 mol \\ \hline & 28.052 \end{array} = 0.178\ mols\ C_2H_4$$

   b. $Ca(C_2H_3O_2)_2$

$$\begin{array}{l|l} 5.00g\ Ca(C_2H_3O_2)_2 & 1 mol \\ \hline & 158.168g \end{array} = 0.0316\ mols\ Ca(C_2H_3O_2)_2$$

   c. C_5H_{12}

$$\begin{array}{l|l} 5.00g\ C_5H_{12} & 1 mol \\ \hline & 72.146g \end{array} = 0.0693\ mols\ C_5H_{12}$$

437. Determine the percent composition by mass of carbon in the following compounds:

a. C_2H_4

$$\frac{2 \times C}{C_2H_4} \times 100 = \frac{24.02}{28.052} \times 100 = 85.6\% \ C$$

b. $Ca(C_2H_3O_2)_2$

$$\frac{4 \times C}{Ca(C_2H_3O_2)_2} \times 100 = \frac{48.04}{158.168} \times 100 = 30.4\% \ C$$

c. C_5H_{12}

$$\frac{5 \times C}{C_5H_{12}} \times 100 = \frac{60.05}{72.146} \times 100 = 83.2\% \ C$$

438. How many grams of carbon would be in 12.8 grams of $Ca(C_2H_3O_2)_2$?

$$12.8 g \times .304 = 3.89 g \ C$$

439. A 12.0 g sample of the hydrocarbon C_7H_{16} is combusted and the products analyzed. The products of the combustion are water and carbon dioxide, CO_2. Using the percent composition of C in C_7H_{16}, determine the mass of CO_2 formed in the reaction.

$$\frac{7 \times C}{C_7H_{16}} = \frac{84.07}{100.198} = 0.839 \text{ C}$$

$$12.0g \times 0.893 = 10.07 g\, C$$

$$\% C = \frac{g\,C}{g\,CO_2}$$

$$\frac{12.01}{44.01} = \frac{10.07 g}{x\, g\, CO_2}$$

$$0.273 = \frac{10.07}{x}$$

$$x = 36.9g\ CO_2$$

440. A 0.3920 g sample of an organic compound consisting of only C, H, and O is analyzed via combustion analysis. The combustion reaction yields 0.7182 g of CO_2 and 0.1960 g of H_2O. Determine the empirical formula of the organic compound.

$$C_xH_yO_z + O_2 \rightarrow CO_2 + H_2O$$

$$.3920g\ C_?H_?O_?$$
$$-\ .196\,g\ C$$
$$-\ .0219g\ H$$
$$0.1741g\ O$$

$$.7182g\ CO_2 \times \frac{12.01}{44.01} = 0.196g\ C$$

$$.1960g\ H_2O \times \frac{2.016}{18.016} = 0.0219g\ H$$

$$\frac{.196g\ C}{} \Big| \frac{1 mol}{12.01 g} = .0163 mols\ C / .0109 = 1.5\ C$$

$$\frac{.0219g\ H}{} \Big| \frac{1 mol}{1.008 g} = .0217 mols\ H / .0109 = 2\ H$$

$$\frac{.1741g\ O}{} \Big| \frac{1 mol}{16 g} = .0109 mols\ O / .0109 = 1\ O$$

$$C_{1.5}H_2O$$
$$\times 2$$
$$C_3H_4O_2$$

Section 6.02

Topic 6.4 Worksheet

441. Describe each variable in the equation below, including units.

$$q = m\, c\, \Delta T$$

heat (J)

mass (g)

Specific heat capacity (J/g°C)

change in temp (°C or K)

442. What is the first law of thermodynamics?

Energy is conserved in chemical and physical processes.

443. A reaction is carried out in a Styrofoam calorimeter. The reaction is known to release 15.2 kJ of energy. However, when the reaction is carefully measured and all calculations performed the reaction only gave off 15 kJ of energy.

a. Explain why it is not possible to lose 0.2 kJ of energy.

The 1st Law of Thermodynamics states that energy is conserved.

b. If not to the surroundings, where could that energy have gone?

Into the calorimeter.

444. Two different metals of equal mass at the same initial temperature are given the same amount of energy.

a. Explain why the two metals would not have the same temperature.

The two metals most likely have two different specific heat capacities. The different specific heat capacities means the metals will gain energy at different rates.

b. Would the metal with the higher specific heat capacity have a higher temperature, lower temperature, or the same temperature as the metal with the lower specific heat capacity? Justify your answer with a calculation for both metals.

The metal with the higher specific heat capacity would have a lower temperature.

higher specific heat

$q = mc\Delta t$

$1 = (1)(1)\Delta t$

$1\,°C = \Delta t$

lower specific heat

$q = mc\Delta t$

$1 = (1)(.5)\Delta t$

$2\,°C = \Delta t$

445. A closed system of a piece of ice and metal is created. The 5.00 g piece of ice at −5.0 °C melts to liquid water at 10. °C when placed on the metal. The metal has an initial temperature of 75 °C.

a. Why did the ice melt?

Energy was transferred from the metal to the ice.

b. Is the energy of the entire system after the ice melts less than, greater than, or equal to the energy of the system before the ice melts? Explain your reasoning.

The energy is equal. It is a closed system. Energy lost by the metal is gained by the ice/water.

c. Which has more energy, the ice or the liquid water?

The liquid water has more energy.

d. What happened to the overall energy of the metal?

It decreased.

e. What happened to the overall energy of the H_2O?

It increased.

f. Explain your answers to d and e in terms of the first law of thermodynamics.

Energy is conserved therefore the energy lost by the metal will be gained by the water.

446. Convert the following to either specific heat capacity (J/g °C) or molar heat capacity (J/mol °C) .

a. Convert to molar heat capacity: $H_2O(l)$ = 4.184 J/g °C

$$\frac{4.184\ \text{J}}{\text{g}\ ^\circ\text{C}} \cdot \frac{18.016\ \text{g}}{1\ \text{mol}} = 75.4\ \text{J}/\text{mol}\ ^\circ\text{C}$$

b. Convert to specific heat capacity: $Al(s)$ = 24.3 J/mol °C

$$\frac{24.3\ \text{J}}{\text{mol}\ ^\circ\text{C}} \cdot \frac{1\ \text{mol}}{26.98\ \text{g}} = .901\ \text{J}/\text{g}\ ^\circ\text{C}$$

c. Convert to molar heat capacity: Au(s) = 0.129 J/g °C

$$\frac{.129\ \text{J}}{\text{g}\ ^\circ\text{C}} \cdot \frac{197\ \text{g}}{1\ \text{mol}} = 25.4\ \text{J}/\text{mol}\ ^\circ\text{C}$$

d. Convert to specific heat capacity: Ethylene glycol = 78 J/mol °C (molar mass = 62.07 g/mol)

$$\frac{78\ \text{J}}{\text{mol}\ ^\circ\text{C}} \cdot \frac{1\ \text{mol}}{62.07\ \text{g}} = 1.26\ \text{J}/\text{g}\ ^\circ\text{C}$$

447. A 200g metallic cube is placed into a coffee cup calorimeter with 100 g of water at 20 °C. If the metal releases 7.2 kJ of energy to the water, what is the final temperature of the water? (The specific heat capacity of water is 4.2 J/g °C)

$q = mC\Delta t$

$7200\text{ J} = (100\text{g})(4.2)\Delta t$

$\frac{7200}{(100)(4.2)} = \Delta t$

$17.1°C = \Delta t$

$\Delta t = t_{final} - t_{initial}$

$17.1 = t_f - 20$

$37.1°C = t_f$

448. A neutralization reaction between NaOH and HCl is carried out in a coffee cup calorimeter. A 10.0 mL sample of a 1.0 M NaOH solution is titrated against a 10.0 mL sample of a 1.0 M HCl solution, as shown in the equation below.

$$HCl(aq) + NaOH(aq) \rightarrow NaCl(aq) + H_2O(l)$$

If both solutions started at a temperature of 20 °C and ended at 26.0 °C, what is the ΔH_{rxn} in kJ/mol$_{rxn}$? (Assume the specific heat capacity of the total solution is the same as water, 4.18 J/g °C, and the density is 1.0 g/mL.)

$q = mc\Delta t$

$q = (20)(4.18)(26-20)$

$q = 501.6\text{J}$

energy per 10mL of 1.0M HCl

$\frac{501.6\text{ J}}{.01 \times 1}$

volume

molarity

used to find moles of HCl (M x L)

$\frac{50160\text{ J}}{\text{mol HCl}} \times \frac{1\text{ kJ}}{1000\text{ J}} \times \frac{1\text{ mol HCl}}{1\text{ mol}_{rxn}} = -50.16\frac{\text{kJ}}{\text{mol}_{rxn}}$

(where did the "-" come from in the answer? I know it's exothermic so I just put it there. The temp increased so it has to be exothermic.)

449. A student conducts a reaction between hydrochloric acid and calcium hydroxide, as shown in the reaction below:

$$2\ HCl(aq) + Ca(OH)_2(aq) \rightarrow CaCl_2(s) + H_2O(l)$$

In the reaction, the student mixes 10.0 mL of 1.00 M HCl and 10.0 mL of 0.500 M $Ca(OH)_2$. It is determined that 0.870 kJ of energy is absorbed in the reaction.

a. Is the reaction endothermic or exothermic? Should the temperature of the reaction increase or decrease?

Endothermic. The temp of the reaction will decrease.

b. Determine the ΔH_{rxn}, in kJ/mol_{rxn}, by using the moles of HCl.

$$\frac{.870\ kJ}{.010L \cdot \times 1M} = \frac{87\ kJ}{mol\ HCl} \left| \frac{2\ HCl}{1\ mol_{rxn}} \right. = 174\ kJ/mol_{rxn}$$

c. Determine the ΔH_{rxn}, in kJ/mol_{rxn}, by using the moles of $Ca(OH)_2$.

$$\frac{.870\ kJ}{.010L \times .5M} = \frac{174\ kJ}{mol\ Ca(OH)_2} \left| \frac{1\ Ca(OH)_2}{1\ mol_{rxn}} \right. = 174\ kJ/mol_{rxn}$$

450. Determine the amount of heat needed to melt 125 g of Au originally at 298 K.

Au melting point (°C)	1064
C_{Au} (J/g °C)	0.128
ΔH_{fusion} (kJ mol^{-1})	12.55

solid to solid (heating up)

$q = mC\Delta t$

$q = (125)(.128)(1039K)$

$q = 16\,624\ J$

solid to liquid

$q = n\,\Delta H$

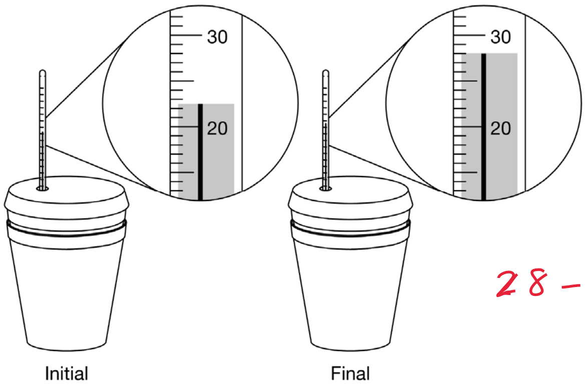
$q = \left(\frac{125}{197}\right) 12.55$

$q = 7.96\ kJ$

$q_{tot} = 16.624\ kJ + 7.96\ kJ$

$q_{tot} = 24.6\ kJ$

451. What is the temperature change for the reaction shown below?



$28 - 22.5 = 5.5\,°C$

452. A student conducts an experiment by mixing 10.0 mL of 1.0 M HCl and 10.0 mL of 1.0 M NaOH. The energy of the reaction is determined. She then mixes 20.0 mL of 1.0 M HCl and 10.0 mL of 1.0 M NaOH. The results of the experiment are shown below.

Experiment	HCl	NaOH	Energy
#1	10.0 mL of 1.0 M HCl	10.0 mL of 1.0 M NaOH	X kJ
#2	20.0 mL of 1.0 M HCl	10.0 mL of 1.0 M NaOH	??? kJ

Would the energy of experiment #2 be greater than, less than, or equal to the energy for experiment #1?

The energy would be equal to expt. #1. The NaOH is the limiting reactant in expt #2.

453. The hydrocarbon pentane, C_5H_{12} (molar mass 72.15g) is combusted to produce carbon dioxide and water, as shown in the unbalanced reaction below.

$$\_\_\_\_\_ C_5H_{12}(l) + \underline{\ 8\ } O_2(g) \rightarrow \underline{\ 6\ } H_2O(l) + \underline{\ 5\ } CO_2(g)$$

a. The complete combustion of 5.00 g of pentane releases 243 kJ of heat. On the basis of this information, calculate the value of ΔH for the complete combustion of one mole of pentane.

$$\begin{array}{c|c} 243\text{ kJ} & 72.146\ \cancel{\text{g}} \\ \hline 5.00\ \cancel{\text{g}}\ C_5H_{12} & 1\text{ mol} \end{array} = -3510\text{ kJ}/\text{mol}\ C_5H_{12}$$

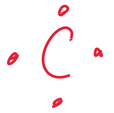
b. Determine the ΔH_{rxn}, in kJ/mol_{rxn}.

$$\begin{array}{c|c} -3510\text{ kJ} & 1\text{ mol } C_5H_{12} \\ \hline \text{mol}\ C_5H_{12} & 1\text{ mol}_{rxn} \end{array} = -3510\text{ kJ}/\text{mol}_{rxn}$$

Review Topic 2.5, 2.7, 2.6

454. Give the Lewis electron dot diagram for the following elements.

a. Carbon

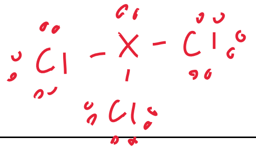
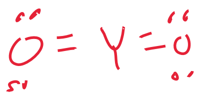
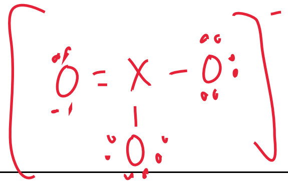
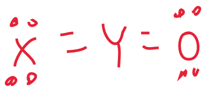


c. Nitrogen

b. Magnesium

d. Oxygen

455. Draw the Lewis diagram for the following compounds. For each compound, X has 5 valence electrons, Y has 4 valence electrons

Compound	Lewis Structure	Molecular Geometry	Polar or Nonpolar	Hybridization (central atom)
XCl_3 5 21		trigonal pyramid	polar	sp^3
YO_2 4 12		linear	non polar	sp
XO_3^- 5 18+1		trigonal planar	non polar	sp^2
OYX^- 6 4 5+1		linear	polar	sp

456. Determine the formal charge on each atom for the Lewis structures you drew in 455 above.

XCl_3

$X = 5 - (\frac{1}{2}(6) + 2) = 0$

$Cl: 7 - (\frac{1}{2}(2) + 6) = 0$

YO_2

$Y: 4 - (\frac{1}{2}(4) + 0) = 0$

$O = 6 - (\frac{1}{2}4 + 4) = 0$

XO_3^-

$X: 5 - (\frac{1}{2}(8) + 0) = +1$

$O_{d.b.} = 6 - (\frac{1}{2}(4) + 4) = 0$

$O_{s.b}: 6 - (\frac{1}{2}(2) + 6) = -1$

OYX^-

$O = 6 - (\frac{1}{2}(4) + 4) = 0$

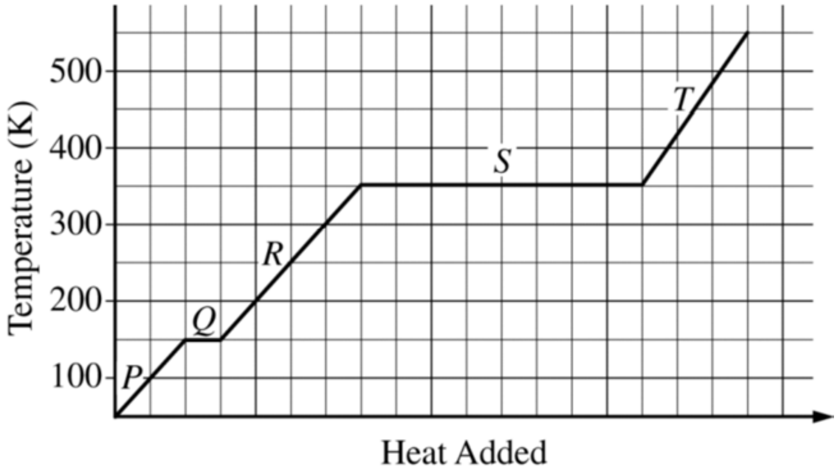
$Y = 4 - (\frac{1}{2}(8) + 0) = 0$

$X = 5 - (\frac{1}{2}(4) + 4) = -1$

Section 6.03

Topic 6.5 Worksheet

Use the heating curve of pure ethanol, CH_3CH_2OH , given below to answer questions 457 to 461 below.



457. What is the melting point of ethanol? What is the boiling point of ethanol?

150 K

350K

458. Explain why line Q is shorter than line S by referring to the specific intermolecular forces of ethanol.

Line Q is the weakening of the IMF, including hydrogen bonds, dipole-dipole forces, and London dispersion forces. Line S is breaking all IMF which requires more energy than just weakening the forces.

459. Which line segment represents …

- a. the enthalpy of fusion?

  Q

- b. the enthalpy of vaporization?

  S

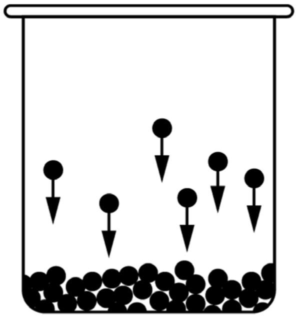
460. Explain what is happening on the molecular level at points P, R, and T.

The particles are speeding up.

461. Is the specific heat capacity of gaseous ethanol greater than, less than, or equal to the specific heat capacity of liquid ethanol? Justify your answer by referring to the slope of line segment R and line segment T.

The specific heat capacity of gaseous ethanol is less than the specific heat capacity of liquid ethanol. You can tell by the slope of the line. Since T is more steep it heats up faster over the same period of time which means it has a lower specific heat capacity.

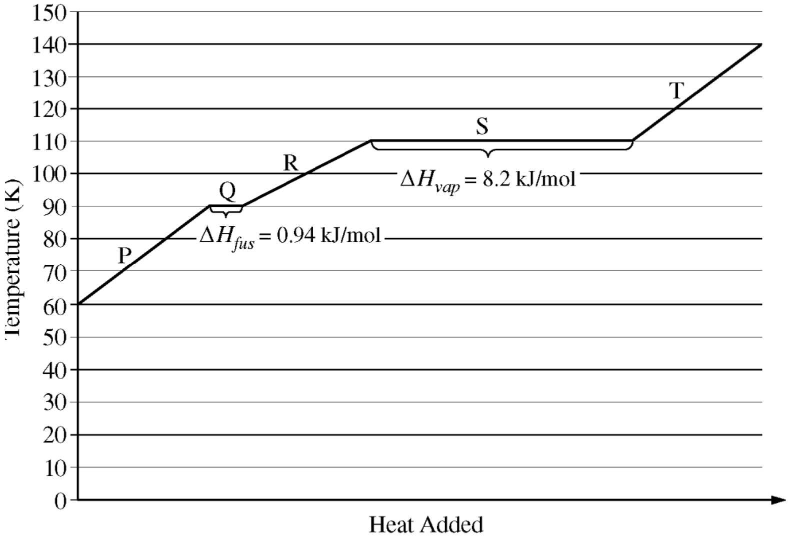
462. Three moles of a gas undergoes a phase transition, as shown below. How much energy is absorbed or released if the enthalpy of vaporization is +25 kJ/mol?



Energy is released because forces of attraction are being made.

$$\frac{3 \text{ mols}}{} \Big| \frac{25 \text{ kJ}}{1 \text{ mol}} = 75 \text{ kJ}$$

Use the heating curve of methane, CH_4, shown below to answer questions 463 to 468 below.



463.How much energy does it take to melt 1 mole of methane? Is this process exothermic or endothermic?

.94 kJ endothermic

464.How much energy does it take to vaporize 1 mole of methane? Is this process exothermic or endothermic?

8.2 kJ endothermic

465.How much energy does it take to freeze 1 mole of methane? Is this process exothermic or endothermic?

.94 kJ exothermic

466.How much energy does it take to condense 1 mole of methane? Is this process exothermic or endothermic?

8.2 kJ exothermic

467.How much energy does it take to freeze 9.0 g of methane?

$$\frac{9.0\,g\ CH_4 \mid 1\,mol}{\mid 16.042\,g} = \frac{.561\,mol\ CH_4 \mid .94\,kJ}{\mid 1\,mol} = -.527\,kJ$$

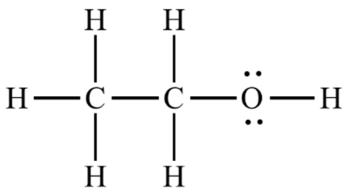
468.How much energy is released when 45 g of methane condenses?

$$\frac{45\,g\ CH_4 \mid 1\,mol}{\mid 16.042\,g} = \frac{2.81\,mol\ CH_4 \mid 8.2\,kJ}{\mid mol} = -23\,kJ$$

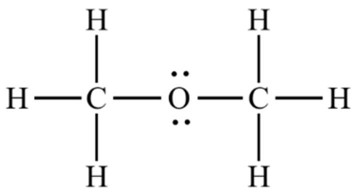
469.Complete the chart below about intermolecular forces.

Compound	All intermolecular and interparticle forces present	Strongest IMF present
Nonpolar substance	London dispersion	London dispersion
Polar substance, no H present	London dispersion dipole-dipole	dipole-dipole
Polar substance, H bonded to N, O, or F	London dispersion dipole-dipole H-bonding	H-bonding
Ionic substance in water	ion-dipole	ion-dipole
Nonpolar compound next to polar compound	dipole-induced dipole London dispersion forces	dipole-induced dipole

470.Consider the two compounds shown below.



Ethanol
Boiling point = 78.4 ℃



Dimethyl Ether
Boiling point = − 24.0 ℃

a. Which substance has the stronger intermolecular forces? Use data to support your answer.

Ethanol has the stronger IMF. Ethanol has a higher boiling point (78.4^{o}C) compared to dimethyl ether (-24.0^{o}C). A higher boiling point means more energy is required to break the IMF. More energy means stronger IMF.

b. What intermolecular forces does ethanol have?

London dispersion, dipole-dipole, hydrogen bonding

c. What intermolecular forces does dimethyl ether have?

London dispersion, dipole-dipole

d. What intermolecular force is responsible for the difference in boiling point? Explain your reasoning.

Hydrogen bonding. The molecules are isomers of each other. As a result, the only difference is the hydrogen bonding that is present in ethanol. Hydrogen bonding is the strongest IMF and is the reason for the difference.

Topic 6.7 Worksheet

471. Explain why a carbon-carbon triple bond requires more energy to break than a carbon-carbon single bond.

A carbon-carbon triple bond is stronger than a carbon-carbon single bond and therefore requires more energy to break. A carbon-carbon triple bond is a shorter bond which causes an increase in the Coulombic forces of attraction.

472. Is a carbon-carbon single bond longer than, shorter than, or equal in length to a carbon-carbon double bond?

A carbon-carbon single bond is longer than a carbon-carbon double bond.

473. Define bond dissociation energy, also known as bond enthalpy.

The energy required to break a bond.

474. Why are bond dissociation energies positive?

Bond breaking is endothermic. Energy must be absorbed to break a bond.

475. A certain bond has a bond enthalpy of 125 kJ/mol.
    a. How much energy is required to break the bond?

    125 kJ

    b. Is bond breaking endothermic or exothermic?

    Endothermic

476. A reaction is exothermic. What can be said about the sum of the bond enthalpies of the reactants compared to the sum of the bond enthalpies of the products?

If the reaction is exothermic then the change in enthalpy (dH) is negative. That means the bond enthalpy of the products (making bonds) is greater than the bond enthalpy of the reactants (breaking bonds).

477. In the chemical reaction below, it is known that A−B has a bond enthalpy of 85 kJ/mol. Determine the bond enthalpy of A_2.

$$A_2 + B \rightarrow ABA \quad \Delta H_{rxn} = -150 \text{ kJ/mol}_{rxn}$$

reactants	products
A_2	$A\text{-}B \times 2 = 85 \times 2$

Σ reactant − Σ prod = ΔH

$X - 170 = -150$

$X = 20 \frac{kJ}{mol}$

The bond enthalpy of A_2 is 20 kJ/mol

478. Predict the sign of ΔH for the following processes:

a. $2\,H(g) \rightarrow H_2(g)$ $\Delta H = -$

b. $Na(g) \rightarrow Na^+(g) + e^-$ $\Delta H = +$

c. $HBr(g) \rightarrow H(g) + Br(g)$ $\Delta H = +$

Bond Enthalpy Values

Bond	ΔH (kJ/mol)	Bond	ΔH (kJ/mol)	Bond	ΔH (kJ/mol)	Bond	ΔH (kJ/mol)
C−H	413	N−H	391	O−H	463	F−F	155
C−C	348	N−N	163	O−O	146		
C=C	614	N−O	201	O=O	495	Cl−F	253
C−N	293	N−F	272	O−F	190	Cl−Cl	242
C−O	358	N−Cl	200	O−Cl	203		
C=O	799	N−Br	243	O−I	234	Br−F	237
C−F	485					Br−Cl	218
C−Cl	328	H−H	436			Br−Br	193
C−Br	276	H−F	567				
C−I	240	H−Cl	431			I−Cl	208
		H−Br	366			I−Br	175
		H−I	299			I−I	151

479. Explain why the bond enthalpy of C−C is less than the bond enthalpy of C=C.

A C-C is weaker than a C=C. A weaker bond takes less energy to break.

480. Explain why the bond enthalpy of H–Cl is greater than the bond enthalpy of H–Br using principles of atomic structure.

The Cl is smaller than the Br. Since it's smaller there is a greater force of attraction between H-Cl than there is between H-Br

481. Determine ΔH_{rxn} for the reactions given below.

a. $\mathrm{H{-}H}\,(g) + \mathrm{Br{-}Br}\,(g) \longrightarrow 2\ \mathrm{H{-}Br}\,(g)$

rct	prod
H-H = 436	H-Br x2 = 366x2
Br-Br = 193	

629 — 732 = −103 kJ/mol rxn

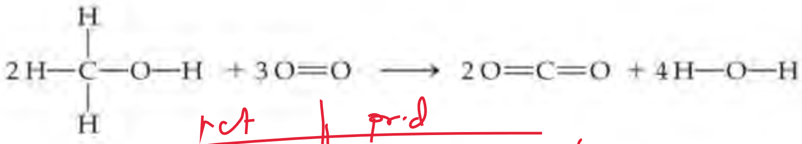
b. $\mathrm{H{-}CH_3}\,(g) + \mathrm{Cl{-}Cl}\,(g) \longrightarrow \mathrm{Cl{-}CH_3}\,(g) + \mathrm{H{-}Cl}\,(g)$

rct	prod
H-C: 413	C-Cl = 328
Cl-Cl = 242	H-Cl = 431

655 — 759 = −104 kJ/mol rxn

482. Use the bond enthalpy values to estimate ΔH_{rxn} for the reactions given below:

a.



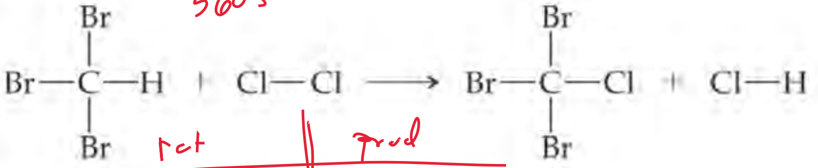
$$2\,\mathrm{H{-}\overset{\displaystyle H}{\underset{\displaystyle H}{C}}{-}O{-}H} + 3\,\mathrm{O{=}O} \longrightarrow 2\,\mathrm{O{=}C{=}O} + 4\,\mathrm{H{-}O{-}H}$$

rct	prod
6 x C-H = 413x6	4xC=O = 799x4
2 xC-O = 358x2	8x O-H = 463x8
2 xO-H = 463x2	
3 xO=O = 495x3	

5605 − 6900 = −1295

ΔH: −1295 kJ/mol$_{rxn}$

b.



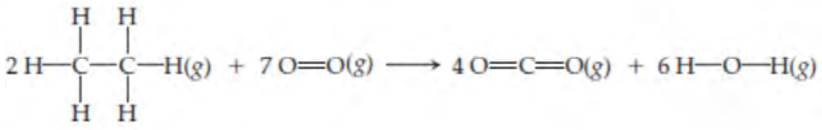
$$\mathrm{Br{-}\overset{\displaystyle Br}{\underset{\displaystyle Br}{C}}{-}H} + \mathrm{Cl{-}Cl} \longrightarrow \mathrm{Br{-}\overset{\displaystyle Br}{\underset{\displaystyle Br}{C}}{-}Cl} + \mathrm{Cl{-}H}$$

rct	prod
3 xC-Br = 276x3	3xC-Br = 276x3
C-H = 413	C-Cl = 328
Cl-Cl = 242	H-Cl = 431

1483 − 1587 = − 104 kJ/mol$_{rxn}$

ΔH = −

483. Determine the bond enthalpy of the O=O for the reaction below if $\Delta H_{rxn} = -2800$ kJ/mol$_{rxn}$.



$$2\,\mathrm{H{-}\overset{\displaystyle H}{\underset{\displaystyle H}{C}}{-}\overset{\displaystyle H}{\underset{\displaystyle H}{C}}{-}H}(g) + 7\,\mathrm{O{=}O}(g) \longrightarrow 4\,\mathrm{O{=}C{=}O}(g) + 6\,\mathrm{H{-}O{-}H}(g)$$

rct	prod
12 xC-H = 413x12	8 xC=O = 799x8
2xC-C = 348x2	12xO-H = 463x12
7xO=O = 7x	

5652 + 7x − 11948 = −2800

x = 499 kJ/mol

484. Two nitrogen atoms combine to form a nitrogen molecule, as represented by the following equation.

$$2\ N(g) \rightarrow N_2(g)$$

Using the table of average bond energies below, determine the enthalpy change, ΔH, for the reaction.

Bond	Average Bond Energy (kJ mol^{-1})
N — N	160
N ═ N	420
N ≡ N	950

rct	prod
0	N≡N = 950

0 − 950

$\Delta H = -950\ kJ/mol_{rxn}$

Review Topic 3.12

485. A photon has a wavelength of 317 nm.
   a. Determine the frequency of the photon.

$C = \lambda W$

3×10^{17} nm/s = 317 nm W

9.46×10^{14} Hz = W

   b. Determine the energy of the photon.

$$E = \frac{hc}{\lambda} = \frac{(6.626\times10^{-34})(3\times10^{17})}{317}$$

$$E = 6.27\times10^{-19}\ \text{J}$$

   c. How much energy would 1 mole of photons have?

$$\frac{6.27\times10^{-19}\ \text{J}}{\text{photon}}\ \Bigg|\ \frac{6.022\times10^{23}\ \text{photon}}{1\ \text{mole}} = 378\ \text{kJ/mol}$$

486. To break a mole of C−C single bonds 348 kJ of energy must be absorbed. Determine the lowest frequency of light that can be used to break a single C−C bond.

$$\frac{348\ \text{kJ}}{\text{mol}}\ \Bigg|\ \frac{1\ \text{mol}}{6.022\times10^{23}} = 5.78\times10^{-19}\ \text{J}$$

$E = h\nu$

$$\frac{E}{h} = \nu \qquad \frac{5.78\times10^{-19}}{6.626\times10^{-34}} = \nu = 8.72\times10^{14}\ \text{Hz}$$

487. Use the equation below to answer the questions that follow.

$$2\,H_2S(g) + 3\,O_2(g) \rightarrow 2\,H_2O(l) + 2\,SO_2(g) \qquad \Delta H^{\circ} = -1120\ \text{kJ/mol}_{rxn}$$

a. Determine the amount of energy released when 4.5 moles of $O_2(g)$ is used in the reaction.

$$4.5\ \text{mls}\ O_2\ \frac{1\ \text{mol rxn}}{3 O_2} = 1.5\ \text{mol rxn}\ \frac{-1120\ \text{kJ}}{\text{mol}_{rxn}} = -1680\ \text{kJ}$$

b. Determine the amount of energy released when 18 g of $H_2O(l)$ are formed. (MM_{H2O} = 18 g/mol)

$$18\ \text{g}\ H_2O\ \frac{1\ \text{mol}}{18\ \text{g}} = 1\ \text{ml}\ H_2O\ \frac{1\ \text{ml rxn}}{2\ \text{mol}\ H_2O} = .5\ \text{mol}_{rxn}\ \frac{-1120\ \text{kJ}}{\text{mol}_{rxn}} = -560\ \text{kJ}$$

c. Determine the amount of energy released when 1.45 grams of $SO_2(g)$ are formed.

$$1.45\ \text{g}\ SO_2\ \frac{1\ \text{mol}}{64.07\ \text{g}} = 0.0226\ \text{mol}\ SO_2\ \frac{1\ \text{mol rxn}}{2SO} = .0113\ \text{mols}_{rxn}\ \frac{-1120\ \text{kJ}}{\text{mol}_{rxn}} = 12.7\ \text{kJ}$$

d. Determine the mass of $H_2S(g)$ used when 120 kJ of energy are released.

$$-120\ \text{kJ}\ \frac{1\ \text{mol rtn}}{-1120\ \text{kJ}} = .107\ \text{mol}_{rxn}\ \frac{2\ H_2S}{1\ \text{mol}_{rxn}} = 0.214\ \text{mols}\ H_2S\ \frac{34.086\ \text{g}}{1\ \text{mol}} = 7.30\ \text{g}\ H_2S$$

water absorbs 12.7kJ

e. For the reaction in c., the energy released is used to heat 250 g of water (C = 4.2 J/g °C). If the water is initially at a temperature of 25 °C, determine the final temperature of the water. Assume no energy is lost.

$q = mC\Delta t$

$12\,700 = (250)(4.2)\Delta T$

$12.1°C = \Delta T$

$\Delta t = t_f - t_i$

$12.1 = T_f - 25$

$37.1°C = T_f$

f. Determine the mass of H_2S*(g)* required to react in order to bring 100. mL of water to a boil. The initial temperature of the water (d = 1.00 g/mL) is 20.0 °C and the specific heat capacity of the water is 4.18 J/g °C.

$q = mc\Delta t$

$= (100)(4.18)(100 - 20)$

$q = 33400 J$

$$-33.4 kJ \times \frac{1 \cancel{mol_{rxn}}}{-1120 kJ} \times \frac{2 \cancel{H_2S}}{1 \cancel{mol_{rxn}}} \times \frac{34.076 g}{1 mol} = 2.03 g\ H_2S$$

g. Determine the moles of SO_2*(g)* formed from the reaction above if the energy from the reaction is used to melt 120. g of ice initially at 0 °C. The heat of fusion of ice is 334 J/g.

$q = m\Delta H_f$

$= (120)(334 J/g)$

$q = 40\,080 J$

$$-40.08 kJ \times \frac{1 \cancel{mol_{rxn}}}{-1120 kJ} \times \frac{2\ SO_2}{1 \cancel{mol_{rxn}}} = .0716\ mols\ SO_2$$

488. Use the equation below to answer the questions that follow.

$$\mathrm{HCl}(aq) + \mathrm{NaOH}(aq) \rightarrow \mathrm{NaCl}(aq) + \mathrm{H_2O}(l) \qquad \Delta H^{\circ} = -57.1\ \mathrm{kJ/mol}_{rxn}$$

a. Determine the amount of energy released when 20.0 mL of 0.75 M HCl is used.

$n = .02\text{ L} \times .75\text{ M}$

$n = .015\text{ mls HCl}$

$$\frac{.015\text{ mols HCl}}{}\left|\frac{1\text{ mol rxn}}{1\text{ mol HCl}}\right|\frac{-57.1\text{ kJ}}{1\text{ mol}_{rxn}} = -.857\text{ kJ}$$

b. When 100. mL of 1.0 M HCl and 100. mL of 1.0 M NaOH are reacted, 5.71 kJ of energy is released. When 100. mL of 2.0 M HCl is reacted with 100. mL of 1.0 M NaOH the energy released is still 5.71 kJ. Explain why.

The NaOH is the limiting reactant. Once the NaOH runs out no more energy is given off.

c. When 25.0 mL of an unknown molarity of HCl is reacted with excess NaOH, 2.55 kJ of energy is released. Determine the molarity of the HCl solution.

$$\frac{-2.55\text{ kJ}}{}\left|\frac{1\text{ mol rxn}}{-57.1\text{ kJ}}\right|\frac{1\text{ HCl}}{1\text{ mol}_{rxn}} = .0447\text{ mol HCl}$$

$$M = \frac{n}{L} = \frac{.0447}{.025} = 1.79\text{ M}$$

489. Explain why Argon has a smaller radius than Phosphorus.

Argon has a greater nuclear charge (18 protons) than Phosphorus (15 protons) that attracts the electrons with a greater force. Both have the same energy shells and shielding electrons.

490. Explain why Xenon has a larger radius than Argon.

Xenon is further down the periodic table than Argon and therefore has more energy shells with electrons further from the nucleus. Argon has 3 energy shells while Xenon has 5 energy shells.

491. Explain why it takes more energy to remove the second electron from aluminum than it does to remove the first electron from aluminum.

Once the first electron from aluminum is removed there is less electron-electron repulsion which makes the Al^+ ion smaller than the Al atom. Since the Al^+ ion is smaller it takes more energy to remove the next electron because of Coulomb's Law.

Element	First Ionization Energy (kJ/mol)	Atomic Radius (pm)
C	1086	77
N	1400	75
O	1314	73
F	1680	72

492. The table above shows the first ionization energy and atomic radius of four elements in the same period. Explain why oxygen deviates in the trend for first ionization energy.

Oxygen has a p^4 with a set of electrons pairing up. Those paired electrons create more electron-electron repulsion and forces the electrons further away from the nucleus than expected. When the electron is further from than nucleus than expected the energy to remove that electron is less than expected.

493. The fluoride ion, F^-, is smaller than the oxide ion, O^{2-} even though they are isoelectronic. Explain why.

Both F^- and O^{2-} have the same number of electrons. However, F^- has a greater nuclear charge because of more protons (9 versus 8). The greater nuclear charge attracts the electrons with more force making the ion smaller.

1st I.E.	2nd I.E.	3rd I.E.	4th I.E.	5th I.E.
737.7	1,450.7	7,732.7	10,542.5	13,630

494. The table above shows the successive ionization energies of a third period element.

a. Explain why the 5th ionization is greater than the 4th ionization energy.

With each electron removed the ion gets smaller. The 5th I.E. is removing an electron from a smaller particle than the 4th I.E.

b. Explain why the 3rd ionization energy is significantly greater than the 2nd ionization energy.

The 3rd I.E. is an electron being removed from an inner shell of electrons significantly closer to the nucleus than the 1st and 2nd electrons.

c. What is the name of the element? Justify your answer by referring to the given ionization energies.

The element is magnesium. The third electron is removed from an inner shell of electrons. That means that the 1st and 2nd electron are the valence electrons. We know that it is a 3rd period element and the 3rd period element with two valence electrons is magnesium.

Section 6.06

Topic 6.1 Worksheet

495. Determine if the following observations would be from an **endothermic** or **exothermic** reaction. Then determine if energy is flowing into the reaction out of the reaction.

a. A student conducts a chemical reaction in a test tube and feels that the test tube is hot.

☐ Endothermic ☐ Energy flows into the system
☒ Exothermic ☒ Energy flows out of the system

b. A student mixes two chemicals in a beaker. Frost appears on the outside of the beaker.

☒ Endothermic ☒ Energy flows into the system
☐ Exothermic ☐ Energy flows out of the system

c. The temperature on a thermometer drops when placed into a test tube that contains a reaction.

☒ Endothermic ☒ Energy flows into the system
☐ Exothermic ☐ Energy flows out of the system

496. Mike Eilerman is a peach farmer. If peaches freeze they are ruined. When temperatures drop below freezing peach farmers will often spray their crop with water that will then freeze.

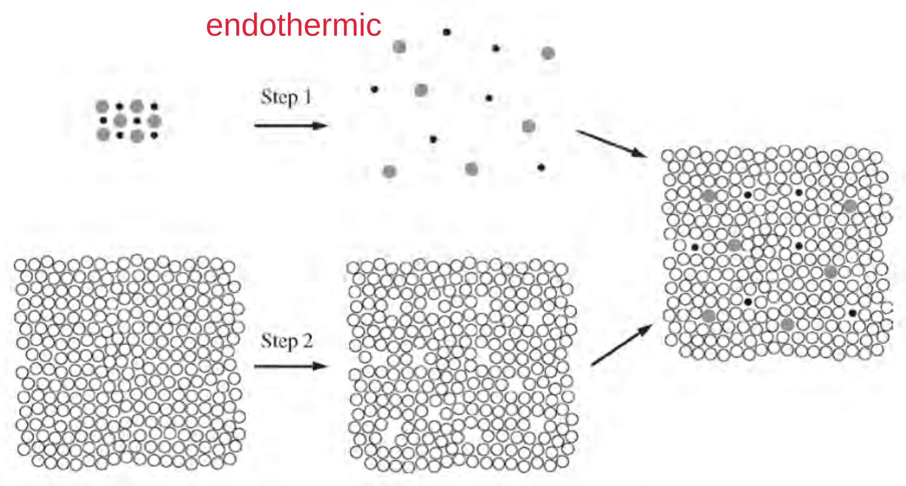
a. Is water freezing endothermic or exothermic? Explain why by referring to intermolecular forces.

Water freezing is exothermic. Freezing involves making IMF. Creating bonds/forces of attraction is exothermic.

b. Explain how spraying water on peaches keeps the peaches from freezing.

When the water freezes it gives off energy. That energy is absorbed by the peach keeping the temperature of the peack above freezing.

497. Shown below are the steps for an ionic solute dissolving into a polar solvent. In Step 1, the ions in the solute separate from each other, in Step 2 the polar solvent is expanded to make space for the solute, and then in Step 3 (unlabeled) the solute combines with the solvent. Determine if each step is exothermic or endothermic.



endothermic

exothermic

endothermic

498. If the formation of a solution is endothermic …

a. Which would require more energy, separating the ions in the crystal lattice or forming the solution?

Separating the ions

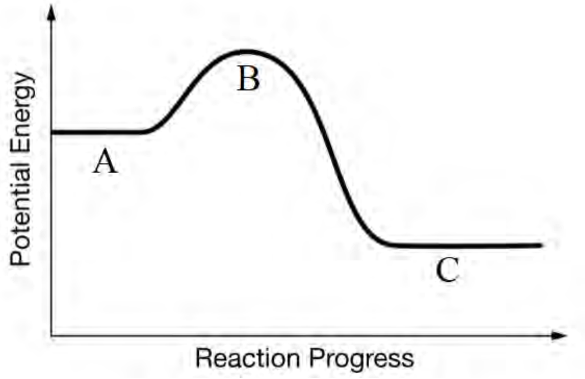
b. Explain why your answer to 4a would lead to an endothermic solution.

If it takes more energy to separate the ions than to form the solution then the value of dH for step 1 would be larger than for step 2.

499. Determine if the following processes are endothermic or exothermic by placing a + or – in the box.

Process	+ or –
$K(s) \rightarrow K(g)$	+
$K(g) \rightarrow K^+(g) + e^-$	+
$Cl_2(g) \rightarrow 2\ Cl(g)$	+
$Cl(g) + e^- \rightarrow Cl^-(g)$	-
$K^+(g) + Cl^-(g) \rightarrow KCl(s)$	-

500.Shown below is a potential energy diagram for an uncatalyzed reaction.



a. According to the diagram, is the reaction exothermic or endothermic? Justify your answer.

The reaction is exothermic. The products (C) are lower in energy than the reactants (A). Therefore energy must be released.

b. A solid, insoluble catalyst is added to the reaction mixture.

  i. Which part or parts of the potential energy diagram will change, A, B, or C?

  Only B.

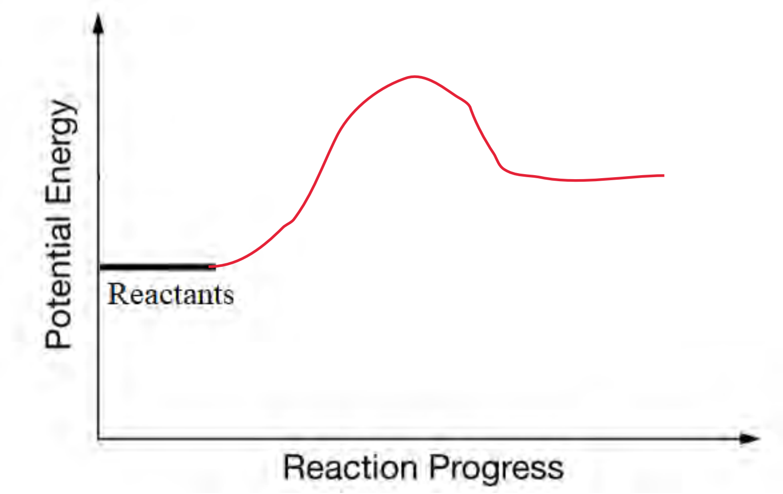
  ii. Describe how the part of the potential energy diagram will change, i.e. will the part be higher, lower, to the left, etc.?

  Part B will have a lower peak.

  iii. How will the catalyst change the quantity of product formed?

  A catalyst does not change the amount of product formed only how fast it is formed.

501. In 500.a above you determined whether the reaction was endothermic or exothermic. On the incomplete potential energy diagram below, draw the opposite type of reaction.



502. Determine the number of nitrogen atoms in 250. mL of a 0.95 M solution of $Ca(NO_3)_2$. Assume no other source of nitrogen in the solution.

$M \times L = n$

$(.95)(.25) = n$

$$\begin{array}{l|l} .2375 \text{ mols } Ca(NO_3)_2 & 2N \\ \hline & 1\ Ca(NO_3)_2 \end{array} = \begin{array}{l|l} 0.475 \text{ mls } N & 6.022 \times 10^{23} \text{ atm} \\ \hline & 1 \text{ mol} \end{array} = 7.86 \times 10^{23} \text{ atoms } N$$

503. Perform the following calculations. (Hint: You can use fractions to estimate your answer first.)

a. The number of atoms in 0.4 moles of Ca.

$$\begin{array}{l|l} .4 \text{ mols } Ca & 6.022 \times 10^{23} \text{ atoms} \\ \hline & 1 \text{ mol} \end{array} = 2.41 \times 10^{23} \text{ atoms } Ca$$

b. The number of moles in 18.5 grams of $Ca(OH)_2$ (molar mass = 74 g/mol).

$$\begin{array}{l|l} 18.5\text{g } Ca(OH)_2 & 1 \text{ mol} \\ \hline & 74\text{g} \end{array} = 0.25 \text{ moles } Ca(OH)_2$$

c. Determine the molarity of a solution made by dissolving 62.5 grams of a compound (molar mass = 125 g/mol) into 250 mL total solution.

$$\begin{array}{l|l} 62.5\text{g} & 1 \text{ mol} \\ \hline & 125\text{g} \end{array} = 0.5 \text{ mols}$$

$M = \frac{n}{L}$

$= \frac{.5}{.25} = 2.00 \text{ M}$

Compound	Molar Mass (grams)
NaCl	58.4
LiCl	42.4
CuCl	98.9
CsCl	168

504. According to the information provided in the table above, a 1.00 g sample of which compound would contain the greatest mass of chlorine? Justify your answer with calculations. Then, make a general statement about the percent of chloride and molar mass.

$\underline{NaCl}$ $\frac{35.45}{58.44} = .607$

$LiCl$ $\frac{35.45}{42.4} = 0.836$

$CuCl$ $\frac{35.45}{98.9} = 0.358$

$CsCl$ $\frac{35.45}{168} = .211$

LiCl would contain the most chloride because it has the greater % of chloride by mass. Since all of these substances contain one Cl, the percent of Cl decreases as the molar mass increases.

505. A student adds an excess of HCl(aq) to 150.0 mL of a solution containing Ag^+ and precipitates the Ag^+ as silver(I) chloride, AgCl. The student collects the precipitate in a pre-weighed piece of filter paper. The precipitate and filter paper are rinsed with distilled water, dried, and weighed. The data from the experiment are given in the table below.

Initial mass of filter paper	0.851 g
Mass of filter paper + AgCl(s) ppt. after drying	2.311 g

a. Determine the moles of AgCl that precipitated.

$$\begin{array}{r} 2.311\,g \\ -\ 0.857\,g \\ \hline 1.46\,g\ AgCl \end{array}$$

$$\frac{1.46\,g\ AgCl}{} \left| \frac{1\,mol}{143.32\,g} \right. = 0.01019\ mols\ AgCl$$

b. Calculate the original molarity of Ag^+ in the original solution. Assume that all Ag^+ ions precipitated.

$$\frac{0.01019\ mols\ AgCl}{} \left| \frac{1\,Ag^+}{1\,AgCl} \right. = .01019\ mols\ Ag^+$$

$$M = \frac{n}{L} = \frac{.01019\ mols}{.150\ L} = 0.0679\ M\ Ag^+$$

Section 6.07

Topic 6.8 Worksheet

506. What is the definition of the standard enthalpy of formation?

The energy require to form 1 mole of a substance from its elements.

507. Write reactions for the standard enthalpy of formation of the following compounds from their constituent elements in their standard states.

a. $MgCO_3(s)$

$$Mg(s) + C(graphite) + 1\tfrac{1}{2}O_2(g) \longrightarrow MgCO_3(s)$$

b. $C_6H_{12}O_6(s)$

$$6C(graphite) + 6H_2(g) + 3O_2(g) \longrightarrow C_6H_{12}O_6(s)$$

c. $CO_2(g)$

$$C(graphite) + O_2(g) \longrightarrow CO_2(g)$$

d. $NH_3(g)$

$$\tfrac{1}{2}N_2(g) + 1\tfrac{1}{2}H_2(g) \longrightarrow NH_3(g)$$

Compound	Standard Enthalpy of Formation (kJ mol^{-1})	Compound	Standard Enthalpy of Formation (kJ mol^{-1})	Compound	Standard Enthalpy of Formation (kJ mol^{-1})
$C_2H_4(g)$	52.30	$H_2O(l)$	−285.83	$NO_2(g)$	33.84
$C_2H_6(g)$	−84.68	$H_2O(g)$	−136.10	$NO(g)$	90.37
$CO_2(g)$	−393.5	$HNO_3(g)$	−134.3	$N_2O_4(g)$	9.66

508. Use the standard enthalpies of formation above to calculate the ΔH_{rxn} for each reaction given below.

a. $C_2H_4(g) + H_2(g) \rightarrow C_2H_6(g)$

$\Delta H_{rxn} = \Sigma \Delta H_{f\,prod} - \Sigma \Delta H_{f}\,rct$

$\Delta H = C_2H_6 - (C_2H_4 + H_2)$

$= -84.68 - (52.30 + 0)$

$\Delta H = -136.98$ kJ/mol rxn

b. $3\,NO_2(g) + H_2O(l) \rightarrow 2\,HNO_3(g) + NO(g)$

$\Delta H_{rxn} = \Sigma \Delta H_{f\,prd} - \Sigma \Delta H_{f}\,rct$

$\Delta H = (2(HNO_3) + NO) - (3(NO_2) + H_2O)$

$= (2(-134.3) + 90.37) - (3(33.84) + {}^{-}285.83)$

$\Delta H = 6.08$ kJ/mol rxn

c. $N_2O_4(g) + 4\,H_2(g) \rightarrow N_2(g) + 4\,H_2O(g)$

$\Delta H_{rxn} = \Sigma \Delta H_{f\,prod} - \Sigma \Delta H_{f}\,rct$

$\Delta H = (4(H_2O) + N_2) - (N_2O_4 + 4(H_2))$

$= 4(-136.10) + 0) - (9.66 + 0)$

$\Delta H = -554$ kJ/mol rxn

509. Determine the standard enthalpy of formation of nitromethane, CH_3NO_2, using the equation below and the standard enthalpies of formation from the table above.

$$2\,CH_3NO_2(l) + {}^{3}/_{2}\,O_2(g) \rightarrow 2\,CO_2(g) + 3\,H_2O(l) + N_2(g) \qquad \Delta H^{o}_{rxn} = -1418 \text{ kJ/mol}_{rxn}$$

$\Delta H = \Sigma \Delta H_{f\,prod} - \Sigma \Delta H_{f}\,rct$

$\Delta H = (2(CO_2) + 3(H_2O) + N_2) - (2(CH_3NO_2) + 3/2(O_2))$

$-1418 = (2(-393.5) + 3(-285.83) + 0) - (2x + \frac{3}{2}(0))$

$-1418 = -1644.49 - 2x$

$x = -113$ kJ/mol $= \Delta H_f\, CH_3NO_2$

510. In an experiment, liquid heptane, $C_7H_{16}(l)$, is completely combusted to produce $CO_2(g)$ and $H_2O(l)$, as represented by the following equation.

$$C_7H_{16}(l) + 11\ O_2(g) \rightarrow 7\ CO_2(g) + 8\ H_2O(l)$$

The heat of combustion, ΔH^o_{comb}, for one mole of $C_7H_{16}(l)$ is -4.85×10^3 kJ. Determine the enthalpy of formation, ΔH^o_f, for $C_7H_{16}(l)$

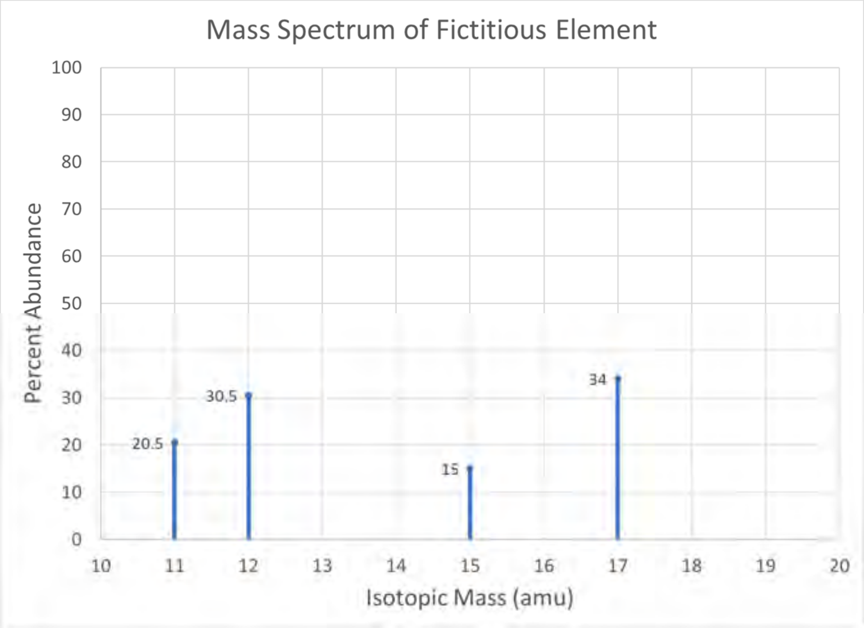
$$\Delta H = \left(7(CO_2) + 8(H_2O)\right) - \left(C_7H_{16} + 11(O_2)\right)$$

$$-4.85 \times 10^3 = \left(7(-393.5) + 8(-285.83)\right) - x$$

$$x = -191.14\ \frac{kJ}{mol} = \Delta H^\circ_f\ C_7H_{16}$$

Review Topic 1.2

511. Use the completely fictitious mass spectrum of an element below to answer the questions that follow.



a. How many isotopes does the fictitious element have?

4

b. What is the percent abundance of the most abundant isotope?

34

c. What is the atomic mass of the element?

$$\frac{(20.5 \times 11) + (30.5 \times 12) + (15 \times 15) + (34 \times 17)}{100} = \text{amu}$$

13.9 amu

512. A different fictitious element is made of the four isotopes shown below.

Mass of Isotope (amu)	Percent Abundance
80	10
82	32
84	35
??	??

a. What is the percent abundance of the missing isotope?

23 $(100 - 10 - 32 - 35)$

b. What is the mass of the missing isotope if the atomic mass of the element is 83.88 amu?

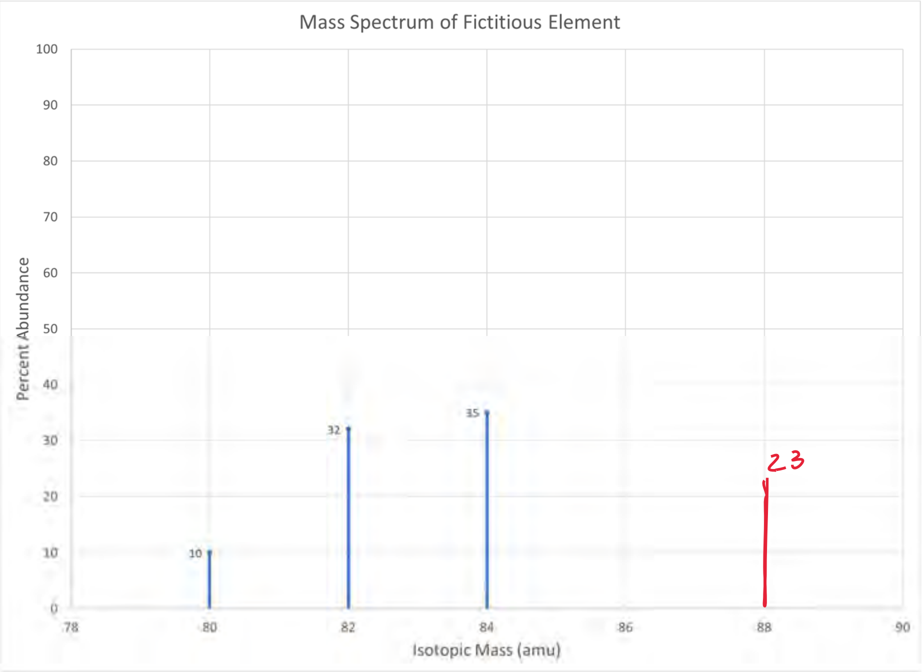
$$83.88 = \frac{(80 \times 10) + (82 \times 32) + (84 \times 35) + (x \times 23)}{100}$$

$$8388 = 800 + 2624 + 2940 + 23x$$

$$2024 = 23x$$

88 amu = x = mass of the missing isotope

c. On the graph below, sketch in the missing isotope.



Topic 6.9 Worksheet

513. A certain reaction has a ΔH of −25 kJ/mol$_{rxn}$. What would be the ΔH for the opposite reaction?

+ 25 kJ/mol

514. Is energy transferred to the system or the surroundings …

  a. in an endothermic reaction?

  To the system

  b. in an exothermic reaction?

  To the surroundings

515. The reaction below is the sublimation of iodine.

$$I_2(s) \rightarrow I_2(g)$$

  a. Would you expect the sublimation of iodine to be endothermic or exothermic? Justify your answer by referring to the bonds/forces of attraction.

  Endothermic. The London dispersion forces are being broken. Bond/Force breaking requires energy to be put into the system.

  b. The energy required to sublime 1 mole of iodine is the same amount of energy that is required to turn 1 mole of solid iodine to liquid iodine and then liquid iodine to gaseous iodine. Explain why by referring to what is occurring with the intermolecular forces.

  In both cases the same amount of IMF are being broken.

  c. Use the reactions given below to determine the ΔH°$_{rxn}$ in kJ/mol$_{rxn}$.

$\frac{1}{2} H_2(g) + \frac{1}{2} I_2(s) \rightarrow HI(g)$ ΔH = 26 kJ/mol$_{rxn}$

$\frac{1}{2} H_2(g) + \frac{1}{2} I_2(g) \rightarrow HI(g)$ ΔH = −5.0 kJ/mol$_{rxn}$

$\frac{1}{2}\cancel{H_2(g)} + \frac{1}{2} I_{2(s)} \rightarrow \cancel{HI(g)}$ $\Delta H = 26\ kJ/mol_{rxn}$

$\cancel{HI(g)} \rightarrow \frac{1}{2}\cancel{H_2(g)} + \frac{1}{2} I_{2(g)}$ $\Delta H = 5.0\ kJ/mol_{rxn}$

$\frac{1}{2} I_{2(s)} \rightarrow \frac{1}{2} I_{2(g)}$ $\Delta H = 31\ kJ/mol_{rxn}$

must multiply final reaction by 2

$2\times$ $\frac{1}{2} I_{2(s)} \rightarrow \frac{1}{2} I_{2(g)}$ $2\times$

$\Delta H = 31\ kJ/mol_{rxn}$

$I_{2(s)} \rightarrow I_{2(g)}$

$\Delta H = 62\ kJ/mol_{rxn}$

516. Calculate the standard enthalpy change, ΔH°, for the formation of 1 mol of strontium carbonate (the material that gives the red color in fireworks) from its elements.

$$Sr\ (s) + C\ (graphite) + 3/2\ O_2\ (g) \rightarrow SrCO_3\ (s)$$

The information available is

$Sr\ (s) + \frac{1}{2} O_2\ (g) \rightarrow SrO\ (s)$	ΔH° = −592 kJ
$SrO\ (s) + CO_2\ (g) \rightarrow SrCO_3\ (s)$	ΔH° = −234 kJ
$C\ (graphite) + O_2\ (g) \rightarrow CO_2\ (g)$	ΔH° = −394 kJ

$Sr(s) + \frac{1}{2} O_2(g) \longrightarrow \cancel{SrO}(s)$ ΔH = −592 kJ

$C(graphite) + O_2(g) \longrightarrow \cancel{CO_2}(g)$ ΔH = −394 kJ

$\cancel{SrO}(s) + \cancel{CO_2}(g) \longrightarrow SrCO_3(s)$ ΔH = −234 kJ

$Sr(s) + 3/2\ O_2(g) + C(graphite) \longrightarrow SrCO_3(s)$ ΔH = −1220 kJ

517. In designing a chemical plant for manufacturing the plastic polyethylene, you need to know the enthalpy change for the removal of H_2 from C_2H_6 (ethane) to give C_2H_4 (ethylene), a key step in the process.

$$C_2H_6\ (g) \rightarrow C_2H_4\ (g) + H_2\ (g) \qquad \Delta H^o = ?$$

From experiments you know the following thermochemical equations:

$2\ C_2H_6\ (g) + 7\ O_2\ (g) \rightarrow 4\ CO_2\ (g) + 6\ H_2O\ (l)$	ΔH° = −3119.4 kJ
$C_2H_4\ (g) + 3\ O_2\ (g) \rightarrow 2\ CO_2\ (g) + 2\ H_2O\ (l)$	ΔH° = −1410.9 kJ
$2\ H_2\ (g) + O_2\ (g) \rightarrow 2\ H_2O\ (l)$	ΔH° = −571.66 kJ

Use this information to find the value of ΔH° for the formation of ethylene from ethane.

÷2 $\frac{2}{2} C_2H_6(g) + \frac{\cancel{7O_2}}{2}(g) \longrightarrow \frac{\cancel{4CO_2}}{2}(g) + \frac{\cancel{6H_2O}}{2}(l)$ ΔH = −3119.4 kJ ÷2

$2\cancel{CO_2}(g) + 2\cancel{H_2O}(l) \longrightarrow C_2H_4(g) + \cancel{3O_2}(g)$ ΔH = +1410.9 kJ

ΔH = +571.66 kJ ÷2

÷2 $\frac{2}{2}\cancel{H_2O}(l) \longrightarrow \frac{2}{2} H_2(g) + \frac{1}{\cancel{2}}\cancel{O_2}(g)$

$C_2H_6(g) \longrightarrow H_2(g) + C_2H_4(g)$ ΔH = 137.03 kJ/mol rxn

518. The combustion of carbon monoxide is represented by the equation below.

$$CO(g) + \frac{1}{2} O_2(g) \rightarrow CO_2(g)$$

Determine the value of the standard enthalpy change, ΔH°_{rxn}, for the combustion of CO(g) at 298 K using the following information.

$C(s) + \frac{1}{2}O_2(g) \rightarrow CO(g)$ $\qquad \Delta H^\circ_{298} = -110.5 \text{ kJ mol}^{-1}$

$C(s) + O_2(g) \rightarrow CO_2(g)$ $\qquad \Delta H^\circ_{298} = -393.5 \text{ kJ mol}^{-1}$

$CO_{(g)} \rightarrow \cancel{C}_{(s)} + \cancel{\frac{1}{2}O_2}_{(g)}$ $\qquad \Delta H = +110.5 \text{ kJ/mol}$

$\cancel{C}_{(s)} + \frac{1}{2}O_{2(g)} \rightarrow CO_{2}(g)$ $\qquad \Delta H = -393.5$

$CO_{(g)} + \frac{1}{2}O_{2}(g) \rightarrow CO_{2(g)}$ $\qquad \Delta H = -283 \frac{\text{kJ}}{\text{mol}}$

519. Consider the two processes represented below.

Process 1: $H_2O(l) \rightarrow H_2O(g)$ $\qquad \Delta H^\circ = +44.0 \text{ kJ mol}^{-1}$

Process 2: $H_2O(l) \rightarrow H_2(g) + \frac{1}{2}O_2(g)$ $\qquad \Delta H^\circ = +286 \text{ kJ mol}^{-1}$

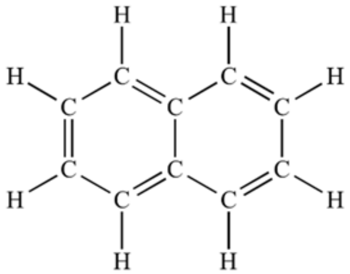
a. For each of the two processes, identify the type(s) of intermolecular or intramolecular attractive forces that must be overcome for the process to occur.

Process 1: Hydrogen bonds are being broken

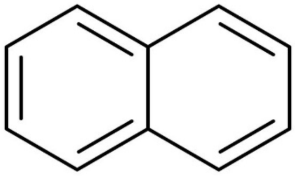
Process 2: Covalent bonds are being broken

b. Explain why Process 2 requires much more energy than Process 1.

Covalent bonds are much stronger than hydrogen bonds. A stronger bond requires more energy.



Lewis structure of naphthalene



Skeletral structure of naphthalene

520. The compound above is naphthalene and is commonly used as moth balls. Naphthalene has a melting point of 80.3 ℃ and a boiling point of 218 ℃ at 1 atm.

a. What state of matter, solid, liquid, or gas, is naphthalene at 60.0 ℃ and 1 atm?

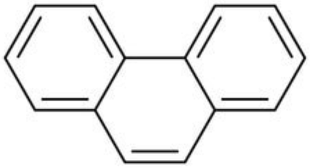
Naphthalene is a solid at 60.0^{o}C.

b. What intermolecular forces are present for naphthalene?

Naphthalene has London dispersion forces present.

c. In which state of matter are the intermolecular forces the greatest, solid, liquid, or gas? The weakest?

The IMF are the greatest in a solid and the weakest in a gas.



Skeletal structure of phenanthrene

d. The skeletal structure above is phenanthrene.

   i. Would you expect phenanthrene at 60.0 ℃ to have stronger, weaker, or the same strength of intermolecular forces as naphthalene at 60.0 ℃? Justify your answer.

I would expect phenanthrene to have stronger IMF than naphthalene. Phenanthrene has more electrons that naphthalene which makes it more polarizable with stronger London dispersion forces.

   ii. Which compound would you expect to have a higher melting point, phenanthrene or naphthalene? Explain your answer.

Phenanthrene would have a higher melting point. Since phenanthrene has stronger IMF than naphthalene it takes more energy to separate the particles from each other. More energy means a higher melting point.

   iii. Which intermolecular force, if any, is responsible for your answer to 520.d.ii?

London dispersion forces.

Unit 7

Section 7.01

Topic 7.1 Worksheet

521. A reaction begins with only reactants. As the reaction proceeds toward equilibrium …

a. the rate of the forward reaction increases [✓ decreases] until equilibrium is reached.

b. the rate of the reverse reaction [✓ increases] decreases until equilibrium is reached.

c. the concentration of the reactants will increase [✓ decrease] until equilibrium is reached.

d. the concentration of the products will [✓ increase] decrease until equilibrium is reached.

522. At equilibrium …

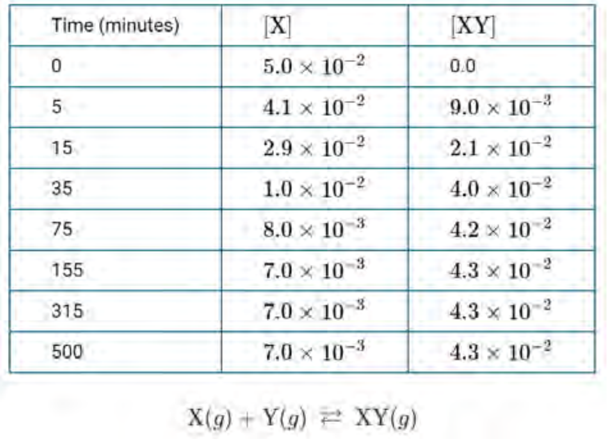
a. the rate of the forward reaction is greater than [✓ equal to] less than the rate of the reverse reaction.

b. the rate of the reverse reaction is greater than [✓ equal to] less than the rate of the forward reaction.

c. the concentration/pressure of the reactants is increasing. [✓ remains the same.] is decreasing.

d. the concentration/pressure of the products is increasing. [✓ remains the same.] is decreasing.

523. The reaction shown below is conducted and data collected.



Time (minutes)	[X]	[XY]
0	5.0×10^{-2}	0.0
5	4.1×10^{-2}	9.0×10^{-3}
15	2.9×10^{-2}	2.1×10^{-2}
35	1.0×10^{-2}	4.0×10^{-2}
75	8.0×10^{-3}	4.2×10^{-2}
155	7.0×10^{-3}	4.3×10^{-2}
315	7.0×10^{-3}	4.3×10^{-2}
500	7.0×10^{-3}	4.3×10^{-2}

$$X(g) + Y(g) \rightleftarrows XY(g)$$

a. How does the rate of the forward reaction compare to the rate of the reverse reaction from 0 to 5 minutes? Explain your reasoning?

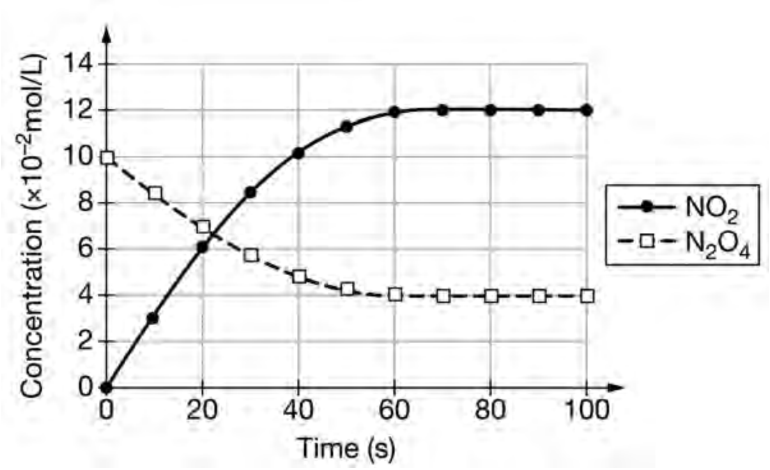
The rate of the forward reaction is very fast compared to the rate of the reverse reaction. The forward reaction causes a drop in concentration of X by 9 x 10^{-3}. The reverse reaction sees no change in concentration from 0 to 5 minutes as the [XY] changes from 0 to 9 x 10^{-3}. The reverse reaction isn't measurably occurring.

b. At what point does the system reach equilibrium? Justify your answer by referring to the data.

155 s. At that time the concentrations stop changing.

c. At equilibrium, what visible changes can be observed?

None. At equilibrium it appears as though the system has stopped but in reality both the forward and reverse reactions are proceeding.



524. The reaction $N_2O_4\,(g) \rightleftharpoons 2\,NO_2\,(g)$ occurs in an evacuated container at 373 K. The concentration of each species is measured over time and the data are used to make the graph shown above.

a. When does the system reach equilibrium? Justify your answer.

60 s. The concentration stops changing at 60 s.

b. At 60 s, is the rate of the forward reaction greater than, less than, or equal to the rate of the reverse reaction? Explain your reasoning.

The rate of the forward reaction is equal to the rate of the reverse reaction. At equilibrium the rates are equal.

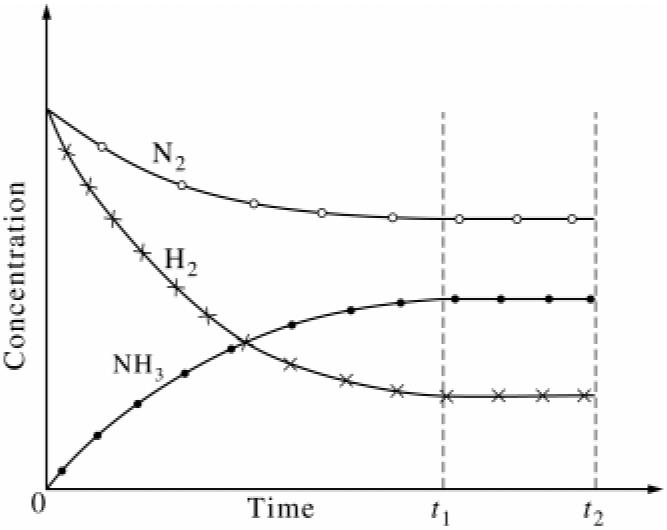
c. Explain why the slope of the [NO_2] is greater than the slope of the [N_2O_4].

You form more moles of NO_2 than you lose moles of N_2O_4 because of the stoichiometry of the reaction.

525. Two different containers at the same temperature have the same volume of a volatile organic compound. The liquid organic compound is in equilibrium with the vapor phase. Explain why, regardless of size or surface area, the equilibrium vapor pressure of both containers is the same.

The liquid will always achieve the same equilibrium vapor pressure at that temperature regardless of amount of liquid present (assuming enough liquid is present to reach equilibrium). The actual number of particles as a vapor may change but not the concentration/pressure.

526. Use the graph below to answer the questions that follow about the reaction $3\ H_2\ (g) + N_2\ (g) \rightleftharpoons 2\ NH_3\ (g)$



a. What is significant about t_1?

Equilibrium

b. Why does the $[H_2]$ decrease faster than the $[N_2]$?

For every 1 mole of N_2 consumed there are 3 moles of H_2 consumed.

c. What is the rate of disappearance of H_2 compared to the rate of disappearance of N_2?

$\frac{1}{3}\ \text{Rate}_{H_2} = \frac{1}{1}\ \text{Rate}_{N_2}$

d. How does the rate of disappearance of H_2 compare to the rate of appearance of NH_3?

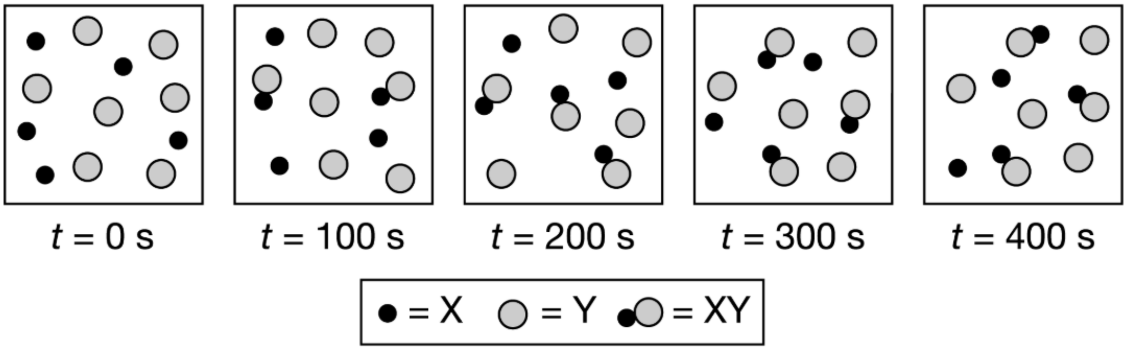
$$\frac{1}{3}\text{Rate}_{H_2} = \frac{1}{2}\text{Rate}_{NH_3}$$

e. At what point does the rate of the forward reaction equal the rate of the reverse reaction?

At point t_1, equilibrium.

Topic 7.2 Worksheet

527. Use the figures below to answer the questions that follow.



$X(g) + Y(g) \rightleftarrows XY(g)$

a. At what time is equilibrium established? Justify your answer.

200 s. The number of X, Y, and XY stops changing.

b. Would the value of K be greater than, less than, or equal to 1? Justify your answer.

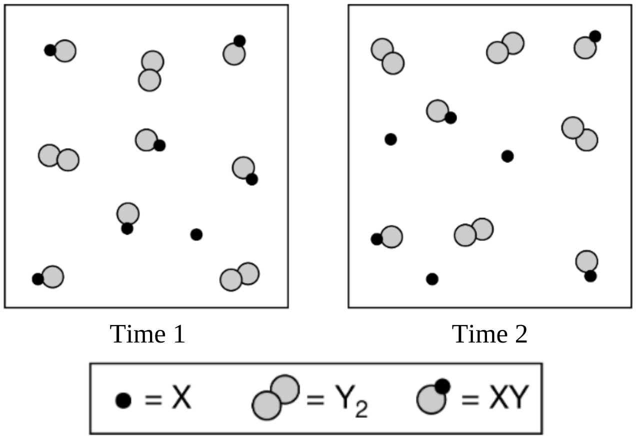
K would be less than 1.

$$K = \frac{[XY]}{[X][Y]} = \frac{3}{(2)(4)} = \frac{3}{8}$$

c. What does the value of K indicate about the amounts of reactants and products?

If K > 1 then more products are present.
If K < 1 then more reactants are present.
In this case K < 1 so more reactants are present than products.

528. Use the figures below to answer the questions that follow.



$2\,X(g) + Y_2(g) \rightleftarrows 2\,XY\,(g)$

a. Has the system reached equilibrium? Justify your answer.

No. The system is still changing; not enough data to determine.

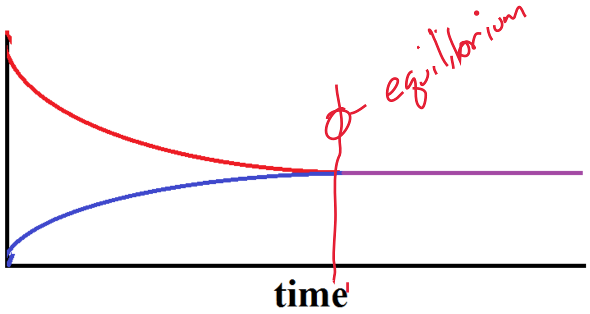
b. What can be inferred about the relative rates of the forward and reverse reactions between time 1 and time 2?

The rate of the forward reaction is less than the rate of the reverse reaction.

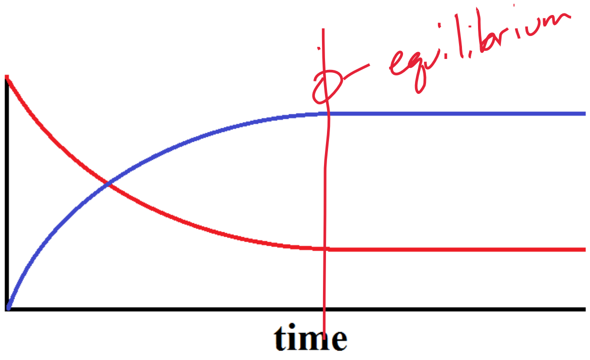
529. A student makes the claim shown in the box. Do you agree or disagree with the statement? Explain your reasoning.

> At equilibrium, the forward and reverse reaction stop occurring because the concentrations of reactants and products no longer changes.

I disagree. At equilibrium the reactants are still changing into products and vice versa. But the rate of the forward reaction is equal to the rate of the reverse reaction.



Graph A



Graph B

530. Of the two graphs above …

a. which shows how the concentration of a reaction changes over time to reach equilibrium? Justify your answer.

Graph B. The concentrations stop changing but they do NOT have to be equal.

Note: Graph A could be the concentration but graph B is more likely. Graph A would give a K of 1 at equilibrium.

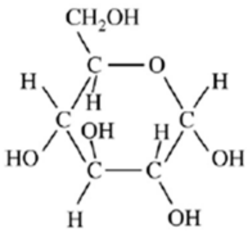
b. which shows how the rates of a reaction changes over time to reach equilibrium? Justify your answer.

Graph A. The rates will become equal as they reach equilibrium.

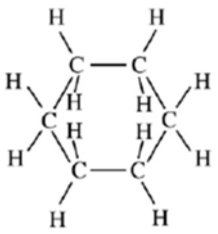
c. On both graphs, draw a vertical line to show where equilibrium is established.

531. What is the relationship between bond polarity and miscibility of a liquid in water?

As bond polarity increases the miscibility of a liquid in water increases.



Glucose



Cyclohexane

532. Determine all intermolecular forces present in glucose.

Glucose has London dispersion, dipole-dipole, and hydrogen bonding IMF.

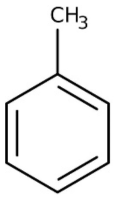
533. Determine all intermolecular forces present in cyclohexane.

Cyclohexane has London dispersion forces only.

534. Which substance, glucose or cyclohexane, would be more soluble in water? Explain your reasoning.

Glucose would be more soluble. Glucose has IMF more similar to water than cyclohexane.

535. Shown below is toluene. Which substance, glucose or cyclohexane, would be more soluble in toluene? Explain your reasoning.



Cyclohexane would be more soluble in toluene. Cyclohexane and toluene have similar IMF - both have just London dispersion forces.

Section 7.02

Topic 7.4, 7.5, 7.7 Worksheet

536. Answer the following questions based on the value of K given.

Value of K	Forward or Reverse reaction favored	Proceeds to Completion or Barely proceeds at all	More Reactants or More Products
1 x 10^{-3}	Forward [✓ Reverse]	Proceeds to Completion [✓ Barely Proceeds]	[✓ More Reactants] More Products
10	[✓ Forward] Reverse	Proceeds to Completion [✓ Barely Proceeds]	More Reactants [✓ More Products]
1 x 10^{15}	[✓ Forward] Reverse	[✓ Proceeds to Completion] Barely Proceeds	More Reactants [✓ More Products]
1	Forward, neither Reverse	Proceeds to Completion [✓ Barely Proceeds]	More Reactants neither More Products
1 x 10^{-16}	Forward [✓ Reverse]	Proceeds to Completion [✓ Barely Proceeds]	[✓ More Reactants] More Products

537. Use the reaction given below to answer the questions that follow.

$$2\,X\,(g) + Y\,(g) \rightleftharpoons 3\,Z\,(g)$$

a. Write the expression for the equilibrium constant, K_C, for the reaction.

$$K_C = \frac{[Z]^3}{[X]^2[Y]}$$

b. Determine the value of K_c if the molar concentrations at equilibrium are [X] = 2.0 M, [Y] = 0.5 M, [Z] = 4.0 M.

$$K_C = \frac{(4)^3}{(2)^2(.5)} = \frac{64}{2} = 32$$

c. Determine the partial pressure of each gas if the total pressure is 17 atm and the gases are in a 5.0 L container.

$X = M \times L = n = 5.0L \times 2M = 10mols$

$Y = M \times L = n = 5.0L \times .5M = 2.5mols$

$Z = M \times L = n = 5.0L \times 4M = 20mols$

$P_x = \chi_x P_{tot} = \left(\frac{10}{32.5}\right)(17)$

$P_x = 5.23atm$

$P_y = \chi_y P_{tot} = \left(\frac{2.5}{32.5}\right)(17)$

$P_y = 1.31 atm$

$P_z = \chi_z P_{tot} = \left(\frac{20}{32.5}\right)(17)$

$P_z = 10.46atm$

d. Determine the value of K_p.

$$K_p = \frac{(10.46)^3}{(5.23)^2(1.31)} = 32$$

Note: Same # of moles of gas on each side means $K_c = K_p$

$$C\ (s) + CO_2\ (g) \rightleftharpoons 2\ CO\ (g)$$

538. Solid carbon and carbon dioxide were placed in a rigid 2.00 L container and the reaction represented above occurred. As the reaction proceeded the total pressure in the container was monitored. When equilibrium was reached there was still some C(s) remaining in the mixture. Results are recorded in the table below.

Time (hours)	Total Pressure of Gases in Container at 1,160 K (atm)
0.0	5.00
2.0	6.26
4.0	7.09
6.0	7.75
8.0	8.37
10.0	8.37

a. Write the expression for the equilibrium constant, K_p, for the reaction.

$$K_p = \frac{(P_{CO})^2}{P_{CO_2}}$$

b. Calculate the number of moles of CO_2 (g) initially placed in the container. (Assume that the volume of the solid carbon is negligible.)

All pressure at time 0 is from CO_2

$$PV = nRT$$

$$(5)(2) = n(0.08206)(1160)$$

$$n = 0.105 \text{ moles}$$

c. At what time does the system reach equilibrium? Justify your answer.

8.0 hours   Pressure stops changing at 8.0 hours

d. For the reaction mixture at equilibrium at 1,160 K, the partial pressure of the CO_2 is 1.63 atm. Calculate …

i. the partial pressure of CO *(g)* and

$P_{tot} = P_{CO_2} + P_{CO}$ OR

$8.37 = 1.63 + P_{CO}$

$P_{CO} = 6.74$ atm

	I	C	E
CO	0	+2x	2x
CO_2	5	-x	5-x

$5 - x = 1.63$

$x = 3.37$

$2x = E$

$2(3.37) = E$

$6.74 = E$

ii. the value of the equilibrium constant, K_p

$$K_p = \frac{(6.74)^2}{1.63} = 27.9$$

e. According to the K value you calculated in d is the forward reaction or reverse reaction favored?

Forward Reaction. K > 1 means more product.

539. The same number of moles of HCl and O_2 are used in the reaction below. Answer the questions that following about the reaction *after* it has reached equilibrium.

$$4\,HCl\,(g) + O_2\,(g) \rightleftharpoons 2\,Cl_2\,(g) + 2\,H_2O\,(g)$$

a. How would the molarity of H_2O compare to the molarity of Cl_2? Explain why by referring to the reaction stoichiometry.

The molarities would be the same because they have the same coefficient.

b. How would the molarity of HCl compare to the molarity of O_2? Explain why by referring to the reaction stoichiometry.

The molarity of HCl would be less than the molarity of O_2. For every 1 mole of O_2 used there are 4 moles of HCl used.

540. The reaction below occurs when two aqueous solutions are mixed, forming an aqueous complex ion.

$$A^{2+}\ (aq) + B^{4-}\ (aq) \rightleftharpoons AB^{2-}\ (aq) \qquad K = 7.7 \times 10^7$$

a. The value of K is *very* large. Explain why this reaction can be considered to go to completion even though it is actually in equilibrium.

Since K is so large we have almost entirely products and very few reactants. So much so that the reaction proceeds almost to completion and can be treated as a typical stoichiometry problem.

b. Determine the concentration of AB^{2-} in the equilibrium mixture if 50.0 mL of 0.35 M B^{4-} is mixed with 50.0 mL of 0.30 M A^{2+} to produce 100.0 mL of total solution.

$$K_c = \frac{[AB^{2-}]}{[A^{2+}][B^{4-}]} = 7.7 \times 10^7$$

$$n_{B^{4-}} = M \times V = (.35)(.05) = \frac{.0175\ \text{mols}}{.1\ L} = .175\ M$$

$$n_{A^{2+}} = M \times V = (.30)(.05) = \frac{.015\ \text{mols}}{.1\ L} = .15\ M$$

using stoich, $n_{A^{2+}} = n_{AB^{2-}} = .015\ \text{mols}/.1\ L = .15\ M = [AB^{2-}]$

	I	C	E
B^{4-}	.175	$-x$	$.175-x$
A^{2+}	.15	$-x$	$.15-x$
AB^{2-}	0	$+x$	x

$$K_c = \frac{x}{(.15-x)(.175-x)}$$ (quadratic!)

$x = .1499\ M = [AB^{2-}]$

$$CO\ (g) + H_2O\ (g) \rightleftharpoons CO_2\ (g) + H_2(g) \qquad K_c = 1.5 \times 10^3$$

541. An equal number of moles of CO and H_2O are placed into a rigid reaction vessel and allowed to reach equilibrium. Answer the following questions about the concentration of each species at equilibrium.

a. Would the equilibrium concentration of CO be greater than, less than, or equal to the equilibrium concentration of H_2O? Explain your reasoning.

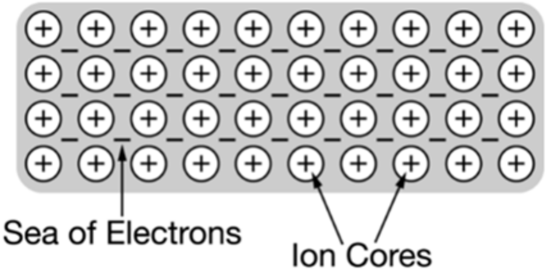
It would be equal. The reaction starts with equal moles of CO and H_2O. For every 1 mole of CO used there is 1 mole of H_2O used.

b. Would the equilibrium concentration of CO_2 be greater than, less than, or equal to the equilibrium concentration of H_2? Explain your reasoning.

The concentration would be equal. For every 1 mole of CO_2 formed 1 mole of H_2 is formed.

c. Would the equilibrium concentration of CO_2 be greater than, less than, or equal to the equilibrium concentration of H_2O? Explain your reasoning.

The concentration would be greater in CO_2. Since the reaction favors the products more CO_2 will form compared to H_2O.



542. Use the figure above to help describe the following:

   a. Sea of electrons

The electrons are free to move about the entire structure and are not bound to a single atom.

   b. Nondirectional bonds

The electrons are shared between all of the atoms. As a result the metallic bonds are nondirectional.

543. What are the relative sizes of metallic atoms used to create an interstitial alloy?

A large atom and a small atom.

544. What are the relative sizes of metallic atoms used to create a substitutional alloy?

They are relatively the same size.

545. A metal is used to make two different alloys, an interstitial alloy with copper and a substitutional alloy with silver. It is determined that the interstitial alloy is harder than the substitutional alloy. Use the structural compositions of the two types of alloys to explain why the interstitial alloy is harder than the substitutional alloy.

The interstitial alloy has copper atoms between all of the other atoms of the metal. These copper atoms in between the other atoms makes the alloy less malleable as the copper atoms block the movement. In the substitutional alloy there are no atoms between the silver and other metal. This allows the atoms to move more and makes it more malleable.

546. Steel is an alloy made of iron, Fe, and carbon, C. Chromium, Cr, is often used to make steel less likely to rust. Fe and Cr have similar atomic radii while Fe and C have different atomic radii.

a. What type of alloy would Fe and Cr form?

Fe and Cr would make a substitutional alloy.

b. What type of alloy would Fe and C form?

Fe and C would form an interstitial alloy.

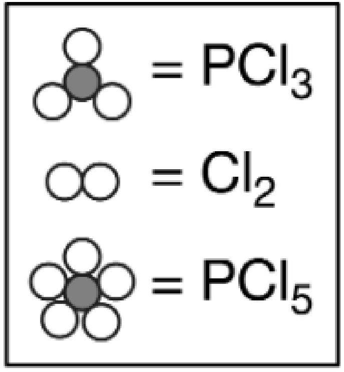
Section 7.03

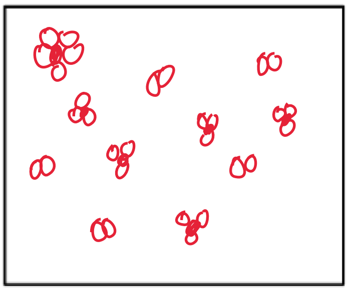
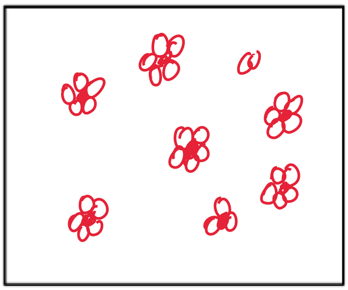
Topic 7.8 Worksheet

547. The reaction shown below is carried out at various temperatures and the value of K_C determined. Draw particulate representations that show the relative amounts of each reactant and product that would **roughly** correspond to the value of K.

$$PCl_3\,(g) + Cl_2\,(g) \rightleftharpoons PCl_5\,(g)$$

Legend



Magnitude of K	1×10^{-3}	1×10^{5}
Particulate diagram of equilibrium mixture		

Review Topic 3.7

548. Determine the molarity of all of the ions for each of the solutions given below. Assume solutions are made of strong electrolytes.

a. 14.9 grams of KCl (molar mass = 74.5 g/mol) in 133.3 mL of total solution

$\frac{14.9\text{g KCl} \mid 1\text{mol}}{\mid 74.5\text{g}} = 0.2\text{mols KCl}$ $M = \frac{n}{L} = \frac{.2\text{moles}}{.1333\text{L}} = 1.5\text{M KCl}$

1.5M K^+

1.5M Cl^-

b. 63.75 grams of $Ba(C_2H_3O_2)_2$ (molar mass = 255 g/mol) in 0.1667 L of total solution

$\frac{63.75\text{g } Ba(C_2H_3O_2)_2 \mid 1\text{mol}}{\mid 255\text{g}} = 0.25\text{mols}$

$M = \frac{n}{L}$

$= \frac{.25}{.1667} = 1.5\text{M } Ba(C_2H_3O_2)_2$

$1.5\text{M } Ba^{2+}$

$3.0\text{M } C_2H_3O_2^-$

c. 33.3 grams of Na_3P (molar mass = 100 g/mol) in 222 mL of total solution

$\frac{33.3\text{g } Na_3P \mid 1\text{mol}}{\mid 100\text{g}} = .333\text{mols}$

$M = \frac{n}{L}$

$= \frac{.333}{.222} = 1.5\text{M}$

$Na^+ = 4.5\text{M}$

$P^{3-} = 1.5\text{M}$

549. If 25 mL of each solution in #548 were compared, Na_3P would have the greatest conductivity. Why?

Gretest # of ions present in solution

550. How many grams of $Ba(C_2H_3O_2)_2$ are required to make 25 mL of a 1.5 M solution?

$n = M \times L$

$= (1.5)(.025)$

$= .0375\text{mols } Ba(C_2H_3O_2)_2$

$\frac{.0375\text{mols } Ba(C_2H_3O_2)_2 \mid 255\text{g}}{\mid 1\text{mol}} = 9.6\text{g } Ba(C_2H_3O_2)_2$

551. How would you prepare 10. mL of 0.70 M $Ba(C_2H_3O_2)_2$ using the solution made in #550?

$M_1V_1 = M_2V_2$

$(1.5)V_1 = (.70)(10)$

$V_1 = 4.67\text{mL}$

I would put 4.67 mL of the 1.5M solution from #550 into a 10 mL volumetric flask. I would then add DI water up to the 10mL line on the flask and mix thoroughly.

Note: A 10 mL graduated cylinder can also be used.

552. A reaction is carried out between 20. mL of 0.25 M $Na_2CO_3(aq)$ and 20. mL of 0.25 M $CaCl_2(aq)$ as shown below.

$$Na_2CO_3(aq) + CaCl_2(aq) \rightarrow CaCO_3(s) + 2\ NaCl(aq)$$

a. Determine the mass of $CaCO_3$ created.

$n_{Na_2CO_3} = M \times L = (.25)(.02) = .005\text{ mols } Na_2CO_3 \left| \frac{1\ CaCO_3}{1\ Na_2CO_3} \right. = .005\text{ mols } CaCO_3 \left| \frac{100.09\text{ g}}{1\text{ mol}} \right. = .50\text{ g } CaCO_3$

note: could have also used $CaCl_2$ and get the same answer. Perfect stoich ratio.

b. Determine the final molarity of NaCl(aq).

$.005\text{ mols } Na_2CO_3 \left| \frac{2\ NaCl}{1\ Na_2CO_3} \right. = .010\text{ mols } NaCl$

$M = \frac{n}{L} = \frac{.010}{.04\text{ L}}$

$M = 0.25\text{ M}$

553. Which of these is the best piece of glassware for preparing a 1.00 L aqueous solution?

[✓ a.] 1000 mL Volumetric flask
b. 1000 mL Erlenmeyer flask
c. 1000 mL Graduated beaker
d. 1000 mL Graduated cylinder

554. A 50. mL sample of 0.20 M $CaCl_2$ is mixed with a 25 mL sample of 0.30 M LiCl. No reaction occurs.

a. What is the $[Ca^{2+}]$ in the final solution?

$n = M \times L = (.20)(.05)$

$= .01\text{ mols } Ca^{2+}$

$M = \frac{n}{L} = \frac{.01}{.075} = 0.133\text{ M } Ca^{2+}$

OR

$M_1V_1 = M_2V_2$

$(.2)(50) = M_2(75)$

$M_2 = 0.133\text{M} = [Ca^{2+}]$

b. What is the $[Li^{+}]$ in the final solution?

$n = M \times L$

$= (.30)(.025)$

$= .0075\text{ mols}$

$M = \frac{n}{L}$

$= \frac{.0075}{.075}$

$= 0.1\text{ M}$

c. What is the [Cl$^-$] in the final solution?

$$\frac{.01 \text{ mols } CaCl_2 \mid 2Cl^-}{\mid 1CaCl_2} = .02 \text{ mols } Cl^-$$

$$\frac{.0075 \text{ mls } LiCl \mid 1Cl^-}{\mid 1LiCl} = .0075 \text{ mls } Cl^-$$

$n_{Cl^-} = .0275 \text{ mols}$ $\quad M = \frac{n}{L} = \frac{.0275 \text{ mols}}{.075 \text{ L}} = 0.37 \text{ M } Cl^-$

d. How many grams of AgNO$_3$ (s) would need to be added to the final solution to precipitate all of the Cl$^-$ assuming the reaction goes to completion?

$$Ag^+ (aq) + Cl^- (aq) \rightarrow AgCl(s)$$

$$\frac{.0275 \text{ mols } Cl^- \mid 1Ag^+}{\mid 1Cl^-} = .0275 \text{ mols } Ag^+ \qquad 1Ag^+ : 1AgCl$$

$$\frac{.0275 \text{ mls } AgCl \mid 143.32g}{\mid 1 \text{ mol}} = 3.9 \text{ g } AgCl$$

Section 7.04

Topic 7.3, 7.10 Worksheet

$$2\ H_2S\ (g) + CH_4\ (g) \rightleftharpoons CS_2\ (g) + 4\ H_2\ (g) \qquad K_c = 3.4 \times 10^{-4}$$

Use the reaction above to answer the questions that follow about Q, K_c, and the direction of the reaction.

555. A 0.10 mol sample of each of the four species in the reaction represented above is injected into a rigid 1.0 L container.

a. What is the value of Q?

$$Q = \frac{[CS_2][H_2]^4}{[H_2S]^2[CH_4]} = \frac{(.1)(.1)^4}{(.1)^2(.1)} = .01$$

b. In which direction will the reaction proceed to reach equilibrium? Justify your answer by comparing K to Q.

K Q

$3.4 \times 10^{-4} < .01$

The reaction will proceed to the left. Q is greater than K so reactants will form to bring Q in line with K to reestablish equilibrium.

c. Which species will have the highest concentration at equilibrium? Justify your answer.

The H_2S will have the highest concentration. The reaction shifts toward reactants. More H_2S will form because of the coefficient.

$2\ XY\ (g) \rightleftharpoons X_2\ (g) + Y_2\ (g)$ $K_p = 230$

556. Determine the direction the reaction will proceed to establish equilibrium given the initial conditions of each experiment.

Initial Conditions (atm)	Value of Q	Direction reaction will proceed to establish equilibrium	Will more reactants or products form in order to reach equilibrium?	The rate of the forward reaction is greater than, less than, or equal to the reverse reaction?
XY = 2.0 X_2 = 0.0 Y_2 = 0.0	0	Right	products	[✓ greater than] less than equal to
XY = 0.010 X_2 = 0.20 Y_2 = 2.0	4000	Left	reactants	greater than [✓ less than] equal to
XY = 0.0 X_2 = 5.0 Y_2 = 5.0	very large	Left	reactants	greater than [✓ less than] equal to
XY = 12.0 X_2 = 0.05 Y_2 = 0.05	1.74×10^{-5}	Right	products	[✓ greater than] less than equal to
XY = 1.0 X_2 = 1.0 Y_2 = 1.0	1	Right	products	[✓ greater than] less than equal to
XY = 0.064 X_2 = 0.97 Y_2 = 0.97	230	neither	neither	greater than less than [✓ equal to]

557. For any situation in 556 …

a. where K > Q, explain your reasoning for the direction the reaction will proceed.

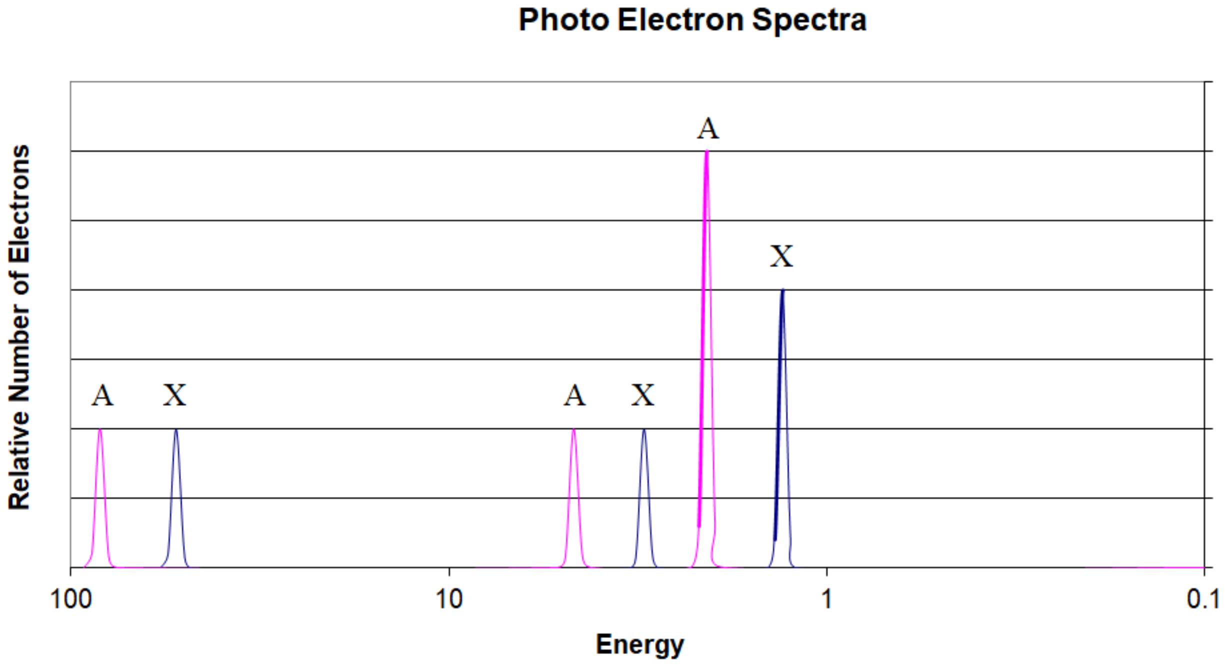
The reaction will shift right and proceed to make products until Q = K in order to establish equilibrium.

b. where K < Q, explain your reasoning for the direction the reaction will proceed.

The reaction will shift left and proceed to make reactants until Q = K in order to establish equilibrium.

c. where K = Q, explain your reasoning for the direction the reaction will proceed.

The system is in equilibrium.



558. Shown above are the complete PES of two different elements. All of the peaks labeled A belong to the same element and all of the peaks labeled X belong to the other element.

  a. What is the electron configuration of Element A? What is the name of the element?

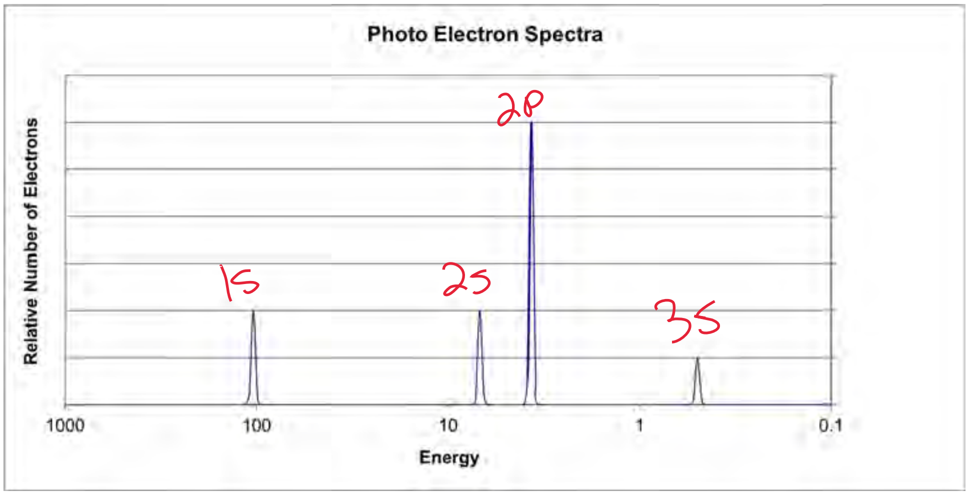
$1s^2 2s^2 2p^6$. The element is neon.

  b. What is the electron configuration of Element X? What is the name of the element?

$1s^2 2s^2 2p^4$. The element is oxygen.

  c. Explain why the last peak for A is further to the left than the last peak for X in terms of Coulombic attraction.

The last peak for A is further to the left because of greater Coulombic attraction compared to X. It takes more energy to remove the 2p electrons in A than it does in X. A has a greater nuclear charge than X because of more protons. This is the cause of the greater Coulombic attraction.



559. Use the complete PES above to answer the questions that follow.

a. How many total electrons does this atom have?

11

b. What is the symbol of the element?

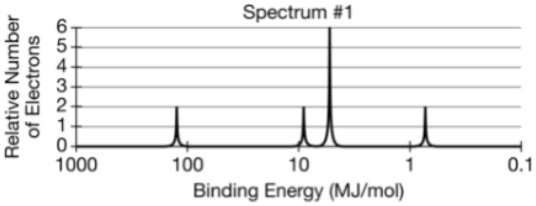
Na

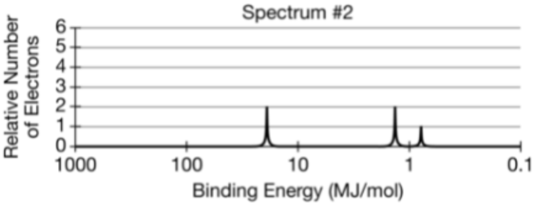
c. Label each peak with the sublevel represented.

d. Would you expect the first peak of the previous element to be to the left, to the right, or in the same spot as the first peak of this element? Justify your answer by discussing the nuclear charge of each element.

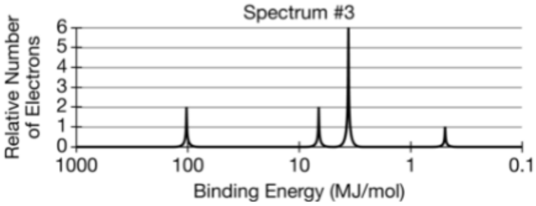
I would expect the first peak (the 1s peak) of the previous element to be to the right of the first peak of this element. The previous element has a smaller nuclear charge (fewer protons) and therefore has a smaller force of attraction for the first peak electrons. This element has a greater nuclear charge which causes a greater force of attraction for the first peak electrons. A greater force of attraction requires more energy to remove.

560. In a reaction, N_2 reacts with a pure metal X to form the compound X_3N. Which of the following complete photoelectron spectra could be that of metal X? Justify your choice.





Spectrum #3. The X forms a +1 cation (X^+) which means that it has 1 valence electron. Spectrum #3 has one valence electron ($1s^22s^22p^63s^1$).



561. A series of chemical reactions in equilibrium are used to determine the equilibrium constant of a third reaction, as shown below.

$2\ A\ (aq) + B\ (aq) \rightleftharpoons A_2B\ (aq)$	$K_1 = 25.0$
$CB\ (aq) + Y\ (aq) \rightleftharpoons CY\ (aq) + B\ (aq)$	$K_2 = 1.0 \times 10^{-3}$
$2A\ (aq) + CB\ (aq) + Y\ (aq) \rightleftharpoons A_2B\ (aq) + CY\ (aq)$	$K_3 = ???$

Determine the value of K

$K_3 = K_1 \times K_2 = .025$

562. Complete the table below

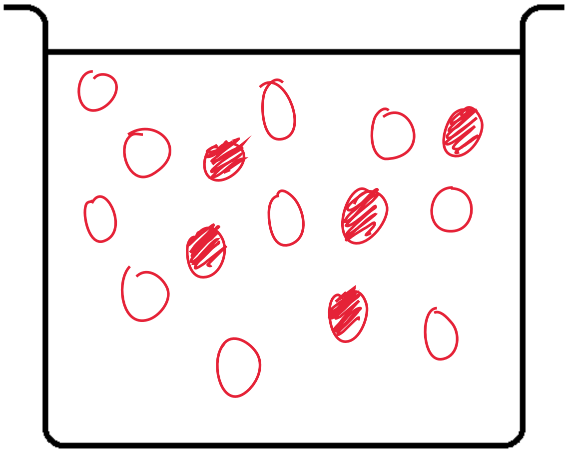
	Original Reaction		
Reaction	$PCl_5\ (g) \rightleftharpoons PCl_3\ (g) + Cl_2\ (g)$	$3\ PCl_5\ (g) \rightleftharpoons 3\ PCl_3\ (g) + 3\ Cl_2\ (g)$	$PCl_3\ (g) + Cl_2\ (g) \rightleftharpoons PCl_5\ (g)$
K expression	$\frac{(P_{PCl_3})(P_{Cl_2})}{P_{PCl_5}}$	$\frac{(P_{PCl_3})^3 (P_{Cl_2})^5}{(P_{PCl_5})^3}$	$\frac{P_{PCl_5}}{(P_{PCl_3})(P_{Cl_2})}$
K value (in terms of $K_{original}$)	↑ $K_{original}$	$K = (K_{original})^3$	$K = \frac{1}{K_{original}}$

Reaction	$1/3\ PCl_5\ (g) \rightleftharpoons 1/3\ PCl_3\ (g) + 1/3\ Cl_2\ (g)$	$2\ PCl_3\ (g) + 2\ Cl_2\ (g) \rightleftharpoons 2\ PCl_5\ (g)$
K expression	$\frac{(P_{PCl_3})^{1/3} (P_{Cl_2})^{1/3}}{(P_{PCl_5})^{1/3}}$	$\frac{(P_{PCl_5})^2}{(P_{PCl_3})^2 (P_{Cl_2})^2}$
K value (in terms of $K_{original}$)	$K = (K_{original})^{1/3}$	$K = \left(\frac{1}{K_{original}}\right)^2$

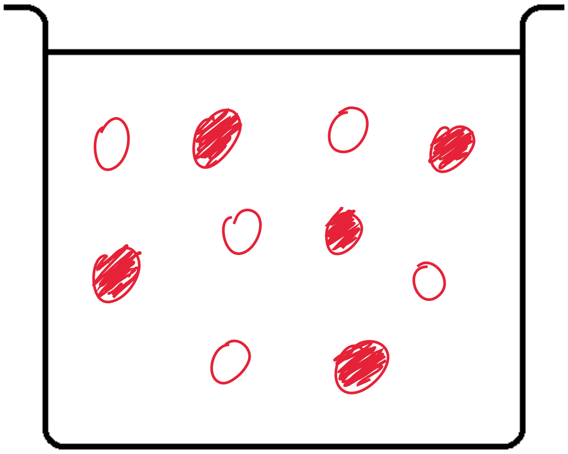
Review Topic 3.8

563. In the beakers below draw particulate representations of solutions. Use the symbols below for each solution:

a. Five particles of $(NH_4)_2S$ (s) are placed in the solution. Draw in the species needed to represent the major ionic species in the solution.



b. Five particles of $NaC_2H_3O_2$ (s) are placed in the water. Draw in the species needed to represent the major ionic species in the solution.



c. Five particles of $Pb(NO_3)_2$ (aq) react with eight particles of KI (aq). In the beaker below draw the resulting solution after the reaction has gone to completion. Draw in the species needed to represent all ionic species and any precipitate. Label all species.

L.R.

$$\underset{5}{Pb(NO_3)_2\,(aq)} + \underset{8}{2\,KI\,(aq)} \rightarrow \underset{8}{2\,KNO_3\,(aq)} + \underset{4}{PbI_2\,(s)}$$

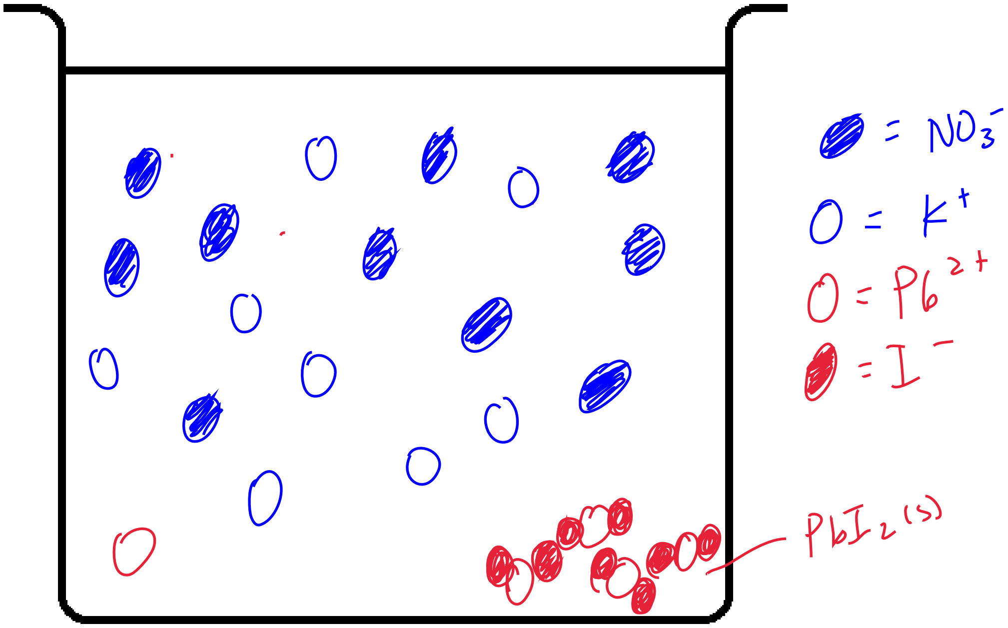
Final solution

1 Pb^{2+}

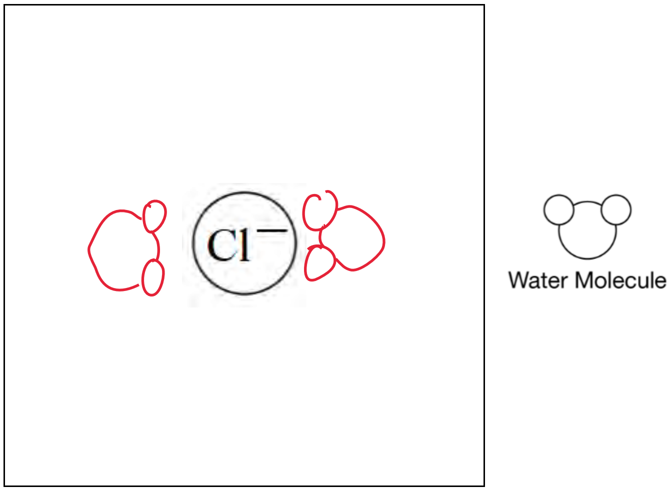
10 NO_3^-

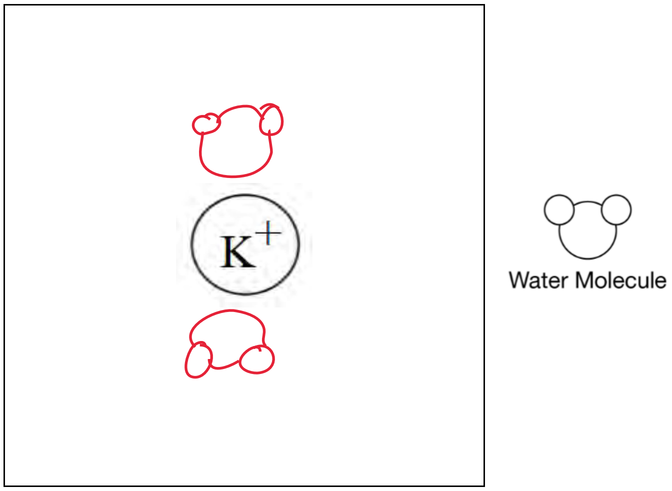
8 K^+

4 PbI_2



d. A particulate drawing of a single Cl^- ion and K^+ ion are shown in the boxes below. Draw two H_2O molecules in each box to show the most likely orientation of H_2O (l) molecules around the Cl^- ion.





Section 7.06

Topic 7.9 Worksheet

564. For the following situations, determine the direction the reaction will shift to reestablish equilibrium, explain the shift in terms of Q and K, and determine whether the reactants or the products will increase in concentration/pressure.

$$2\,BaO_2(s) \rightleftharpoons 2\,BaO(s) + O_2(g) \qquad \Delta H = 162\ \text{kJ/mol}_{rxn}$$

System Stress	Direction reaction shifts	Explanation based on Q and K	Increased concentration/pressure of reactants or products
Decrease pressure of O_2	[✓ Forward] Reverse None	K > Q. Reaction will shift right to form more products to bring Q in line with K to establish equilibrium.	[✓ Products increase] Reactants increase No increase
Increase mass of BaO_2	Forward Reverse [✓ None]	$BaO_2(s)$ does not appear in the equilibrium expression so the mass of BaO_2 does not effect equilibrium.	Products increase Reactants increase [✓ No increase]
Decrease temperature	Forward [✓ Reverse] None	A decrease in temperature favors the exothermic reaction because the particles have less energy. Less energy favors the exothermic reaction because of a smaller E_A for the exothermic reaction. K gets smaller as well.	Products increase [✓ Reactants increase] No increase
Increase size of container	[✓ Forward] Reverse None	K > Q. Increased size of container lowers the pressure and the Q as a result. Reaction will shift right to increase pressure and bring Q in line with K.	[✓ Products increase] Reactants increase No increase

$$2\,NO(g) + O_2(g) \rightleftharpoons 2\,NO_2(g) \qquad \Delta H < 0$$

System Stress	Direction reaction shifts	Explanation based on Q and K	Increased concentration/pressure of reactants or products
Decrease the size of the container	[✓ Forward] Reverse None	K > Q. The pressure of the reactants increases faster than the products. Reaction will shift right to increase pressure of products and bring Q back in line with K.	[✓ Products increase] Reactants increase No increase
Increase concentration of O_2	[✓ Forward] Reverse None	K > Q. The increased pressure of O_2 drops Q. The reaction will shift right to increase the pressure of products and bring Q back in line with K.	[✓ Products increase] Reactants increase No increase
Decrease temperature	[✓ Forward] Reverse None	K would get bigger. An increase in temperature increases the rate of the endothermic reaction more than the exothermic reaction. More particles have the required energy to get over the larger endo E_A.	[✓ Products increase] Reactants increase No increase
Use a catalyst	Forward Reverse [✓ None]	A catalyst has no effect on the equilibrium condition. Equilibrium will be established faster.	Products increase Reactants increase [✓ No increase]

565. Explain why an increase in temperature always favors the endothermic reaction.

The increase in temperature increases the rate of both the forward and reverse reaction. However, the increased energy will produce more particles that have enough energy to overcome the larger E_A of the endothermic reaction while having a smaller effect on the smaller E_A of the exothermic reaction. This results in a greater change in the endothermic direction.

566. What effect does an increase in temperature have on the size of K for …

a. an endothermic reaction?

Larger K. More products are formed as the forward reaction is favored.

b. an exothermic reaction?

Smaller K. More reactants are formed as the reverse reaction is favored.

567. Use the redox reaction below to answer the questions that follow.

$$Cu(s) + 2\,Ag^{+}(aq) \rightarrow Cu^{2+}(aq) + 2\,Ag(s) \qquad E^{o} = 0.460\ V$$

a. Would an increase in concentration of Ag^{+} increase, decrease, or have no effect on the voltage of the cell? Justify your answer in terms of Q and K.

Increase the voltage. As $[Ag^{+}]$ increases Q decreases and the voltage gets further away from 0.00 and further from equilibrium. According to the Nernst equation, as Q gets smaller the voltage gets bigger.

b. Would a decrease in concentration of Ag^{+} increase, decrease, or have no effect on the voltage of the cell? Justify your answer in terms of Q and K.

Decrease teh voltage. As $[Ag^{+}]$ decreases Q increases and the voltage gets closer to 0.00 V and closer to equilibrium. According to the Nernst equation, as Q gets bigger the voltage gets smaller.

c. Would an increase in concentration of Cu^{2+} increase, decrease, or have no effect on the voltage of the cell? Justify your answer in terms of Q and K.

Decrease the voltage. Same explanation as b.

d. Would a decrease in concentration of Cu^{2+} increase, decrease, or have no effect on the voltage of the cell. Justify your answer in terms of Q and K.

Increase the voltage. Same explanation as a.

Review Topic 4.4

568. When methanol, CH_3OH (l), evaporates …

   a. Describe what happens to the covalent bond between the oxygen and the hydrogen.

   Nothing happens to the covalent bond between the oxygen and hydrogen. Evaporation changes IMF and not covalent bonds.

   b. Describe what happens to the force of attraction between two different methanol particles.

   The force of attraction is broken between two different methanol particles.

569. A 20.00 gram sample of glycerol, $C_3H_8O_3$ (l) evaporates completely and becomes gaseous $C_3H_8O_3$ (g). Would the mass of $C_3H_8O_3$(g) be greater than, less than, or equal to 20.00 grams? Justify your answer.

   The mass would remain the same and be equal to 20.00 grams. The Law of Conservation of Mass states that mass should not be lost or gained in a physical or chemical change.

570. Do you agree or disagree with the statement in the box below? Explain your reasoning.

   A student makes the claim that when liquid propanol becomes gaseous propanol the chemical bonds between the propanol molecules are broken.

   I disagree. Chemical bonds do not hold the propanol molecules together, IMF hold them together.

Section 7.07

Topic 7.11 Worksheet

571. Define the following terms:

a. Solubility product

K_{sp}. The equilibrium constant of a slightly soluble or "insoluble" solid.

b. Solubility

The amount of solid, in either moles/L or grams/L, that dissolves in water.

572. Give the dissociation reaction, then determine the solubility, in mols L^{-1}, of each of the following solids given their K_{sp}

a. AgCl ($K_{sp} = 1.8 \times 10^{-10}$)

$AgCl_{(s)} \rightleftharpoons Ag^+ + Cl^-$

$x \quad x \quad x$

$K_{sp} = [Ag^+][Cl^-]$

$K_{sp} = (x)(x) = 1.8 \times 10^{-10}$

$\sqrt{x^2} = \sqrt{1.8 \times 10^{-10}}$

$x = 1.3 \times 10^{-5}$ M = solubility of AgCl

b. AgBr ($K_{sp} = 5.0 \times 10^{-13}$)

$AgBr_{(s)} \rightleftharpoons Ag^+ + Br^-$

$x \quad x \quad x$

$K_{sp} = [Ag^+][Br^-]$

$K_{sp} = (x)(x) = 5.0 \times 10^{-13}$

$\sqrt{x^2} = \sqrt{5.0 \times 10^{-13}}$

$x = 7.1 \times 10^{-7}$ M

c. Ag_2CrO_4 ($K_{sp} = 1.1 \times 10^{-12}$)

$Ag_2CrO_{4(s)} \rightleftharpoons 2Ag^+ + CrO_4^{2-}$

$x \quad 2x \quad x$

$K_{sp} = [Ag^+]^2[CrO_4^{2-}]$

$K_{sp} = (2x)^2(x) = 1.1 \times 10^{-2}$

$4x^3 = 1.1 \times 10^{-2}$

$\sqrt[3]{x^3} = \sqrt[3]{\frac{1.1 \times 10^{-2}}{4}}$

$x = 6.5 \times 10^{-5}$ M

d. $PbCrO_4$ ($K_{sp} = 2.8 \times 10^{-13}$)

$PbCrO_{4(s)} \rightleftharpoons Pb^{2+} + CrO_4^{2-}$

$x \quad x \quad x$

$K_{sp} = [Pb^{2+}][CrO_4^{2-}]$

$K_{sp} = (x)(x) = 2.8 \times 10^{-13}$

$\sqrt{x^2} = \sqrt{2.8 \times 10^{-13}}$

$x = 5.3 \times 10^{-7}$ M

e. $Cr(OH)_3$ ($K_{sp} = 1.6 \times 10^{-30}$)

$Cr(OH)_{3\,(s)} \rightleftharpoons Cr^{3+} + 3OH^-$

$K_{sp} = [Cr^{3+}][OH^-]^3$

$K_{sp} = (x)(3x)^3 = 1.6 \times 10^{-30}$

$27x^4 = 1.6 \times 10^{-30}$

$\sqrt[4]{x^4} = \sqrt[4]{\frac{1.6 \times 10^{-30}}{27}}$

$x = 1.6 \times 10^{-8}$ M

573. For each of the solids from 572 above, give the solubility of the anion.

a. AgCl ($K_{sp} = 1.8 \times 10^{-10}$)

$[Cl^-] = 1.34 \times 10^{-5}$ M

b. AgBr ($K_{sp} = 5.0 \times 10^{-13}$)

$[Br^-] = 7.1 \times 10^{-7}$ M

c. Ag_2CrO_4 ($K_{sp} = 1.1 \times 10^{-12}$)

$[CrO_4^{2-}] = 6.5 \times 10^{-5}$ M

d. $PbCrO_4$ ($K_{sp} = 2.8 \times 10^{-13}$)

$[CrO_4^{2-}] = 5.3 \times 10^{-7}$ M

e. $Cr(OH)_3$ ($K_{sp} = 1.6 \times 10^{-30}$)

$[OH^-] = 3 \times 1.6 \times 10^{-8}$

$= 4.8 \times 10^{-8}$

574. How would the solubility of #573c change if it were in 50.0 mL of total solution?

The solubility would not change. The amount dissolved would but not the solubility because the ratio of mols to L remains the same

575. What condition must be true to be able to directly compare K_{sp} values to determine relative solubility?

Same # of ions for both species being compared

576. For the salts given below, circle the more soluble salt based solely on K_{sp}.

a. [✓ AgCl] ($K_{sp} = 1.8 \times 10^{-10}$) or AgI ($K_{sp} = 8.3 \times 10^{-17}$)

b. [✓ Ag_2CrO_4] ($K_{sp} = 1.1 \times 10^{-12}$) or $Co(OH)_2$ ($K_{sp} = 1.3 \times 10^{-15}$)

c. [✓ CaF_2] ($K_{sp} = 3.9 \times 10^{-11}$) or Ag_2S ($K_{sp} = 6.0 \times 10^{-51}$)

577. A 1.00 L saturated solution of calcium oxalate (CaC_2O_4) at 25 °C contains 0.0061 g of CaC_2O_4.

a. Write the dissociation reaction of CaC_2O_4, including phase symbols.

$$CaC_2O_{4(s)} \rightleftharpoons Ca^{2+} + C_2O_4^{2-}$$

b. Calculate the solubility product constant, K_{sp}, for this salt at 25 °C.

$$\frac{.0061\text{ g } CaC_2O_4}{1\text{ L}} \cdot \frac{1\text{ mol}}{128.06\text{ g}} = 4.76 \times 10^{-5}\text{ M}$$

$$CaC_2O_{4(s)} \rightleftharpoons Ca^{2+} + C_2O_4^{2-}$$

$x \quad x \quad x$

$4.76 \times 10^{-5} \quad 4.76 \times 10^{-5} \quad 4.76 \times 10^{-5}$

$$K_{sp} = [Ca^{2+}][C_2O_4^{2-}]$$

$$= (4.76 \times 10^{-5})(4.76 \times 10^{-5})$$

$$= 2.3 \times 10^{-9}$$

578. The solubility of fluoride in a saturated solution of SrF_2 is 1.76 x 10^{-3} M.

a. Write the dissociation reaction of SrF_2, including phase symbols.

$SrF_{2(s)} \rightleftharpoons Sr^{2+} + 2F^-$

b. Determine the solubility of SrF_2.

$SrF_{2(s)} \rightleftharpoons Sr^{2+} + 2F^-$

x x 2x

8.8×10^{-4} 8.8×10^{-4} 1.76×10^{-3}

$2x = 1.76 \times 10^{-3} M = [F^-]$

$x = 8.8 \times 10^{-4} M = [SrF_2]$

c. Determine the K_{sp} of SrF_2.

$K_{sp} = [Sr^{2+}][F^-]^2 = (8.8 \times 10^{-4})(1.76 \times 10^{-3})^2$

$K_{sp} = 2.7 \times 10^{-9}$

579. If 0.0490 g of $AgIO_3$ dissolves per liter of solution, calculate the solubility product constant, K_{sp} .

$AgIO_{3(s)} \rightleftharpoons Ag^+ + IO_3^-$

x x x

$\frac{.0490 g\ AgIO_3 \mid 1 mol}{\mid 282.8 g} = 1.73 \times 10^{-4} M = x$

$K_{sp} = [Ag^+][IO_3^-]$

$= (1.73 \times 10^{-4})(1.73 \times 10^{-4})$

$K_{sp} = 3.0 \times 10^{-8}$

580. Determine if a precipitate will form for the following solutions. In each case, justify your answer based on Q.

a. A 500 mL solution of 0.0250 M $Pb(NO_3)_2$ mixed with a 255 mL solution of 0.0045 M solution of Na_2CrO_4. The K_{sp} of $PbCrO_4$ is 2.8 x 10^{-13}.

$PbCrO_{4(s)} \rightleftharpoons Pb^{2+} + CrO_4^{2-}$

$(.5)(.025)$ $(.255)(.0045)$

$\frac{.0125}{.755 L} = .0166 M$ $\frac{.00115}{.755} = .00152 M$

$K_{sp} = [Pb^{2+}][CrO_4^{2-}] = Q$

$(.0166)(.00152) = Q$

$2.5 \times 10^{-5} = Q$

K Q

$2.8 \times 10^{-13} < 2.5 \times 10^{-5}$

a ppt. forms b/c K < Q. The $PbCrO_4$ will ppt. out until K = Q

b. A 255 mL solution of 0.0150 M $Co(NO_3)_2$ mixed with a 150 mL solution of 0.0122 M solution of NaOH. The K_{sp} of $Co(OH)_2$ is 1.3 x 10^{-15}.

$Co(OH)_{2(s)} \rightleftharpoons Co^{2+} + 2OH^-$

$(.255)(.015)$ $(.150)(.0122)$

$\frac{.00383}{.405} = .00944 M$ $\frac{.00183}{.405} = .00452 M$

$K_{sp} = [Co^{2+}][OH^-]^2 = Q$

$= (.00944)(.00452)^2$

$1.93 \times 10^{-7} = Q$

K Q

$1.3 \times 10^{-15} < 1.93 \times 10^{-7}$

a ppt. forms b/c K < Q. The $Co(OH)_2$ will ppt. out until K = Q

581. In each of the following situations, determine the concentration, in mols/L, of the anion needed to precipitate out the solid.

a. What molarity of Cl^- is needed to precipitate out AgCl ($K_{sp} = 1.8 \times 10^{-10}$) in a solution that is 0.05 M Ag^+?

$AgCl_{(s)} \rightleftharpoons Ag^+ + Cl^-$

$x \quad .05M \quad x$

$1.8 \times 10^{-10} = (.05)(x)$

$x = 3.6 \times 10^{-9} = [Cl^-]$ needed to reach equilibrium.

Any amount greater than 3.6×10^{-9} M will cause precipitation because Q will be > K

$Ksp = [Ag^+][Cl^-]$

b. How many moles of NaI are needed to form a AgI ($K_{sp} = 8.3 \times 10^{-17}$) precipitate in 50.0 mL of total solution if the $[Ag^+]$ is 1.5×10^{-4} M? (Assume the NaI does not change the volume of the solution.)

$AgI_{(s)} \rightleftharpoons Ag^+ + I^-$

$x \quad 1.5 \times 10^{-4} \quad x$

$8.3 \times 10^{-17} = (1.5 \times 10^{-4})(x)$

$x = 5.53 \times 10^{-13} M = [I^-]$

$K_{sp} = [Ag^+][I^-]$

$M = \frac{n}{L}$

$M \times L = n$

$(5.53 \times 10^{-13})(.05) = n = 2.77 \times 10^{-14}$ mols NaI

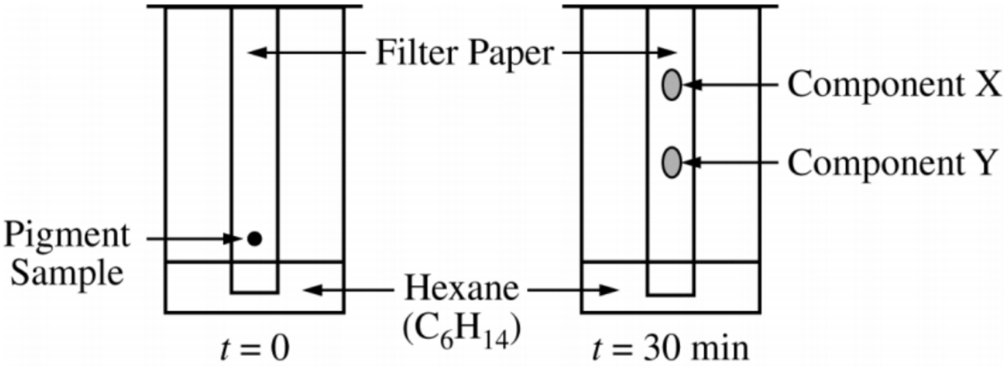
582. A student performs a distillation procedure on a mixture composed of the two miscible liquids hexane (C_6H_{14}) and decane ($C_{10}H_{22}$).

a. Explain why distillation would be used and not filtration.

The two liquids are miscible/dissolved into each other. Filtration is used when a solid is in an aqueous mixture. There is no solid in this mixture.

b. Which liquid would be initially present in higher concentration in the distillate? Justify your answer by referring to the relative strengths of the intermolecular forces present in each molecule.

Hexane would be present in higher concentration in the distillate. The IMF of hexane are weaker than the IMF of decane. As a result hexane will boil/evaporate away first because it takes less energy to reach the boiling point of hexane compared to decane.



583. A student performs a paper chromatography experiment to separate two components of a pigment. The stationary phase is moderately polar and hexane is used for the mobile phase.

a. Explain why the student can conclude that component Y is more polar than X.

Component Y is not as attracted to the mobile phase, hexane. Hexane is nonpolar. Since component Y did not travel as far it has IMF less similar than component X.

b. Explain why the student can conclude that component X is more nonpolar than Y.

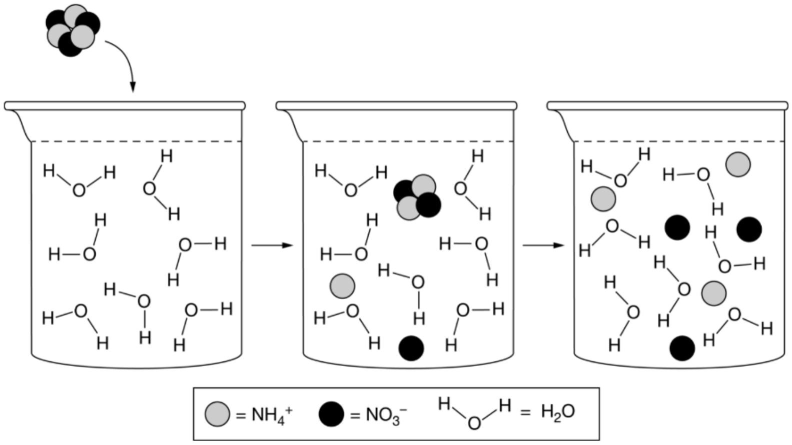
Component X is more attracted to the mobile phase, hexane. Hexane is nonpolar. Since component X traveled further than Y we can assume that X has IMF more similar to hexane.

c. How would the results after 30 minutes change if a 50% ethyl acetate solution (dipole moment of 1.88 D) had been used instead of hexane (dipole moment of 0.09 D)?

Component Y would travel further than Component X since Component Y is more polar than Component X.

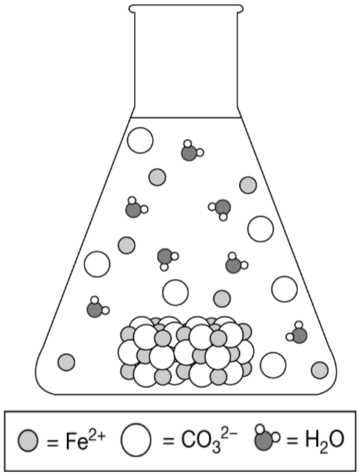
Section 7.08

Topic 7.14 Worksheet



584. Explain how the image above illustrates an increase in entropy.

In the final beaker there are more particles. Since there are more particles there are more possible microstates. A greater number of microstates means greater entropy.



585. The Erlenmeyer flask above is allowed to sit out on a lab bench such that all of the water evaporates.

a. What would happen to the aqueous ions in solution?

The aqueous ions would precipitate out and form a crystal of $FeCO_3$.

b. Explain why this process would result in a decrease in entropy.

Once the water has evaporated there are fewer particles and fewer microstates. Also, the more mobile aqueous ions become a solid that has a more restrictive movement. Both decrease the amount of entropy present.

Element	Electronegativity
H	2.1
C	2.5
S	2.5
F	4.0
Cl	3.0
Si	1.8

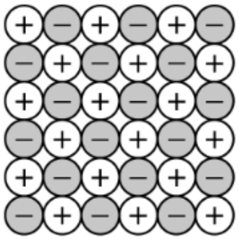
586. According to the table above …

a. Which two different elements would create the most polar bond? Justify your answer by referring to the electronegativities in the table above.

Si and F would create the most polar bond. Si and F would have the greatest difference in electronegativity.

b. Which two different elements would create the most nonpolar bond? Justify your answer by referring to the electronegativities in the table above.

Si and H would create the most nonpolar bond. Si and H would have the smallest difference in electronegativity.



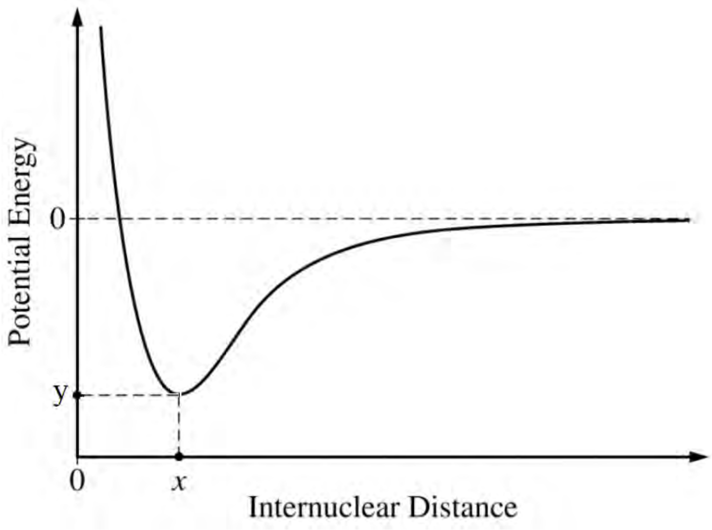
587. The figure above represents KF.

a. How would the figure change if it were KCl instead? Explain the reason for the difference using principles of atomic structure.

The only difference would be that the Cl^- ions would be larger than the F^- ions. Cl^- has an additional energy shell compared to F^-. The additional energy shell would make the ion larger as the electrons are located further from the nucleus.

b. How would the figure change if it were NaF instead? Explain the reason for the difference using principles of atomic structure.

The only difference would be that the Na^+ ions would be smaller than the K^+ ions. Na^+ has one fewer energy shell compared to K^+. The fewer number of energy shells would make the ion smaller as the electrons are located closer to the nucleus.



588. The potential energy of a system of two third period, single bonded atoms as a function of their internuclear distance is shown in the diagram above.

   a. Consider two second period atoms bonding instead.

      i. What would happen to the placement of x on the x-axis? Explain the reason for the difference using principles of atomic structure.

The "x" would be further to the left. The second period atoms are smaller than the 3rd period atoms. Since they are smaller the distance between the nuclei would be smaller as well.

      ii. What would happen to the placement of y on the y-axis? Explain the reason for the difference using principles of atomic structure.

The "y" would be lower (further from 0). The second period elements are closer to each other because of their smaller size compared to the third period elements. The smaller bond makes a stronger bond. The stronger bond would require more energy to break apart or would give off more energy when it forms. As a result "y" would be lower.

   b. Consider two of the same third period atoms that now create a double bond.

      i. What would happen to the placement of x on the x-axis? Explain the reason for the difference by discussing the forces of attraction that occur in the double bond.

The "x" of the double bond would be to the left. A double bond is stronger than a single bond and as a result makes a smaller bond. A smaller distance between the atoms would make "x" appear to the left.

      ii. What would happen to the placement of y on the y-axis? Explain the reason for the difference by discussing the forces of attraction that occur in the double bond.

The "y" would be lower. A double bond is stronger since the atoms are closer to each other. A stronger bond takes more energy to break or more energy is released when it is formed. As a result "y" would be lower.

$$CdF_2\ (s) \rightleftharpoons Cd^{2+}\ (aq) + 2\ F^-\ (aq)$$

589. A saturated solution of CdF_2 is prepared. The equilibrium in the solution is represented above. In the solution $[Cd^{2+}]_{eq}$ = 0.0585 M and $[F^-]_{eq}$ = 0.117 M.

a. Determine the solubility of CdF_2.

$CdF_{2(s)} \rightleftharpoons Cd^{2+} + 2F^-$ $[CdF_2] = .0585\ M$

x x 2x

.0585M .0585 M

b. Some 0.90 M NaF is added to the saturated solution. Does the solubility of CdF_2 increase, decrease, or remain the same? Justify your answer based on Q.

The solubility of CdF_2 decreases. As F^- is added Q increases. The reaction will shift left to the reactants, precipitating out CdF_2 to bring Q down and back in line with K.

590. Is the solubility of AgCl(s) greater in distilled water or in tap water where the $[Cl^-]$ = 0.010 M? Justify your answer based on Q. The K_{sp} of AgCl = 1.8×10^{-10}.

The solubility of AgCl is greater in distilled water. The extra Cl^- in tap water will increase Q. AgCl will have to precipitate out as the reaction shifts left to the reactants to decrease Q until it equals K.

591. The solubility of CuBr(s) is to be measured in four different solutions: distilled water, an NaBr(aq) solution, an $NaNO_3$(aq) solution, and a $CuNO_3$(aq) solution. Determine if the solubility of CuBr will be greater, less, or the same as in distilled water. Justify your answer based on Q or Le Chatelier's Principle.

a. NaBr(aq) solution

Less soluble. Increased [Br^-] will increase Q. CuBr will have to ppt. out to reestablish equilibrium.

b. $NaNO_3$(aq) solution

The solubility will be the same. $NaNO_3$ has no effect on the equilibrium position.

c. $CuNO_3$(aq) solution

Less soluble. The increased [Cu^+] will increase Q. See answer to a. above.

592. A solution of $AgNO_3(aq)$ reacts with a lump of Zn (s) as shown in the reaction below.

$$2AgNO_3(aq) + Zn(s) \rightarrow Zn(NO_3)_2(aq) + 2Ag(s)$$

a. Give the net-ionic equation for the reaction.

$2\,Ag^+ + Zn(s) \longrightarrow Zn^{2+} + 2\,Ag(s)$

b. What are the spectator ions?

NO_3^-

c. Does the mass of Zn(s) increase, decrease, or remain the same as the reaction proceeds? Explain your reasoning.

The mass of Zn(s) will decrease as the reaction proceeds. Zn(s) is converted to Zn^{2+} and becomes aqueous. As a result less Zn(s) will remain at the end of the reaction.

Note: The total mass of Zn atoms remains the same.

593. A 20. mL sample of acetic acid, $HC_2H_3O_2$ ($K_a = 1.8 \times 10^{-5}$), is titrated with NaOH(aq) and the reaction below proceeds.

$$HC_2H_3O_2(aq) + NaOH(aq) \rightarrow NaC_2H_3O_2(aq) + H_2O(l)$$

a. Acetic acid dissociates in water as shown below,

$$HC_2H_3O_2(aq) + H_2O(l) \rightleftharpoons H_3O^+(aq) + C_2H_3O_2^-(aq) \qquad K_a = 1.8 \times 10^{-5}$$

According to the K_a would there be more $HC_2H_3O_2$ or $C_2H_3O_2^-$ in a solution of acetic acid?

According to the K_a there would be more $HC_2H_3O_2$ in the solution. A small K_a means more reactants.

b. In the net-ionic equation the reactant acetic acid appears as a molecule and not as an ion. Explain why by referring to the equilibrium constant, K_a.

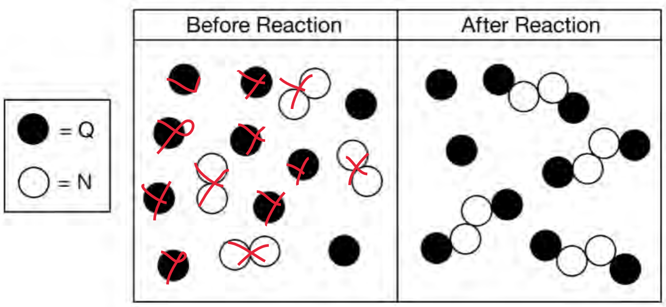
Since the K_a is small we have mainly reactants. This means that $HC_2H_3O_2$ is present more than the ion. As a result $HC_2H_3O_2$ appears as a molecule since very little of it breaks apart.

c. Give the net-ionic equation for the reaction between $HC_2H_3O_2$ and NaOH.

$$HC_2H_3O_2(aq) + OH^-(aq) \rightarrow H_2O(l) + C_2H_3O_2^-(aq)$$

594. The reaction between aqueous barium chloride ($BaCl_2$) and aqueous sodium sulfate (Na_2SO_4) forms a precipitate of barium sulfate ($BaSO_4$). Give the net-ionic equation for the reaction.

$Ba^{2+} + SO_4^{2-} \rightarrow BaSO_4 (s)$



Use the figure above to answer the questions that follow.

595. Give the balanced equation for the reaction.

$2Q + N_2 \rightarrow Q_2N_2$

596. Determine the limiting reactant. Justify your answer by referring to the particles.

The limiting reactant is N_2. The "After Reaction" box still has unreacted Q particles.

597. Assuming that all reactants and products are gases, would you expect the pressure of the container to increase, decrease, or remain the same? Explain your reasoning.

I would expect the pressure of the container to decrease. 3 moles of gas (2Q and 1 N_2) are being converted to 1 mole of gas (Q_2N_2). Fewer particles means less pressure.

OR 14 particles of gas becomes 6 particles of gas. Fewer particles means less pressure.

598. Would you expect Q to be a metal or nonmetal? Explain your reasoning.

I would expect Q to be a nonmetal. The product is a covalently bonded particle. If it were ionic the particles would need to be arranged so that the positive ions would be near the negative ions. The particle has two N atoms next to each other which would not occur in an ionic bond.

Section 7.10

Topic 7.13 Worksheet

599. For the solutions below at 25℃, circle whether they have a high concentration of H^+ or a high concentration of OH^-.

		High Concentration of …		
a.	A solution with a low pH.	[✓ H^+]	OH^-	Neither
b.	A solution with a high pH.	H^+	[✓ OH^-]	Neither
c.	A solution with a pH of 10.2	H^+	[✓ OH^-]	Neither
d.	A solution with a pH of 1.3	[✓ H^+]	OH^-	Neither
e.	A solution with a pH of 7.00	H^+	OH^-	[✓ Neither]

600. Determine if the solubility would increase, decrease, or remain the same for the following solids placed into the solution indicated.

a. Would the solubility of $Cu(OH)_2$ (s) increase, decrease, or remain the same in the solutions given below. In each case, explain your reasoning. (Relevant equations are given below.)

Relevant Equations: $Cu(OH)_2(s) \rightleftharpoons Cu^{2+}(aq) + 2\,OH^-(aq)$ $K_{sp} = 1.6 \times 10^{-19}$
$H^+(aq) + OH^-(aq) \rightarrow H_2O(l)$

i. Low pH

Increase. At a low pH there are H^+ ions that would react w/ the OH^- ions, removing the OH^- ions from solution to form H_2O. As OH^- is removed more $Cu(OH)_2$ will dissolve to reestablish equilibrium.

ii. High pH

Decrease. A high pH has excess OH^- ions. The excess OH^- ions would cause the reaction to shift to the reactants to remove the excess OH^- and reestablish equilibrium.

b. Would the solubility of $FeCO_3$ (s) increase, decrease, or remain the same in the solutions given below? In each case, explain your reasoning. (Relevant equations are given below.)

Relevant Equations:		
	$FeCO_3(s) \rightleftharpoons Fe^{2+}(aq) + CO_3^{2-}(aq)$	$K_{sp} = 3.5 \times 10^{-11}$
	$2\ H^+(aq) + CO_3^{2-}(aq) \rightarrow H_2O(l) + CO_2(g)$	
	$Fe(OH)_2(s) \rightleftharpoons Fe^{2+}(aq) + OH^-(aq)$	$K_{sp} = 4.87 \times 10^{-17}$

i. Low pH Increase. A low pH has H^+ that will react with CO_3^{2-} to come out of solution. More $FeCO_3(s)$ will dissolve to create more CO_3^{2-} to reestablish equilibrium.

ii. High pH Increase. At high pH the Fe^{2+} will precipitate out as $Fe(OH)_2$ since the solubility of $Fe(OH)_2$ (2.3×10^{-6} M) is less than $FeCO_3$ ($5\ 9 \times 10^{-6}$ M). As $[Fe^{2+}]$ decreases more $FeCO_3$ dissolves to reestablish equilibrium.

c. Would the solubility of CaF_2 (s) increase, decrease, or remain the same in the solutions given below? In each case, explain your reasoning. (Relevant equations are given below.)

Relevant Equations:		
	$CaF_2(s) \rightleftharpoons Ca^{2+}(aq) + 2\ F^-(aq)$	$K_{sp} = 3.9 \times 10^{-11}$
	$HF(aq) \rightleftharpoons H^+(aq) + F^-(aq)$	$K_a = 6.6 \times 10^{-4}$
	$Ca(OH)_2(s) \rightleftharpoons Ca^{2+}(aq) + 2\ OH^-(aq)$	$K_{sp} = 6.5 \times 10^{-6}$

i. Low pH Increase. The F^- will form HF. The CaF_2 will dissolve more to increase [F] to reestablish equilibrium

ii. High pH Increase. The Ca^{2+} will ppt. out as $Ca(OH)_2$ More CaF_2 will dissolve to reestablish equilibrium and increase $[Ca^{2+}]$.

$$3\ S_8 + 8\ OH^- \rightarrow 8\ S_3^- + 4\ HOOH$$

601. In an experiment, a student studies the kinetics of the reaction represented above and obtains the data shown in the following table.

Experiment	Initial $[S_8]$ (M)	Initial $[OH^-]$ (M)	Initial Reaction Rate (M/s)
1	0.0500	0.0500	0.334
2	0.150	0.0500	1.02
3	0.150	0.100	2.06

a. Determine the rate law for the reaction.

$Rate = k[S_8]^1[OH^-]^1$

S_8

$$\left(\frac{.15}{.05}\right)^x = \frac{1.02}{.334}$$

$$3^x = 3$$

$$x = 1$$

OH^-

$$\left(\frac{.1}{.05}\right)^x = \left(\frac{2.06}{1.02}\right)$$

$$2^x = 2$$

$$x = 1$$

b. Determine the value of k, along with the units.

$Rate = k[S_8]^1[OH^-]^1$

$.334\ M/s = k(.05)(.05)$

$1.34 \times 10^2\ M^{-1}s^{-1} = k$

c. The student performs experiment 3 again with the same concentrations but at a lower temperature. The student notices that the rate is less than 2.06 M/s. Give two reasons why the rate is lower by referring to the collision theory.

At a lower temperature the particles have less energy. With less energy the particles are traveling slower and have fewer collisions. At lower energy there are also fewer particles with enough energy to overcome the activation energy, E_A.

d. In a different trial, the initial rate of disappearance of S_8 is determined to be 1.35×10^{-2} M/s.

   i. What would be the rate of disappearance of OH^- over the same time period?

$$\frac{1}{3}\,\text{Rate}_{S_8} = \frac{1}{8}\,\text{Rate}_{OH^-}$$

$$\frac{1}{3}(1.35 \times 10^{-2}) = \frac{1}{8}x$$

$$4.5 \times 10^{-3} = \frac{1}{8}x$$

$$x = 3.60 \times 10^{-2}\ M/s = \text{Rate}_{OH^-}$$

   ii. The appearance of S_3^- is determined to be the same as the disappearance of OH^-. Do you agree or disagree with the results? Justify your answer by referring to the balanced equation.

I agree with the results. The coefficient of both S_3^- and OH^- is 8. For every 1 mole of OH^- used 1 mole of S_3^- will be formed. Therefore the rates should be equal but opposite in sign.

602. A reaction mechanism has three elementary steps. The first and third step are fast relative to the second step, which is considered slow. All three steps occur at the same temperature. Give two different reasons the second step could be slow.

The second step could have a large activation energy.

The second step could be termolecular (have 3 molecules colliding).

603. Consider the reaction mechanism below:

Step 1: $SO_2 + \cancel{V_2O_5} \rightarrow SO_3 + \cancel{V_2O_4}$

Step 2: $\cancel{V_2O_4} + \frac{1}{2} O_2 \rightarrow \cancel{V_2O_5}$

a. Give the overall balanced equation.

$$SO_2 + \frac{1}{2} O_2 \longrightarrow SO_3$$

b. Does the concentration of V_2O_5 increase, decrease, or remain the same as the reaction proceeds? Justify your answer by referring to the reaction mechanism.

The [V_2O_5] remains the same as the rxn proceeds. V_2O_5 is formed in Step 1 and disappears in Step 2.

c. Is V_2O_5 a catalyst or an intermediate? Explain your reasoning.

Catalyst. It appears as a reactant and leaves as a product.

604. The photodissociation of formaldehyde occurs in the upper troposphere. A proposed reaction mechanism is shown below.

Fast $H_2CO \rightarrow H + \cancel{HCO}$
Slow $\cancel{HCO} + O_2 \rightarrow \cancel{CO} + HO_2$
Fast $\cancel{CO} + OH \rightarrow CO_2 + H$

a. Give the overall balanced equation.

$$H_2CO + O_2 + OH \longrightarrow CO_2 + HO_2 + 2H$$

b. Determine the rate law expression.

$$Rate = k[H_2CO][O_2]\left[\frac{1}{H}\right]$$

Unit 8

Section 8.01

Topic 8.1, 8.2 Worksheet

605. Write one equation that can be used to calculate …

a. the pH of a solution if [H_3O^+] is known.

$$pH = -\log[H_3O^+]$$

b. the pOH of a solution if [OH^-] is known.

$$pOH = -\log[OH^-]$$

c. the pH of a solution if [OH^-] is known.

$$pH = 14 + \log[OH^-]$$

d. the pOH of a solution if [H_3O^+] is known.

$$pOH = 14 + \log[H_3O^+]$$

e. the pH of a solution if pOH is known.

$$pH = 14 - pOH$$

f. the [H_3O^+] if pH is known.

$$[H_3O^+] = 10^{-pH}$$

g. the [OH^-] if pOH is known.

$$[OH^-] = 10^{-pOH}$$

h. the [H_3O^+] if [OH^-] is known.

$$[H_3O^+] = \frac{1\times10^{-14}}{[OH^-]}$$

$$5\,H_2O_2(aq) + 2\,MnO_4^-(aq) + 6\,H^+(aq) \rightarrow 2\,Mn^{2+}(aq) + 8\,H_2O(l) + 5\,O_2(g)$$

606. Does the pH of the solution in the reaction above increase, decrease, or remain the same as the reaction proceeds? Justify your answer.

The pH increases because as the reaction proceeds H^+ is being used. A decrease in [H^+] will increase pH.

607. A neutral solution of water, with pH = 7.00, is heated to 50 °C and the pH drops to 6.63.

a. Did the ionization of water increase or decrease with an increase in temperature? Justify your answer.

The ionization increased. Since the pH dropped more H^+ ions were released into solution.

b. What is the [H_3O^+]?

$$[H_3O^+] = 10^{-pH} = 10^{-6.63} = 2.34\times10^{-7}\ M$$

c. What is the [OH^-]?

Water is always neutral ∴ $[H_3O^+] = [OH^-] = 2.34\times10^{-7}\ M$

d. Calculate the value of K_w at 50 °C.

$K_w = [H_3O^+][OH^-] = (2.34 \times 10^{-7})^2 = 5.50 \times 10^{-14}$

e. Is the solution still neutral? Justify your answer.

Yes. The $[H_3O^+] = [OH^-]$ therefore it is still neutral

f. Does the value of K_w increase or decrease with an increase in temperature?

The value of K_w increases w/ an increase in temperature

g. Does the value of pK_w increase or decrease with an increase in temperature?

@ 25°C $pK_w = 14$

@ 50°C $pK_w = 13.26$

pK_w decreases w/ an increase in temp

608. By what factor must a solution of a strong acid be diluted to increase the pH by 1? Give an example to justify your answer.

By a factor of 10

$[H^+] = .01\ M$; $pH = 2$

$[H^+] = .001\ M$; $pH = 3$

609. What are the formulas and names of the six strong acids?

HCl hydrochloric acid

HI hydroiodic acid

H_2SO_4 sulfuric acid

HBr hydrobromic acid

$HClO_4$ perchloric acid

HNO_3 nitric acid

610. What physical property of an acid makes it a strong acid?

It dissociates completely

611. What is the general formula of a strong base?

metal hydroxide ; MOH

612. What physical property of a base makes it a strong base?

It dissociates completely

613. Give the reaction of HCl(aq) dissolved in water.

$HCl_{(aq)} + H_2O(l) \longrightarrow H_3O^+(aq) + Cl^-(aq)$

614. Calculate the pH of the strong acids and bases given below:

a. A 0.002 M solution of HCl

$HCl + H_2O \longrightarrow H_3O^+ + Cl^-$

$.002 M \qquad .002 M$

$pH = -\log[H_3O^+]$

$= -\log(.002) = 2.7$

b. A 3.45 x 10^{-4} M solution of HNO_3

$pH = -\log[H_3O^+]$

$pH = -\log(3.45 \times 10^{-4}) = 3.462$

c. A solution made by dissolving 3.2 g of KOH into 450 mL of total solution.

$\frac{3.2g\ KOH \mid 1 mol}{\mid 56.108g} = .0570\ mols\ KOH = n\ OH^-$

$M = \frac{n}{L} = \frac{.0570}{.45\ L} = .127 M\ KOH$

$pOH = -\log[OH^-]$

$pOH = -\log(.127)$

$pOH = .896$

$pH + pOH = 14$

$pH = 13.10$

d. 100 mL of a 1 x 10^{-4} M HBr solution.

$pH = -\log[H_3O^+]$

$= -\log(1 \times 10^{-4})$

$= 4.0$

e. The solution from 614.d that is diluted to a total volume of 1000 mL.

$M_1V_1 = M_2V_2$

$(1\times10^{-4})(100) = M_2(1000)$

$M_2 = 1\times10^{-5}$ M

$pH = -\log(1\times10^{-5})$

$pH = 5.0$

f. 100 mL of a 0.10 M HNO_3 solution added to 100 mL of 0.25 M HCl and diluted to a final volume of 1200 mL.

HNO_3

$M \times L = n$

$(.1)(.1) = n$

.01 mols H^+

HCl

$M \times L = n$

$(.25)(.1) = n$

$.025 = n_{H^+}$

$n_{total} = n_{HCl} + n_{HNO_3}$

$n_{tot} = .025 + .01$

$n_{tot} = .035$ mols H^+

$M = \frac{n}{L}$

$M = \frac{.035}{1.2L}$

$M = 2.92\times10^{-2}$

$pH = -\log[H_3O^+]$

$= -\log(2.92\times10^{-2})$

$pH = 1.5$

g. A solution of $Ca(OH)_2$ made by dissolving 120. g in 3500 mL of total solution.

$\frac{120g\ Ca(OH)_2 \mid 1mol}{\mid 74.096g} = 1.62mols\ Ca(OH)_2$

$Ca(OH)_2 \rightarrow Ca^{2+} + 2OH^-$

1.62 mols   3.24mols

$pOH = -\log\left(\frac{3.24}{3.5}\right)$

$pOH = .0337$

$pH = 13.966$

615. Complete the following table.

[H⁺] or [OH⁻]	pH between …	
	[✓ 1 to 2]	2 to 3
$[H^+] = 1.25 \times 10^{-2}$ M	3 to 4	4 to 5
	5 to 6	6 to 7

	1 to 2	2 to 3
$[H^+] = 4.56 \times 10^{-4}$ M	[✓ 3 to 4]	4 to 5
	5 to 6	6 to 7

	1 to 2	2 to 3
$[OH^-] = 7.88 \times 10^{-11}$	[✓ 3 to 4]	4 to 5
	5 to 6	6 to 7

616.Identify the acid, base, conjugate acid, and conjugate base in the following reactions.

HCl (aq)	+	NH_3(*aq*)	→	NH_4^+(aq)	+	Cl^-(aq)
☒ Acid		☐ Acid		☐ Acid		☐ Acid
☐ Base		☒ Base		☐ Base		☐ Base
☐ Conjugate Acid		☐ Conjugate Acid		☒ Conjugate Acid		☐ Conjugate Acid
☐ Conjugate Base		☐ Conjugate Base		☐ Conjugate Base		☒ Conjugate Base

CH_3CH_2COOH (aq)	+	H_2O(*l*)	→	H_3O^+(aq)	+	$CH_3CH_2COO^-$(aq)
☒ Acid		☐ Acid		☐ Acid		☐ Acid
☐ Base		☒ Base		☐ Base		☐ Base
☐ Conjugate Acid		☐ Conjugate Acid		☒ Conjugate Acid		☐ Conjugate Acid
☐ Conjugate Base		☐ Conjugate Base		☐ Conjugate Base		☒ Conjugate Base

$NaNH_2$ (aq)	+	HF(*aq*)	→	HNH_2(aq)	+	NaF(aq)
☐ Acid		☒ Acid		☐ Acid		☐ Acid
☒ Base		☐ Base		☐ Base		☐ Base
☐ Conjugate Acid		☐ Conjugate Acid		☒ Conjugate Acid		☐ Conjugate Acid
☐ Conjugate Base		☐ Conjugate Base		☐ Conjugate Base		☒ Conjugate Base

617.For the acid, give the conjugate base. For the base, give the conjugate acid.

Acid	Conjugate Base	Base	Conjugate Acid
$HC_2H_3O_2$	$C_2H_3O_2^-$	CH_3NH_2	$CH_3NH_3^+$
HF	F^-	C_5H_5N	$C_5H_5NH^+$
C_5H_5COOH	$C_6H_5COO^-$	H_2O	H_3O^+

618. Equimolar samples of He(g), Ar(g), and Ne(g) are placed into a rigid, 2 L container.

a. Would the initial partial pressure of He(g) be greater than, less than, or equal to the partial pressure of Ar(g)?

The initial pressure would be the same. If an equal number of particles are in the same container they will exert the same pressure because P = nRT/V and n, R, T, and V are all the same.

b. What would occur to the total pressure of the gases in the container if the temperature of the container is decreased?

If the temperature is decreased the total pressure would also decrease.

Gas	Boiling Point (°C)
He	−269
Ar	−186
Ne	−246

c. The container is cooled to –190 °C and the total pressure is 3.4 atm.

i. The partial pressure of Ar is nearly 0 atm. Explain why.

Ar becomes a liquid at temperatures below -186 ^oC. As a liquid Ar does not exert as great a pressure as it does as a gas.

ii. Determine the partial pressure of He.

The partial pressure of helium would be 1.7 atm. There are only two gases left in the container therefore $P_{total} = P_{He} + P_{Ne}$

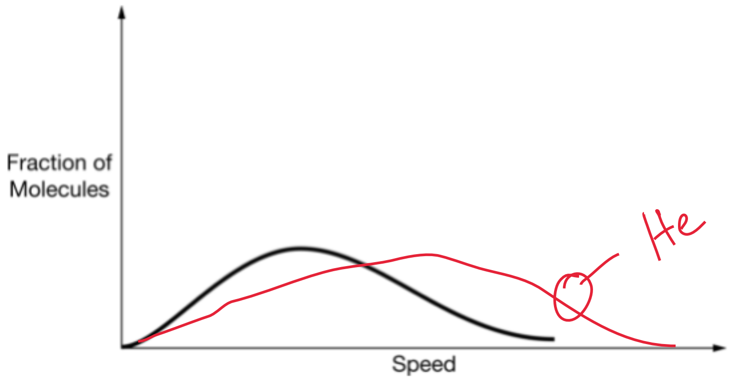
iii. Is the average kinetic energy of Ar(l) greater than, less than or equal to the average kinetic energy of Ne(g)?

The average KE is the same since they are at the same temperature.

iv. How does the average speed of Ne(g) compare to the average speed of He(g)? Explain your reasoning.

The average speed of Ne(g) would be less than He(g). Their KE are the same. KE= $1/2mv^2$. Ne has a greater mass therefore the velocity must be smaller than He.

v. The graph below shows the Maxwell-Boltzman distribution of Ne(g). On the same graph, draw the distribution curve for He(g) at the same temperature.



vi. Do you agree or disagree with the student statement below? Explain your reasoning.

A student makes the claim that the mass of He(g) in the container is the same as the mass of Ne(g) in the container.

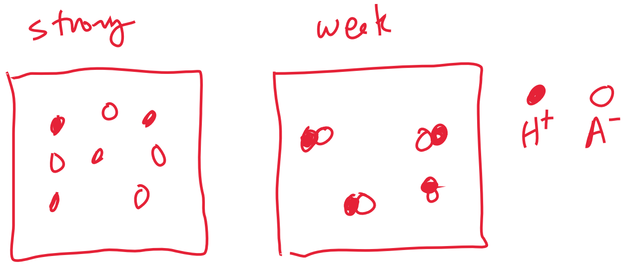
I disagree with the statement. There are the same number of particles in the container which means the same number of moles. However, 1 mole of He does not weigh the same as 1 mole of Ne therefore they will not have the same mass in the container.

vii. A small hole is on the top of the container allowing gases to effuse out. Rank the gases from slowest to effuse out of the container to fastest to effuse out of the container.

He will effuse out faster than Ne. If Ar becomes a gas it will effuse out the slowest.

619. In terms of ionization, what is the difference between a strong acid and a weak acid? Use particle pictures in your explanation.

A strong acid dissociates completely while a weak acid does not dissociate completely.



620. Write the reaction for the following weak acids reacting with water. Then give the K_a expression.

a. $CH_3CH_2COOH(aq)$

$$CH_3CH_2COOH(aq) + H_2O(l) \rightleftharpoons H_3O^+(aq) + CH_3CH_2COO^-(aq)$$

$$K_a = \frac{[H_3O^+][CH_3CH_2COO^-]}{[CH_3CH_2COOH]}$$

b. $HF(aq)$

$$HF(aq) + H_2O(l) \rightleftharpoons H_3O^+(aq) + F^-(aq)$$

$$K_a = \frac{[H_3O^+][F^-]}{[HF]}$$

c. $HSO_3^-(aq)$

$$HSO_3^-(aq) + H_2O(l) \rightleftharpoons H_3O^+(aq) + SO_3^{2-}(aq)$$

$$K_a = \frac{[H_3O^+][SO_3^{2-}]}{[HSO_3^-]}$$

621. Different weak acids have different K_a values.

a. Does the percent ionization of a weak acid increase, decrease, or remain the same as K_a increases? Justify your answer.

The percent ionization increases as K_a increases. As K_a increases, the $[H_3O^+]$ increases. The more H_3O^+ produced the greater the % ionization.

$$\% = \frac{[H_3O^+]}{(HA]} \times 100$$

b. If the solutions are equimolar, does the pH of a weak acid increase, decrease, or remain the same as K_a increases? Justify your answer.

The pH would decrease. As K_a increases the amount of H_3O^+ increases. As H_3O^+ increases, pH decreases.

622. Give the pK_a or pK_b. Then rank them based on pH assuming they are all 0.100 M and the molecule gives or accepts one proton. A high pH should be ranked 5 and a low pH should be ranked 1.

K_a	pK_a value	Relative pH Rank
1.2×10^{-4}	3.921	2
4.22×10^{-2}	1.375	1
5.00×10^{-6}	5.301	4
6.22×10^{-7}	6.206	5
7.8×10^{-6}	5.108	3

K_b	pK_b value	Relative pH Rank
3.8×10^{-7}	6.420	2
2.2×10^{-5}	4.658	3
7.9×10^{-8}	7.102	1
4.11×10^{-3}	2.386	5
6.7×10^{-4}	3.174	4

623. Determine either the K_a or the pK_a for the following acids.

Acid	K_a	pK_a
A	3.2×10^{-4}	3.49
B	1.6×10^{-4}	3.80
C	5.0×10^{-5}	4.30
D	6.2×10^{-6}	5.21
E	9.8×10^{-7}	6.01

624. If each of the acids in 623 above were monoprotic and 0.1 M, which would have the lowest pH? Explain your reasoning.

Acid A. Acid A has the largest Ka value. The larger the Ka the greater the dissociation. The greater the dissociation the lower the pH.

625. Determine the K_a for the following weak monoprotic acids.

in each case

$$HA_{(aq)} + H_2O_{(l)} \rightleftharpoons H_3O^+_{(aq)} + A^-_{(aq)}$$

&

$$K_a = \frac{[H_3O^+][A^-]}{[HA]}$$

a. A 0.10 M solution that has a pH of 4.0.

$[H_3O^+] = 10^{-4}$

$[A^-] = 10^{-4}$

$$K_a = \frac{(10^{-4})(10^{-4})}{.1 - 10^{-4}}$$

$$K_a = 1.00 \times 10^{-7}$$

b. A 0.0010 M solution that has a pH of 6.0.

$$K_a = \frac{(10^{-6})(10^{-6})}{.001 - 10^{-6}} = 1.00 \times 10^{-9}$$

c. A 0.050 M solution that has a pH of 5.0.

$$K_a = \frac{(10^{-5})(10^{-5})}{.050 - 10^{-5}} = 2.00 \times 10^{-9}$$

$$HCN(aq) + H_2O(l) \rightleftharpoons H_3O^+(aq) + CN^-(aq)$$

626. HCN partially ionizes as shown above. For each of the situations, determine if the pH would increase, decrease, or remain the same. In each case, justify your answer by referring to K and Q. (Assume no change in volume.)

a. After the equilibrium has been established, a sample of solid NaCN is added to the solution.

The pH would increase. With the addition of NaCn the Q>K meaning the reaction will shift left to form more reactants to reestablish equilibrium. Forming more reactants consumes H_3O^+ which would result in a higher pH value.

b. After the system has reached equilibrium, a sample containing Pb^{2+} is added to the solution forming $Pb(CN)_2(s)$.

The pH would decrease. As $Pb(CN)_2$ is formed Q<K meaning the reaction would shift right forming more products to reestablish equilibrium. H_3O^+ would increase in concentration which would make the pH drop.

627. Determine the K_b for the following weak bases.

$B + H_2O(l) \rightleftharpoons HB^+ + OH^-$

$$K_b = \frac{[HB^+][OH^-]}{[B]}$$

a. A 0.10 M solution that has a pH of 11.000

pH = 11

pOH = 3

OH = 10^{-3}

$$K_b = \frac{(10^{-3})(10^{-3})}{.1 - 10^{-3}}$$

$$K_b = 1 \times 10^{-5}$$

b. A 0.0010 M solution that has a pH of 10.00

$$K_b = \frac{(10^{-4})(10^{-4})}{.001 - 10^{-4}} = 1 \times 10^{-5}$$

c. A 0.050 M solution that has a pH of 8.00

$$K_b = \frac{(10^{-6})(10^{-6})}{.05 - 10^{-6}} = 2 \times 10^{-11}$$

628. When will the pH of a strong acid be equal to the pH of a weak acid?

When the strong acid is dilute and the weak acid is concentrated

629. Determine the K_a of a 0.15 M weak, monoprotic acid that has a pH of 2.20 .

$$K_a = \frac{[H_3O^+][A^-]}{[HA]} = \frac{(10^{-2.2})(10^{-2.2})}{.15 - 10^{-2.2}}$$

$pH = 2.20$

$[H_3O^+] = 10^{-2.2} = [A^-]$

$K_a = 2.77 \times 10^{-4}$

630. The pH of a 1.15 M weak base is 12.65. Determine the K_b value.

$$K_b = \frac{[OH^-][HB^+]}{[B]}$$

$pH = 12.65$

$pOH = 1.35$

$[OH^-] = 10^{-1.35}$

$[HB^+] = 10^{-1.35}$

$$K_b = \frac{(10^{-1.35})(10^{-1.35})}{1.15 - 10^{-1.35}}$$

$K_b = 1.81 \times 10^{-3}$

631. The ionization constant (K_a) of acid A is 1.8 x 10^{-5} and acid B is 4.5 x 10^{-9}.

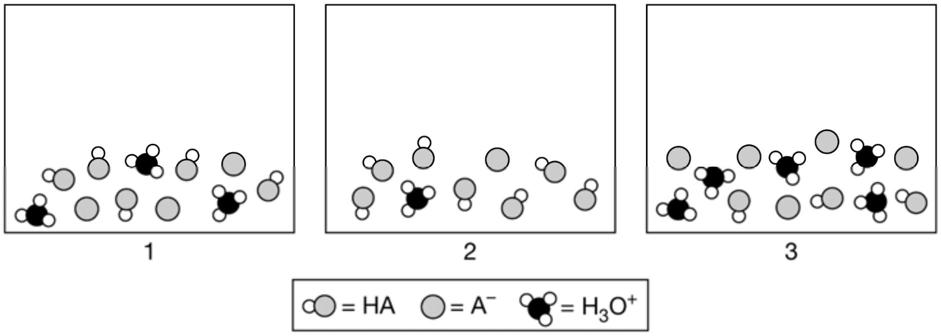
a. Which is a stronger acid?

Acid A. Larger Ka means greater dissociation

b. Which has a stronger conjugate base?

Acid B. The weaker the acid the stronger the conjugate base.

632. The acids shown in the particle diagrams below all have the same molarity. Arrange the acids from weakest to strongest. Justify your answer.



2 < 1 < 3

#2 dissociates the least while #3 dissociates the most.

The more dissociation the stronger the acid

Review Topic 3.11

Use the Phet "Molecules and Light".

633. Select "Microwave". How does microwave radiation effect a nonpolar molecule like N_2? How does microwave radiation effect a polar molecule like NO_2?

No effect on N_2.

makes NO_2 spin

634. Select "Infrared". How does infrared radiation effect a molecule like carbon dioxide?

Makes bonds vibrate

635. Select "Ultraviolet". How does ultraviolet radiation effect a molecule like ozone or NO_2? Notice the action of the specific bonds.

Ozone breaks at apart at the single bond producing O & O_2

It does the same thing to NO_2 producing O & N=O

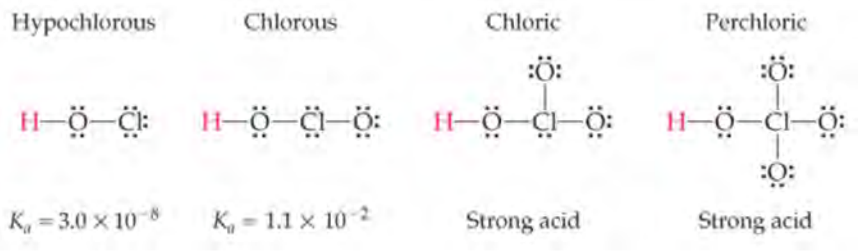
Section 8.03

Topic 8.6 Worksheet

636. Consider HOI and HOCl. Which is a weaker acid? Justify your answer in terms of the electronegativity of the halogen.

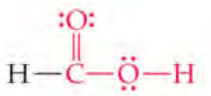
HOI is the weaker acid. Cl is more electronegative than the I. The more electronegative Cl makes OCl^- more stable and less reactive.

637. Consider the four acids shown below. Explain why acidity increases as the number of oxygens added to the halogen increases in terms of electronegativity.

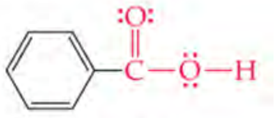


The O is very electronegative. As a result, the anion becomes more stable with increasing number of O atoms.

638. Consider the two carboxylic acids shown below. The K_a of formic acid is 1.8×10^{-4} and the K_a of benzoic acid is 6.3×10^{-5}.



Formic acid



Benzoic acid

a. Which carboxylic acid is a stronger acid? Explain why in terms of molecular structure.

Formic acid is stronger. The benzoic acid pumps electrons into the COOH making it less stable and thus more reactive. (Inductive effect)

b. Which conjugate base is more stable? Justify your answer by referring to the K_b.

$K_b$formic = 5.56×10^{-11}
$K_b$benzoic = 1.59×10^{-10}

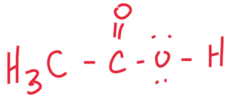
The conjugate base of benzoic acid is more stable. The K_b of formic acid is smaller meaning more reactant which means the conjugate base is less stable/more reactive.

c. Write the equation for the reaction that occurs between benzoic acid and water.

d. Write the equation for the reaction that occurs between formic acid and water.

$HCOOH_{(aq)} + H_2O_{(l)} \rightleftharpoons H_3O^+_{(aq)} + HCOO^-_{(aq)}$

639.Explain why CH_3COOH is an acid while CH_3OH is not.



$H_3C-\overset{\overset{O}{\|}}{C}-\ddot{\underset{\cdot\cdot}{O}}-H$

The additional O in CH_3COOH stabilizes the CH_3COO^- b/c of its electronegativity.

$H_3C-\ddot{\underset{\cdot\cdot}{O}}-H$

640.Explain why CH_3COOH is a stronger acid than CH_3CH_2COOH.

CH_3COOH is a stronger acid b/c CH_3COO^- is more stable than $CH_3CH_2COO^-$ b/c of the inductive effect. The extra "CH_2" group on CH_3CH_2COOH pumps e^- into COO^- making it less stable

$CO(g) + H_2O(g) \rightleftharpoons CO_2(g) + H_2(g)$ $K_c = 1.5 \times 10^3$

641. A 3.0 mol sample of CO(g) and a 3.0 mol sample of $H_2O(g)$ are introduced into an evacuated 100. L rigid container where the temperature is held constant as the reaction proceeds.

a. Give the equilibrium expression.

$$K_c = \frac{[CO_2][H_2]}{[CO][H_2O]}$$

b. Assume the mixture is allowed to reach equilibrium. How does the $[CO_2]$ compare to the [CO]? Justify your answer by referring to the equilibrium constant.

The $[CO_2]$ will be greter than the [CO]. The value of K_c means more products will form than reactants.

c. Determine the equilibrium concentration of $H_2(g)$.

$$K_c = \frac{[CO_2][H_2]}{[CO][H_2O]} = 1.5 \times 10^3.$$

$$\sqrt{\frac{x^2}{(.003-x)^2}} = \sqrt{1.5 \times 10^3}$$

$$\frac{x}{.003-x} = 38.73$$

$x = .116 - 38.73x$

$39.73x = .116$

$x = .00292M$

$[H_2] = .00292\ M$

	I	C	E
CO_2	0	+x	x
H_2	0	+x	x
CO	.003	−x	.003 − x
H_2O	.003	−x	.003 − x

d. The K_c and K_p for this reaction are the same. Explain why.

The change in the number of moles of gas is zero.

e. After reaching equilibrium $CO_2(g)$ is added to the reaction mixture. In which direction must the reaction shift in order to reestablish equilibrium?

The reaction will shift toward the reactants in order to reestablish equilibrium. The excess CO_2 must be consumed as Q>K.

$$2\ NO_2(g) \rightleftharpoons N_2O_4(g)$$

dark brown — colorless

642. The reaction shown above occurs in an enclosed clear glass ampule. The K_c for the reaction at 35 °C is 12.5.

a. What color is the ampule at equilibrium? Explain your reasoning.

The ampule should be clear. A K_c of 12.5 at 35°C means the products are favored. The products are colorless.

b. The temperature is increased and the ampule turns dark brown.

i. Is the reaction exothermic or endothermic? Justify your answer

The reaction is exothermic. An increase in temperature favors the endothermic reaction. Since the ampule became dark brown the reactants were favored which means that the reverse reaction is endothermic. Therefore the forward reaction is exothermic.

ii. Does the value of K increase, decrease, or remain the same?

The value of K decreases as more reactants are formed to reach equilibrium.

$$H_2(g) + I_2(g) \rightleftharpoons 2\ HI(g) \qquad K_c = \frac{[HI]^2}{[H_2][I_2]}$$

643. At 450 °C the value of K_c is 50. Determine if the concentration of $H_2(g)$ will increase, decrease, or remain the same under the following conditions as the system progresses toward equilibrium. Explain your reasoning by discussing the relationship between Q and K.

a. $[H_2]$ = 1.0 M, $[I_2]$ = 1.0 M, $[HI]$ = 1.0 M

K Q

$50 > 1$

The $[H_2]$ will decrease. $K > Q$ so more products need to form to establish equilibrium.

b. $[H_2]$ = 0.005 M, $[I_2]$ = 0.005 M, $[HI]$ = 4.0 M

K Q

$50 < 64000$

The $[H_2]$ will increase. $K < Q$ so more reactants need to form to establish equilibrium.

c. $[H_2]$ = 0.05 M, $[I_2]$ = 1.6 M, $[HI]$ = 2.0 M

K Q

$50 = 50$

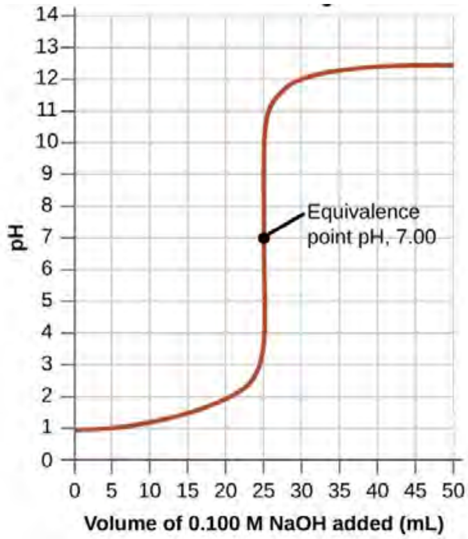
The $[H_2]$ will remain the same $K = Q$ so the system has reached equilibrium.

Section 8.04

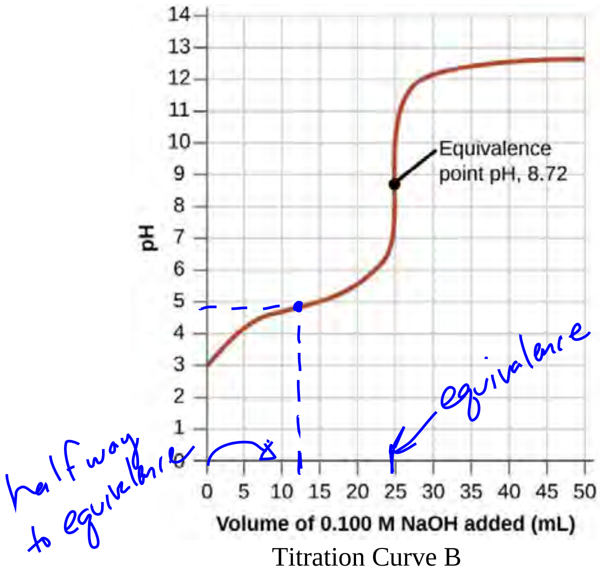
Topic 8.5 Worksheet

644. Explain the process of the titration of an acid with a base.

A titration is used to determine the molarity of a solution with a known chemical reaction. Typically, base is added to the buret and the unknown molarity acid is titrated until a visible change in color occurs due to an acid-base indicator added to the acid.



Titration Curve A



Titration Curve B

645. Consider the two titration curves given above.

a. Which titration curve above is of a strong acid/strong base titration? Explain how you know.

Curve A. The pH at equivalence is 7.

b. Explain why the other titration curve is of a weak acid/strong base titration.

pH at equivalence is greater than 7.

c. What is the pK_a and K_a of the weak acid?

Halfway to equivalence the pH = pKa.

$pH \cong 4.8 \quad \therefore \quad pKa = 4.8$

$Ka = 10^{-4.8} = 1.58 \times 10^{-5}$

only works at equivalence

646.For the following titrations, determine the molarity of the monoprotic acid.

a. In a titration, 15.0 mL of acid reaches equivalence with 23.8 mL of 0.100 M base. (Assuming base only accepts 1 H+)

$M_A V_A \# H^+ = M_B V_B \# OH^-$

$(M_A)(15)(1) = (23.8)(.1)(1)$

$M_A = 0.159M$

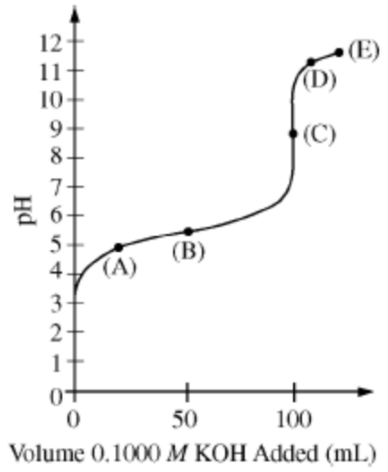
b. 13.4 mL of 0.125 M base reaches equivalence with 25.0 mL of a weak acid.

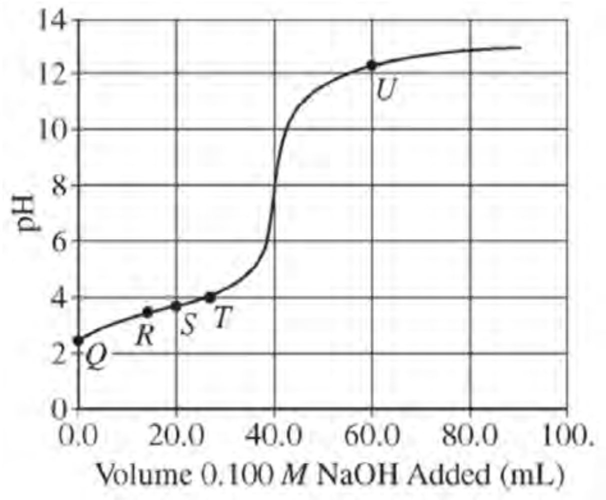
$M_A V_A \# H^+ = M_B V_B \# OH^-$

$(M_A)(25)(1) = (.125)(13.4)(1)$

$M_A = .067M$

647.Determine the pK_a and K_a of the following acids based on the titration curve.



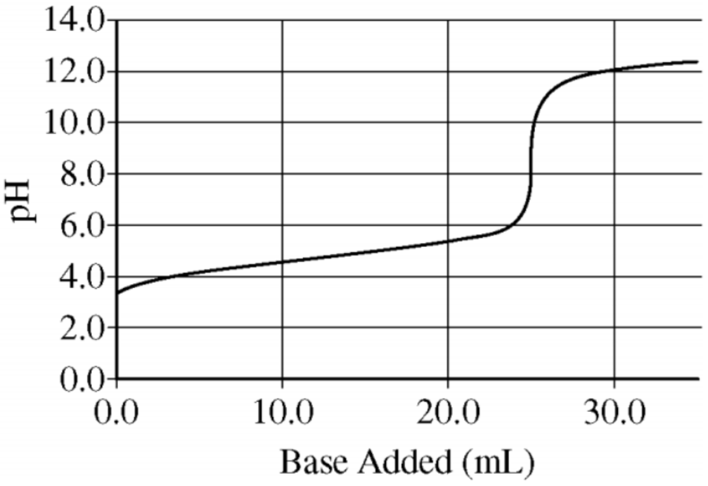


$pKa \cong 5.6$

$Ka = 10^{-5.6} = 2.51 \times 10^{-6}$

$pKa \cong 3.5$

$Ka = 10^{-3.5} = 3.16 \times 10^{-4}$



$pKa \cong 5.0$

$Ka = 10^{-5} = 1 \times 10^{-5}$

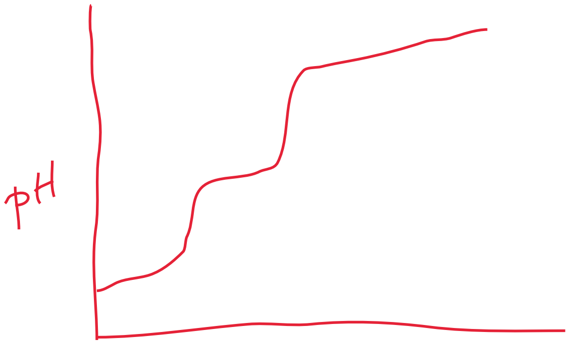
648. For a weak acid/strong base titration, explain why the pH = pK_a at half-way to the equivalence point.

At halfway to equivalence, the molarity of the acid is equal to the molarity of the salt (conjugate base). Thus in the Henderson-Hasselbalch equation pH = pKa + log(1) and the log(1) = 0.

649. Draw a general titration curve for …

a. H_3PO_4

b. H_2SO_4



650. Explain the method to …

a. determine equivalence via titration and acid/base indicator..

At equivalence, the indicator will change color.

b. determine equivalence via pH electrodes.

At equivalence the pH of a strong acid/strong base is 7.00.

651. Define:

a. Titrant

The solution in the buret.

b. Analyte

What's being analyzed, in the flask.

c. Equivalence

All of the acid reacts with all of the base.

d. Equivalence point

The pH when all of the acid reacts with all of the base.

e. End point

The pH when the indicator changes color.

652. At which point in a titration are the concentrations of the weak acid and its conjugate base approximately equal?

Halfway to equivalence.

653. A titration is carried out to determine the molarity of an unknown acid. Determine if the following would increase, decrease, or have no effect on the calculated molarity. Explain your reasoning for each.

a. You use an indicator with an endpoint slightly past the equivalence point.

If you go past the equivalence point you will use too much base. That will increase the moles of acid which would increase the molarity of the acid.

b. You use an indicator with an endpoint slightly before the equivalence point.

The molarity of the acid would decrease. If the indicator changes color before equivalence you will use fewer moles of base. Fewer moles of base means fewer moles of acid. Fewer moles of acid divided by the same volume of acid means a lower molarity.

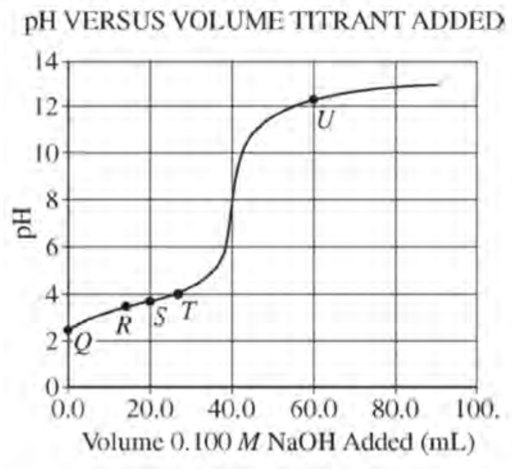
c. You choose the wrong indicator. The indicator you chose should be used for a strong acid/strong base titration but you are carrying out a weak acid/strong base titration.

Same situation as b. because the pH at equivalence for strong acid/base is less than the pH at equivalence for a weak acid/strong base titration.

d. You choose the wrong indicator. The indicator you chose should be used for a weak acid/strong base titration but you are carrying out a strong acid/strong base titration.

Same situation as a. The indicator will change AFTER the equivalence point has been reached.

654. For the titration curve given below, first determine which species are present by placing an “X” in the box then determine which species has the highest concentration by also placing an “O” in the box.



Point	HA	A^-	H_3O^+	OH^-
Q	X, O	X	X	
R	X, O	X	X	
S	X, O	X, O	X	
T	X	X, O	X	
U		X		X, O

$HCOOH(aq) \rightleftharpoons HCOO^-(aq) + H^+(aq)$ $K_a = 1.8 \times 10^{-4}$

655. Formic acid, HCOOH, dissociates in water as shown in the equation above. A 25.0 mL sample of an aqueous solution of pure formic acid is titrated using standardized 0.150 M NaOH.

a. After addition of 15.0 mL of the 0.150 M NaOH, the pH of the resulting solution is 4.37. Calculate each of the following.

i. $[H^+]$ in the solution

$[H^+] = 10^{-pH}$ $[H^+] = 10^{-4.37}$ 4.27×10^{-5} M $= [H^+]$

ii. $[OH^-]$ in the solution

$[OH^-] = \frac{Kw}{[H^+]} = \frac{1 \times 10^{-14}}{4.27 \times 10^{-5}} = 2.34 \times 10^{-10}$ M $= [OH^-]$

iii. The number of moles of NaOH added

$M \times V = mols$

$(.150)(.015) = mols$

2.25×10^{-3} mols NaOH

iv. The number of moles of $HCOO^-$ (aq) in the solution.

Before equivalence

mols OH^- = mols $HCOO^-$

2.25×10^{-3} mols $HCOO^-$

v. The number of moles of HCOOH in the solution.

$pH = pKa + \log\left(\frac{[salt]}{[acid]}\right)$

$4.37 = -\log(1.8 \times 10^{-4}) + \log(x)$

$4.37 = 3.74 + \log(x)$

$.625 = \log(x)$

$10^{.625} = 10^{\log(x)}$

$x = 4.22 = \frac{[salt]}{[acid]} = \frac{\frac{2.25 \times 10^{-3}}{.04}}{x}$

$x = .0133 M$

$M = \frac{n}{L}$

$.0133 = \frac{n}{.04}$

$n = 5.33 \times 10^{-4}$ moles

b. At equivalence, will the pH be greater than 7, less than 7, or equal to 7. Explain your reasoning.

Greater than 7. A weak acid/strong base titration will have an equivalence point above 7. The conjugate base of the weak acid affects the pH at equivalence.

656. Determine the volume of each solution needed to reach equivalence.

a. What volume of 0.15 M HCl is needed to reach equivalence with 23.0 mL of 0.15 M NaOH?

23.0 mL

b. What volume of 0.23 M Benzoic acid ($K_a = 6.3 \times 10^{-5}$) is needed to reach equivalence with 32.92 mL of 0.23 M NaOH?

32.92 mL

c. What volume of 0.10 M HCl is needed to reach equivalence with 20 mL of 0.20 M NaOH?

40 mL

657. What is the pH at equivalence of a …

a. strong acid/strong base titration. Explain why.

7. only water and a neutral salt remain in solution.

b. weak acid/strong base titration. Explain why.

Above 7. At equivalence the conjugate base of the weak acid remains. The conjugate base creates a basic solution.

c. weak base/strong acid titration. Explain why.

Below 7. At equivalence the conjugate acid of the weak base remains. The conjugate acid creates an acidic solution

Review Topic 3.2

658. What are the properties of an ionic bond?

Ionic compounds are held together with strong interactions between ions. They have low vapor pressures, high melting points and high boiling points. Ionic compounds are brittle and can conduct electricity when aqueous or molten.

659. What types of atoms makeup covalent network solids?

Only nonmetal.

660. Give the name of four covalent network solids given in the CED.

diamond, graphite, silicon dioxide, and silicon carbide

661. What are the properties of covalent network solids?

Covalent network solids are held together with covalent bonds. Covalent solids have high melting points, are rigid, and hard.

662. What are the properties of molecular solids?

They are attracted to each other through IMF. They generally have a low melting point, do not conduct electricity, and can be composed of very large molecules.

663. What are the properties of metallic solids?

Metallic solids are held together with metallic bonds. They are good conductors of electricity and heat, are malleable, and ductile.

664. What is the relationship between conductivity and number of ions in solution?

As the number of ions in solution increases the conductivity increases.

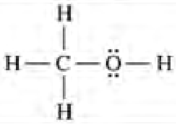
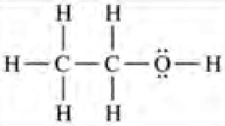
665. What is the relationship between strength of intermolecular force and …

a. Boiling point?

As strength of IMF increases the boiling point increases.

b. Vapor pressure?

As strength of IMF increases the vapor pressure decreases.

Substance	Lewis Diagram	Boiling Point
CH_3OH		338 K
C_2H_5OH		???

666. Answer the following questions about methanol and ethanol shown above.

a. Would you expect the boiling point of C_2H_5OH to be greater than, less than, or equal to the boiling point of CH_3OH? Explain your reasoning.

I would expect the boiling point of C_2H_5OH to be greater than the boiling point of CH_3OH. They both have the same IMF present but C_2H_5OH would have stronger London dispersion forces due to a greater number of total electrons. C_2H_5OH is more polarizable as a result.

b. At 25 °C the vapor pressure of C_2H_5OH is less than the vapor pressure of CH_3OH. What intermolecular force is most responsible for this difference? Justify your answer by referring to each molecule.

London dispersion forces are responsible for the difference. They both have the same types of IMF present - London dispersion, dipole-dipole, and hydrogen bonding. Since C_2H_5OH has more electrons it is more polarizable and has stronger London dispersion forces as a result.

Section 8.05

Topic 8.8, 8.10 Worksheet

667. What types of substances make up a buffer?

Weak acid and salt (conjugate base of a weak acid) make up an acidic buffer and weak base and salt (conjugate acid of the weak base) make up a basic buffer.

668. What does the conjugate acid react with in a buffer?

An added base.

669. What does the conjugate base react with in a buffer?

An added acid.

670. How is a buffer able to resist a change in pH?

For an acidic buffer, any acid added will react with the conjugate base and any base added will react with the weak acid.
For a basic buffer, any acid added will react with the weak base and any base added will react with the conjugate acid.

671. When is a buffer formed in a titration?

As soon as base/acid has been added to the weak acid/base, up to equivalence.

672. Consider a buffer made from acetic acid and sodium acetate.

a. Explain how to make a buffer using acetic acid, $HC_2H_3O_2$, and sodium acetate, $NaC_2H_3O_2$.

A solution made of equal molarities of each. Add $NaC_2H_3O_2$ to $HC_2H_3O_2$.

b. Does the acetic acid react with an acid or a base? Give the reaction of acetic acid reacting with the acid/base. In the case of an acid use HCl and in the case of a base use NaOH.

Base $HC_2H_3O_2(aq) + OH^-(aq) \rightarrow C_2H_3O_2^-(aq) + H_2O(l)$

c. Does the sodium acetate react with an acid or a base? Give the reaction of acetic acid reacting with the acid/base. In the case of an acid use HCl and in the case of a base use NaOH.

Acid $C_2H_3O_2^-(aq) + H^+ \rightarrow HC_2H_3O_2(aq)$

673. Which of the following solutions would be considered a buffer?

Substances	Buffer (Yes or No)
0.10 M HCl + 0.10 M NaCl	No
0.10 M HF + 0.10 M NaF	Yes
0.10 M HBr + 0.10 M NaBr	No
0.10 M C_6H_5COOH + 0.10 M KC_6H_5COO	Yes

674. For the buffers in 673 above, is the pH greater than, less than, or equal to the pKa? Explain your reasoning by referring to the Henderson-Hasselbach equation.

pH=pka b/c same molarity

$$pH = pka + \log\left(\frac{[salt]}{[acid]}\right)$$

$$pH = pka + \log(1)$$

$$pH = pka + 0$$

675. Assume 1.00 L of each of the solutions below. Of the buffers created below, place an "X" in the box that has the greatest buffering capacity.

0.1 M NaH_2PO_4 + 0.1 M Na_2HPO_4	
0.01 M NaH_2PO_4 + 0.01 M Na_2HPO_4	
1.0 M NaH_2PO_4 + 1.0 M Na_2HPO_4	X

676. For the buffers created in 675 above, how does the pH change with the change in molarity?

It doesn't. They are all the same molarity for salt and acid

$$pH = pka + \log\left(\frac{[salt]}{[acid]}\right)$$

$\log(1) = 0$ so $pH = pka$

677. The solids listed below are placed in water. Give the dissociation reaction for the solids and then the K_{sp} expression.

a. $AgI(s)$

$AgI_{(s)} \rightleftharpoons Ag^{+}_{(aq)} + I^{-}_{(aq)}$ $\quad K_{sp} = [Ag^{+}][I^{-}]$

b. $CaF_2(s)$

$CaF_{2(s)} \rightleftharpoons Ca^{2+}_{(aq)} + 2F^{-}_{(aq)}$ $\quad K_{sp} = [Ca^{2+}][F^{-}]^2$

c. $CuBr(s)$

$CuBr_{(s)} \rightleftharpoons Cu^{+}_{(aq)} + Br^{-}_{(aq)}$ $\quad K_{sp} = [Cu^{+}][Br^{-}]$

d. $Mg(OH)_2(s)$

$Mg(OH)_{2\,(s)} \rightleftharpoons Mg^{2+}_{(aq)} + 2OH^{-}_{(aq)}$ $\quad K_{sp} = [Mg^{2+}][OH^{-}]^2$

e. $CdCO_3(s)$

$CdCO_{3(s)} \rightleftharpoons Cd^{2+}_{(aq)} + CO_3^{2-}{}_{(aq)}$ $\quad K_{sp} = [Cd^{2+}][CO_3^{2-}]$

678. The solid $Al(OH)_3$ has a K_{sp} of 1.8×10^{-5}.

a. Determine the $[OH^{-}]$ in a saturated solution.

$Al(OH)_{3\,(s)} \rightleftharpoons Al^{3+}_{(aq)} + 3OH^{-}_{(aq)}$

$x \qquad x \qquad 3x$

$K_{sp} = [Al^{3+}][OH^{-}]^3$

$1.8 \times 10^{-5} = (x)(3x)^3$

$1.8 \times 10^{-5} = 27x^4$

$6.67 \times 10^{-7} = x^4$

$x = 2.86 \times 10^{-2}\,M$

$[OH^{-}] = 3x$

$[OH^{-}] = 8.57 \times 10^{-2}\,M$

b. Determine the pH of the saturated solution.

$pOH = -\log[OH^{-}]$

$pOH = 1.07$ $\qquad pH = 12.93$

c. Determine the mass of Al^{3+} dissolved in 50 mL of solution.

$[Al^{3+}] = 2.86 \times 10^{-2}\,M$

$M \times L = n$

$(2.86 \times 10^{-2})(.05) = n_{Al^{3+}}$

$$\frac{.00143 \text{ mols } Al^{3+} \mid 26.98\,g}{\mid 1 \text{ mol}} = 0.0386\,g\,Al^{3+}$$

d. Determine the solubility, in mols L^{-1}, of $Al(OH)_3$.

$[Al(OH)_3] = 2.86 \times 10^{-2}\,M$

679. Lead(II) arsenate, $Pb_3(AsO_4)_2$ is a slightly soluble solid. A saturated solution has $[Pb^{2+}] = 3.3 \times 10^{-8}$ M.

a. Determine the K_{sp} of lead(II) arsenate.

$Pb_3(AsO_4)_2\,(s) \rightleftharpoons 3Pb^{2+}(aq) + 2AsO_4^{3-}(aq)$

$x \quad 3x \quad 2x$

$3.3 \times 10^{-8}M$

$3x = 3.3 \times 10^{-8}M$

$x = 1.1 \times 10^{-8}$

$K_{sp} = (3x)^3(2x)^2$

$K_{sp} = (3.3 \times 10^{-8})^3(2.2 \times 10^{-8})^2$

$K_{sp} = 1.74 \times 10^{-38}$

b. Determine the solubility of lead(II) arsenate.

$[Pb_3(AsO_4)_2] = 1.1 \times 10^{-8}M$

c. Determine the number of moles of arsenate dissolved in 150. mL of a saturated solution.

$[AsO_4] = 2.2 \times 10^{-8}M$

$M \times L = n$

$(2.2 \times 10^{-8})(.15) = n$

$n = 3.3 \times 10^{-9}$ moles

680. Determine the $[Pb^{2+}]$, in mols L^{-1}, if 0.150 M NaF is added to a saturated solution of lead(II) fluoride, PbF_2 (K_{sp} of $PbF_2 = 2.7 \times 10^{-8}$).

$PbF_2\,(s) \rightleftharpoons Pb^{2+} + 2F^-$

$x \quad x \quad 2x$

$.150M$

$K_{sp} = [Pb^{2+}][F^-]$

$2.7 \times 10^{-8} = x(.150)$

$x = 1.8 \times 10^{-7}M = [Pb^{2+}]$

681. Give the net-ionic reaction of HCl(aq) reacting with NaOH(aq).

$H^+_{(aq)} + OH^-_{(aq)} \rightarrow H_2O(l)$

682. Give the net-ionic reaction of $HC_2H_3O_2$(aq) reacting with NaOH(aq).

$HC_2H_3O_2(aq) + OH^- \rightarrow C_2H_3O_2^-(aq) + H_2O(l)$

683. Give the net-ionic reaction of NH_3(aq) reacting with HCl(aq).

$NH_{3(aq)} + H^+_{(aq)} \rightarrow NH^+_{4(aq)} + Cl^-(aq)$

684. For a strong acid/strong base titration, explain how to calculate the pH when …

a. No base has been added.

$pH = -\log[H_3O^+]$

The molarity of the acid is the molarity of H_3O^+

b. Some base has been added but not enough to reach equivalence.

$n_{acid} - n_{base} = n_{acid}$ left over

$pH = -\log\left(\frac{n_{acid\ left\ over}}{total\ volume}\right)$

c. Enough base has been added to reach equivalence.

$pH = 7$

d. Enough base has been added to go beyond equivalence.

$n_{base} - n_{acid} = n_{base}$ left over

$pH = 14 + \log\left(\frac{n_{base\ left\ over}}{total\ volume}\right)$

685. For a weak acid/strong base titration, explain how to calculate the pH when …

a. No base has been added.

$$k_a = \frac{x^2}{[HA]} \qquad \sqrt{k_a \times [HA]} = x \qquad pH = -\log(x)$$

$$x = [H_3O^+]$$

b. Some base has been added but not enough to reach equivalence.

$$pH = pKa + \log\left(\frac{[salt]}{[acid]}\right)$$

c. Enough base has been added so that it is halfway to equivalence.

$$pH = pKa$$

d. Enough base has been added to reach equivalence.

$$\frac{K_w}{Ka} = K_b = \frac{x^2}{[salt]} \qquad \sqrt{k_b \times [salt]} = x \qquad pH = 14 + \log(x)$$

$$x = [OH^-]$$

e. Enough base has been added to go beyond equivalence.

$$pH = 14 + \log\left(\frac{n_{base} - n_{acid}}{\text{total volume}}\right)$$

686. For a strong acid/weak base titration, explain how to calculate the pH when …

a. No acid has been added.

$$K_b = \frac{x^2}{[B]} \qquad \sqrt{K_b \times [B]} = x \qquad pH = 14 + \log(x)$$

$$x = [OH^-]$$

b. Some acid has been added but not enough to reach equivalence.

$$pOH = pK_b + \log\left(\frac{[salt]}{[base]}\right) \qquad pH = 14 - pOH$$

c. Enough acid has been added so that it is halfway to equivalence.

$$pOH = pK_b \qquad pH = 14 - pOH$$

d. Enough acid has been added to reach equivalence.

$$\frac{K_w}{K_b} = Ka = \frac{x^2}{[salt]} \qquad \sqrt{Ka \times [salt]} = x \qquad pH = -\log(x)$$

$$x = [H_3O^+]$$

e. Enough acid has been added to go beyond equivalence.

$$pH = -\log\left(\frac{n_{acid} - n_{base}}{\text{total volume}}\right)$$

Review Topic 3.6

687. Explain why the following particles deviate from ideal gas behavior.

a. Polar gas particles

Stronger forces of attraction than expected

b. Large radius gas particles

Take up a grate volume than expected

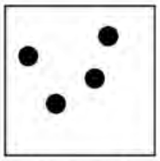
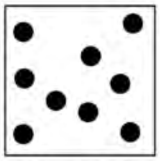
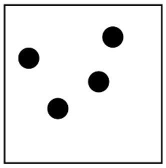
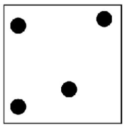
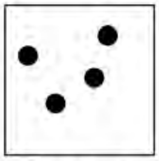
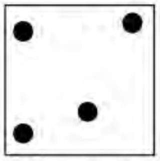
c. Gas particles at low temperature

Stronger forces of attraction then expected

d. Gas particles at high pressure

At high pressure the particles are close to each othe and have higher forces of attraction Than expected.

688. Decide which container would deviate the most from ideal gas behavior. Explain your reasoning.

	Samples	Sample 1 or Sample 2	Reasoning
a.	 Sample 1  Sample 2	Sample 2	higher pressure
b.	 Sample 1  Sample 2	Sample 1	larger particles
c.	 Sample 1 @ 100°C  Sample 2 @ 50°C	Sample 2	lower temperature

Section 8.07

Topic 8.9, 8.7 Worksheet

689. Determine the pH of the buffers given below assuming equimolar concentrations acid/base and salt:

$pH = pKa + \log(1)$

$pH = pKa$

Buffer	K_a of acid or K_b of base	pH
A	$K_a = 1.3 \times 10^{-4}$	3.89
B	$K_a = 5.3 \times 10^{-8}$	7.27
C	$K_b = 7 \times 10^{-3}$	11.85

690. A buffer is created by mixing equal volumes of equimolar weak acid and a salt containing the conjugate base of the weak acid. A little acid or base has been added to change the concentrations of the salt or acid. Does the pH of the buffer increase, decrease, or remain the same when …

a. the concentration of the salt is greater than the concentration of the acid. Explain your reasoning in terms of the Henderson-Hasselbalch equation.

$pH = pKa + \log\left(\frac{[salt]}{[acid]}\right)$

$\left(\frac{[salt]}{[acid]}\right) > 1$

pH of the buffer increases. The ratio ([salt]/[acid]) is greater than 1. The log of a number greater than 1 is a positive number. In the HH equation you would add to the pKa value to determine the pH.

b. the concentration of the acid is greater than the concentration of the salt. Explain your reasoning in terms of the Henderson-Hasselbalch equation.

$\left(\frac{[salt]}{[base]}\right) < 1$

The pH of the buffer decreases. The ratio ([salt]/[acid]) is less than 1. The log of a number less than 1 is a negative number. In the HH equation you would subtract from the pKa value to determine the pH.

c. the concentration of the acid and salt remain in the same ratio. Explain your reasoning in terms of the Henderson-Hasselbalch equation.

The pH remains the same. If the ratio ([salt]/[acid]) remains the same then the pH remains the same because in the HH equation you would take the log of the same value.

691. Determine the pH of a buffer in the following situations. The pKa of the buffer is 3.08.

a. 20 mL of 0.1 M weak acid is mixed with 20 mL of 0.1 M salt.

$pH = pKa + \log(1) \therefore pH = 3.08$

b. 20 mL of 0.1 M weak acid is mixed with 20 mL of 1.0 M salt.

$pH = pKa + \log(10) \quad pH = 4.08$

c. 20 mL of 0.1 M weak acid is mixed with 200 mL of 1.0 M salt.

$pH = pKa + \log(100) \quad pH = 5.08$

d. 200 mL of 0.1 M weak acid is mixed with 20 mL of 0.1 M salt.

$pH = pKa + \log(0.1) \quad pH = 2.08$

692. Determine the molarity of the salt created and the resulting pH for the following situations. In each case, the acid being used is 20 mL of 0.10 M $HC_3H_5O_3$ with a K_a of 8.3×10^{-4}.

a. The acid is mixed with 10 mL of 0.10 M NaOH.

pH = pKa b/c halfway to equivalence

$pH = -\log(8.3 \times 10^{-4}) = 3.081$

b. The acid is mixed with 5 mL of 0.10 M NaOH.

Acid	Base	acid left over	Salt
M×V	M×V	2 − .5	.5 mmols
20 × .1	5 × .1	1.5 mmols	
2 mmols	.5 mmols		

$pH = pKa + \log\left(\frac{[Salt]}{[acid]}\right)$

$pH = -\log(8.3 \times 10^{-4}) + \log\left(\frac{.5}{1.5}\right)$

$pH = 2.604$

c. The acid is mixed with 15 mL of 0.10 M NaOH.

Acid	Base	Acid left over	Salt
M×V	M×V	2 − 1.5	1.5 mmols
20 × .1	15 × .1	.5 mmols	
2 mmols	1.5 mmols		

$pH = pKa + \log\left(\frac{[Salt]}{[acid]}\right)$

$pH = 3.558$

693. Which species is dominant in a buffer, the acid or the conjugate base of the acid, if …

a. the pH < pKa.

acid

b. the pH > pKa.

Conjugate base

c. the pH = pKa.

neither species

Review Topic 4.1

694. What are the signs that a chemical reaction has occurred?

Gives off heat or light.
A precipitate is formed.
A gas is formed.
A permanent change in color.

695. What is the difference between a chemical change and physical change?

A physical change is a change in state, i.e. solid to liquid. No new substances are formed. Only IMF are broken/created.
A chemical change creates new substances. Chemical bonds are broken and/or created.

696. A student uses chromatography to separate a mixture of liquids. Would this be considered a chemical or a physical change? Explain your reasoning.

This would be a physical change. No new substance is formed when they are separated.

697. A student uses distillation to separate a mixture of liquids. Would this be considered a chemical or a physical change?

This would be a physical change. No new substances are formed.

698. A student selects two aqueous solutions that will form a precipitate. Would this be considered a chemical or a physical change?

A chemical change. A new substance, the precipitate, is formed.

699. Is a change in intermolecular forces a physical change or a chemical change? Explain your reasoning.

Physical change. The substances remain the same.

700. Is a change in covalent bonds a physical change or a chemical change? Explain your reasoning.

A chemical change. When covalent bonds are changed new substances are formed.

Unit 9

Section 9.01

Topic 9.1, 9.2 Worksheet

701.Place an “X” in the box for the system with higher entropy.

☐ $NaCl(s)$
Or
☒ $NaCl(aq)$

☐ $H_2O(l)$
Or
☒ $H_2O(g)$

☒ $H_2O(l)$ at 50 °C
Or
☐ $H_2O(l)$ at 25 °C

☐ $CO_2(s)$
Or
☒ $CO_2(g)$

☐ 1 mol of gas in a 1.0 L container
Or
☒ 1 mol of gas in a 5.0 L container

☐ 1.0 L of $CH_4(g)$
Or
☒ 5.0 L of $CH_4(g)$

☐ 1.0 M $C_{12}H_{22}O_{11}$
Or
☒ 10.0 M $C_{12}H_{22}O_{11}$

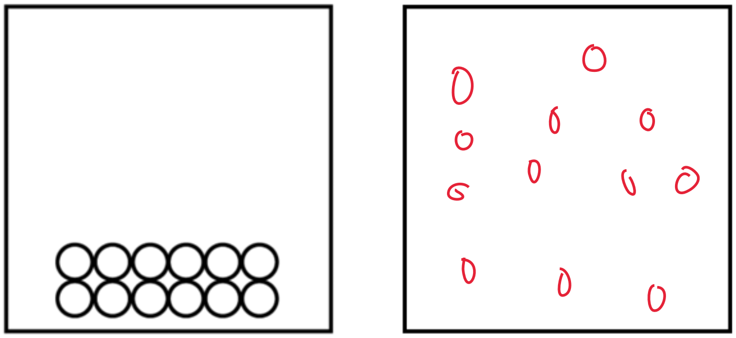
702.Determine if ΔS would be positive, negative, or near zero for the following reactions. Place an “X” in the appropriate box.

Reaction	ΔS		
	+	0	−
$2\ H_2S(g) + SO_2(g) \rightleftharpoons 3\ S(s) + H_2O(g)$			X
$3\ Ag(s) + 4\ HNO_3(aq) \rightarrow 3\ AgNO_3(aq) + NO(g) + 2\ H_2O(l)$	X		
$H_2O(l) \rightarrow H_2O(s)$			X
$I_2(s) + \frac{1}{2}\ Cl_2(g) \rightleftharpoons ICl(g)$	X		
$CO_2(g) + 2\ NH_3(g) \rightarrow CO(NH_2)_2(s) + H_2O(l)$			X
$Cl_2(g) \rightarrow Cl_2(l)$			X
$Mg^{2+}(aq) + 2\ OH^-(aq) \rightarrow Mg(OH)_2(s)$			X
$C_3H_8(g) + 5\ O_2(g) \rightarrow 3\ CO_2(g) + 4\ H_2O(g)$	X		
$Pb(NO_3)_2(s) + 2\ KI(s) \rightarrow PbI_2(s) + 2\ KNO_3(s)$		X	

703.Draw particle diagrams for the following situations.

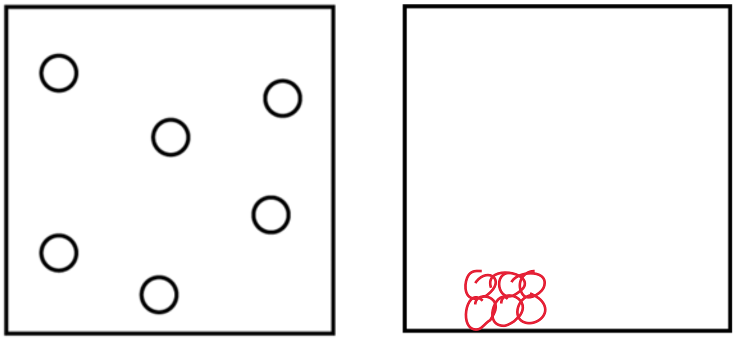
Answers may vary

Draw a situation where ΔS is +.



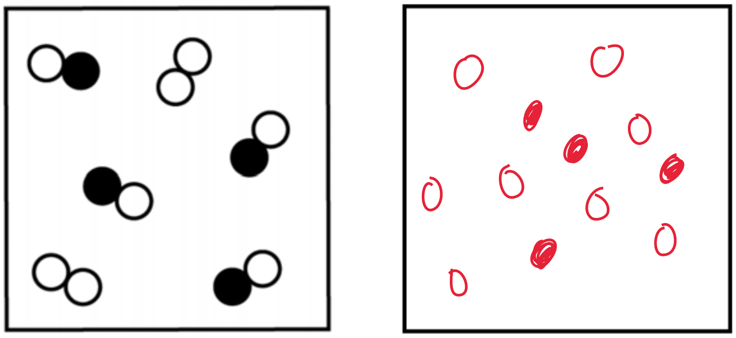
Before After

Draw a situation where ΔS is –



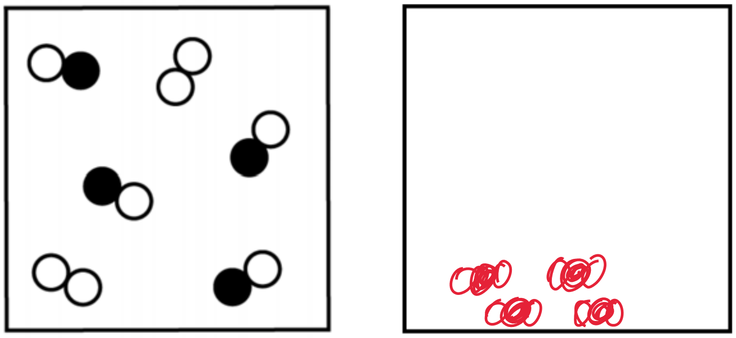
Before After

Draw a chemical reaction where ΔS is +



Before After

Draw a chemical reaction where ΔS is –



Before After

704. What equation is used to calculate ΔSº? What are the units for ΔSº?

$\Delta S^\circ = \Sigma \Delta S^\circ_{prod} - \Sigma \Delta S^\circ_{rct}$ $\Delta S = J/mol_{rxn}$

705. Use the entropy values given to calculate ΔSº values for the following reactions.

a. $C_2H_4(g) + H_2(g) \rightarrow C_2H_6(g)$

Compound	Sº(J/mol•K)
$C_2H_4(g)$	219.4
$H_2(g)$	130.58
$C_2H_6(g)$	229.5

$\Delta S^\circ = (C_2H_6) - (C_2H_4 + H_2)$

$= 229.5 - (219.4 + 130.58)$

$\Delta S^\circ = -120.48\ J/mol_{rxn}$

b. $Be(OH)_2(s) \rightarrow BeO(s) + H_2O(g)$

Compound	Sº(J/mol•K)
$Be(OH)_2(s)$	50.21
$BeO(s)$	13.77
$H_2O(g)$	188.83

$\Delta S^\circ = (13.77 + 188.83) - (50.21)$

$\Delta S^\circ = 152.39\ J/mol_{rxn}$

c. $2\ CH_3OH(g) + 3\ O_2(g) \rightarrow 2\ CO_2(g) + 4\ H_2O(g)$

Compound	Sº(J/mol•K)
$CH_3OH(g)$	237.6
$O_2(g)$	205
$CO_2(g)$	213.6
$H_2O(g)$	188.83

$\Delta S^\circ = (2(213.6) + 4(188.83)) - (2(237.6) + 3(205))$

$\Delta S^\circ = 92.32\ J/mol_{rxn}$

706. Would you expect the entropy of $CH_3OH(l)$ to be greater than, less than, or equal to the entropy of $CH_3OH(g)$? Explain your reasoning.

The entropy of $CH_3OH(l)$ is less than the entropy of $CH_3OH(g)$. A liquid will take fewer microstates than a gas. The fewer microstates the less entropy.

Review Topic 6.4

707. An unknown substance is determined to have a specific heat capacity of 1.22 J/g°C.

a. Determine the amount of heat energy required to raise 1200 g of the substance by 130 °C.

$q = mc\Delta t$

$q = (1200)(1.22)(130)$

$q = 190320\ J$

b. The substance is placed in direct contact with 3.035 L of water (density = 1.00 g/mL) at 25 °C. Determine the temperature of the water after all of the energy has transferred from the substance to the water and both substances have reached thermal equilibrium. Assume the specific heat capacity of water to be 4.18 J/g °C.

$M_{water} = 3035\ mL \times 1.00\ g/mL$

$= 3035\ g$

$q = mc\Delta t$

$190320\ J = (3035)(4.18)\Delta t$

$15.00\ °C = \Delta t$

$\Delta t = t_f - t_i$

$15 = t_f - 25$

$40\ °C = t_f$

c. Determine the molar heat capacity if the molar mass of the substance is 135 g/mol.

$\frac{1.22\ J}{g\ °C} \Big| \frac{135\ g}{1\ mol} = 164.7\ J/mol\ °C$

708. Glycerol, $C_3H_8O_3$, has an enthalpy of combustion, $\Delta H°_{comb}$, of −1650 kJ/mol.

a. Determine the amount of energy released when 15.0 grams of glycerol is combusted.

$\frac{15.0\ g\ C_3H_8O_3}{} \Big| \frac{1\ mol}{92.094} = \frac{0.163\ mols}{} \Big| \frac{-1650\ kJ}{1\ mol} = -269\ kJ$

b. All of the energy released from the combustion of 15.0 grams of glycerol is used to heat 2,500 mL of water. Determine the change in temperature of the water. (Assume the density of water to be 1.00 g/mL and the specific heat capacity of water to be 4.18 J/g °C.)

$q = mc\Delta t$

$269000 = (2500)(4.18)\Delta t$

$25.7\ °C = \Delta t$

709. A student is attempting to melt a 2000 g block of ice such that the initial and final temperature of the block of ice is 0 °C.

a. Determine the amount of energy needed to melt the block of ice if the enthalpy of fusion is 330 J/g.

$q = m\Delta H_f$

$= (2000)(330)$

$q = 660\ 000\ J$

b. The student first uses a 5000 g block of iron (specific heat capacity = 0.451 J/g °C) to melt the ice. To what temperature must the iron block, initially at 25 °C, be raised in order to melt the block of ice?

$q = mc\Delta t$

$660\,000 = (5000)(.451)\Delta t$

$\Delta t = 293\ °C$

$\Delta t = t_f - t_i$

$293°C = t_f - 25°C$

$318\ °C = t_f$

Metal	Specific Heat Capacity (J/g°C)
Copper	0.385
Iron	0.450

c. The specific heat capacities of copper and iron are given above. If copper is raised to the same temperature as the iron block will the mass of copper needed to melt the same block of ice be greater than, less than, or equal to the mass of iron used?

$q_{Fe} = q_{Cu}$

Fe Cu

$mc\Delta t = mc\Delta t$

$mc = mc$

The mass of Cu would have to be greater. Since Cu has a lower specific heat capacity, it takes less energy to raise Cu to the temp of 318°C. As a result the mass would have to be greater to have the same energy as the Fe block.

Section 9.02

Topic 9.3, 9.4 Worksheet

710. What are the conditions for standard state?

1atm, 1M, 25°C

711. What is the sign for ΔG^{o} when …
    a. the reaction is thermodynamically favorable?

    $-$

    b. the reaction is not thermodynamically favorable?

    $+$

712. What does it mean when a reaction is not thermodynamically favorable? Explain your answer in terms of the reaction progress.

A reaction that is not thermodynamically favorable will not proceed in the forward direction. It will proceed in the reverse direction.

713. A reaction has a negative ΔG but does not proceed at an appreciable rate. What could keep the reaction from proceeding even though it is thermodynamically favorable? (Hint: think collision theory)

A large activation energy.

714. What is the formula to calculate $\Delta G^{o}_{reaction}$ when given the ΔG^{o} of the reactants and products?

$$\Delta G^{\circ} = \Sigma \Delta G^{\circ}_{prod} - \Delta G^{\circ}_{rct}$$

715. What is the formula to calculate ΔG^{o} when ΔH^{o} and ΔS^{o} are known?

$$\Delta G^{\circ} = \Delta H^{\circ} - T\Delta S^{\circ}$$

716. Complete the table below by indicating the relative temperature for which the reaction would be considered thermodynamically favorable.

ΔH^{o}	ΔS^{o}	Symbols	$\Delta G^{o} < 0$ at which temperature?
< 0	> 0	< >	all T
> 0	< 0	> <	no T
> 0	> 0	> >	high T
< 0	< 0	< <	low T

717. What is the sign of ΔH …

a. when bonds are broken?

+ endothermic

b. when bonds are formed?

− exothermic

718. Calculate ΔG° for the following reactions given the G° of the reactants and products.

a. $2\,Ag(s) + Cl_2(g) \rightarrow 2\,AgCl(s)$

Compound	ΔG_f^o (kJ/mol)
AgCl(s)	−109.70

$\Delta G^\circ = 2(AgCl) - (2Ag + Cl_2)$

$\Delta G^\circ = 2(-109.7)$

$\Delta G^\circ = -219.4\ kJ/mol_{rxn}$

b. $P_4O_{10}(s) + 16\,H_2(g) \rightarrow 4\,PH_3(g) + 10\,H_2O(g)$

Compound	ΔG_f^o (kJ/mol)
$P_4O_{10}(s)$	−2675.2
$PH_3(g)$	13.4
$H_2O(g)$	−228.57

$\Delta G^\circ = (4(PH_3) + 10(H_2O)) - (P_4O_{10} + 16(H_2))$

$\Delta G^\circ = (4(13.4) + 10(-228.57)) - (-2675.2 + 0)$

$\Delta G^\circ = 443.1\ kJ/ml_{rxn}$

719. Calculate ΔH°, ΔS°, and ΔG° for the reactions given below. Then state if they are thermodynamically favorable. Assume ΔH_f^o and ΔS° do not change with a change in temperature.

a. $NO_2(g) + N_2O(g) \rightarrow 3\,NO(g)$ (carried out at a temperature of 800 K)

Compound	ΔH_f^o (kJ/mol)	ΔS° (J/mol•K)
$NO_2(g)$	33.84	240.45
$N_2O(g)$	81.6	220.0
NO(g)	90.37	210.62

$\Delta H^\circ = 3(NO) - (NO_2 + N_2O)$

$= 3(90.37) - (33.84 + 81.6)$

$\Delta H^\circ = 155.67\ kJ/ml_{rxn}$

$\Delta S^\circ = 3(NO) - (NO_2 + N_2O)$

$= 3(210.62) - (240.45 + 220)$

$\Delta S^\circ = 171.41\ J/ml_{rxn}$

$\Delta G^\circ = \Delta H^\circ - T\Delta S^\circ$

$= (155.67) - ((800)(.17141))$

$\Delta G^\circ = 18.542\ kJ/ml_{rxn}$

The reaction is not thermodynamically favorable b/c ΔG is +.

b. $2\ KClO_3(s) \rightarrow 2\ KCl(s) + 3\ O_2(g)$ (carried out at a temperature of 25 °C)

Compound	ΔH_f^o (kJ/mol)	ΔS^o (J/mol•K)
$KClO_3$	−391.2	143.0
KCl	−435.9	82.7
O_2	0	205

$\Delta H^\circ = (2(-435.9)+0) - (2(-391.2))$

$\Delta H^\circ = -89.4\ kJ/mol_{rxn}$

$\Delta S^\circ = (2(82.7) + 3(205)) - (2(143))$

$\Delta S^\circ = 494.4\ J/mol_{rxn}$

$\Delta G^\circ = -89.4 - (298)(.4944)$

$\Delta G^\circ = -237\ kJ/mol_{rxn}$

The reaction is thermodynamically favorable because ΔG is −.

720. Explain why a reaction may be thermodynamically favorable but not proceed at a measurable rate.

The reaction has a high activation energy. It would be considered under kinetic control.

721. What is the relationship between the magnitude of activation energy and the degree to which a thermodynamically favorable reaction proceeds?

The higher the activation energy the less the reaction will proceed.

$$CaSO_4 \cdot 2\,H_2O(s) \rightleftharpoons CaSO_4(s) + 2\,H_2O(g)$$

722. The hydrate $CaSO_4 \cdot 2\,H_2O(s)$ can be heated to form the anhydrous salt, $CaSO_4(s)$, as shown by the reaction represented above.

a. Using the data in the table below, calculate the value of ΔG°, in kJ/mol_{rxn} , for the reaction at 298 K.

Substance	ΔG_f° at 298 K (kJ/mol)
$CaSO_4 \cdot 2\,H_2O(s)$	−1795.70
$CaSO_4(s)$	−1320.30
$H_2O(g)$	−228.59

$\Delta G^\circ = \Delta G_f^\circ{}_{prod} - \Delta G_f^\circ{}_{rcf}$

$\Delta G^\circ = (-1320.30 + 2(-228.59)) - (-1795.70)$

$\Delta G^\circ = 18.22\ kJ/mol_{rxn}$

b. Given that the value of ΔH° for the reaction at 298 K is $+105\ kJ/mol_{rxn}$, calculate the value of ΔS° for the reaction at 298 K. Include units with your answer.

$\Delta G^\circ = \Delta H^\circ - T\Delta S^\circ$

$18.22 = (105) - (298)(\Delta S^\circ)$

$\Delta S^\circ = -0.291\ kJ/mol_{rxn}$ or $-291\ J/mol_{rxn}$

723. When is thermodynamic favorability determined by …

a. Enthalpy

When dH is negative.

b. Entropy

When dH is positive and dS is negative.

724. For a particular reaction, ΔH = −32 kJ and ΔS = −98 J/K. Assume that ΔH and ΔS do not vary with temperature.

a. At what temperature will the reaction have ΔG = 0?

$\Delta G^\circ = \Delta H^\circ - T\Delta S^\circ$

$0 = -32 - T(-.098)$

$326k = T$

b. If the temperature is increased will the reaction be thermodynamically favored? Explain your reasoning.

If the temperature is increased the reaction will NOT be thermodynamically favored. As T increases the value -T(-S) becomes more and more positive. You are adding an increasingly positive value and causing dG to become more and more positive.

725. Calculate ΔG° for the reaction given below and the indicated temperature.

$2\ PbS(s) + 3\ O_2(g) \rightarrow 2\ PbO(s) + 2\ SO_2(g)$ $\Delta H^\circ = -844\ kJ/mol_{rxn}$ $\Delta S^\circ = -165\ J/K$

a. at a temperature of 0 °C

$\Delta G = \Delta H - T\Delta S$

$\Delta G = -844 - (273(-.165))$

$\Delta G = -799\ kJ/mol_{rxn}$

b. at a temperature of 250 °C

$\Delta G = -844 - (523(-.165))$

$\Delta G = -758\ kJ/mol_{rxn}$

c. at a temperature of 500 °C

$\Delta G = -844 - (773(-.165))$

$\Delta G = -716\ kJ/mol_{rxn}$

d. Does the thermodynamic favorability increase, decrease, or remain the same with an increase in temperature? Justify your answer.

The thermodynamic favorability decreases with an increase in temperature. As the temperature increased from a - c the dG became more positive.

726.Determine at which temperature the reaction is thermodynamically favorable.

Reaction	Sign of ΔH	Sign of ΔS	Temp
A sample of an ionic compound dissolves endothermically into a beaker of water.	+	+	None [✓ High] Low All
2 $H_2(g) + O_2(g) \rightarrow 2\ H_2O(g)$	−	−	None High [✓ Low] All
The exothermic reaction of 4 $Fe(s) + 3\ O_2(g) \rightleftharpoons 2\ Fe_2O_3(s)$	−	−	None High [✓ Low] All
$A(g) + B(g) \rightarrow AB(g)$	−	−	None High [✓ Low] All
2 $A(g) \rightleftharpoons A_2(g)$	−	−	None High [✓ Low] All
$\Delta H < 0$, $\Delta S < 0$	−	−	None High [✓ Low] All
$\Delta H > 0$, $\Delta S < 0$	+	−	[✓ None] High Low All

Review Topic 6.7

727. What is bond enthalpy and why is it always a positive value?

Bond enthalpy is the amount of energy required to break a bond. Bond enthalpy is always positive because energy must be absorbed to break a bond. Bond breaking is endothermic.

728. Which is expected to have a greater bond enthalpy, a C−C or a C=C?

C=C would have a greater bond enthalpy because it's a stronger bond than C-C.

729. A chemical reaction is exothermic, ΔH = −. What does this say about the bond enthalpies of the reactants compared to the bond enthalpies of the products?

The bond enthalpies of the reactants is less than the bond enthalpies of the products.

730. A chemical reaction is endothermic, ΔH = +. What does this say about the bond enthalpies of the reactants compared to the bond enthalpies of the products?

The bond enthalpies of the reactants is greater than the bond enthalpies of the products.

Bond Type	Average Bond Enthalpy (kJ/mol)
C – C	360
C = O	799
C ≡ O	1072
O – O	142
O = O	498

731. Determine the enthalpy of the reaction, ΔH_{rxn}, in kJ/mol_{rxn}, given the bond enthalpies above and the chemical reaction below.

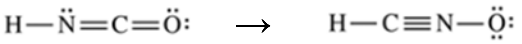
C≡O   O=O   O=C=O

$$2\ CO(g) + O_2(g) \rightarrow 2\ CO_2(g)$$

rct	prod
2 C≡O 2× 1072	4 × C=O 4×799
O=O 498	

2642 − 3196 = ΔH = −554 kJ/mol rxn

732. The reaction to convert isocyanic acid to fulminic acid is shown below:



$H-\ddot{N}=C=\ddot{O}: \rightarrow H-C\equiv N-\ddot{O}:$ $\Delta H = 58.80$ kcal/mol

a. Is the bond enthalpy of the reactants greater than, less than, or equal to the bond enthalpy of the products?

The bond enthalpy of the reactants is less than the bond enthalpy of the products.

b. How does the strength of the bonds of the reactants compare to the strength of the bonds of the products?

The reactant bonds are weaker than the product bonds. Since the bond enthalpy of the products is greater than the bond enthalpy of the reactants it takes more energy to break apart the product than the reactant.

Bond	Enthalpy (kcal/mol)		Bond	Enthalpy (kcal/mol)
N−O	48.04		C=O	178.1
C=N	147.0		C≡N	???
H−C	98.71		H−N	93.45

c. The table above lists the average bond enthalpies for the constituent bonds of isocyanic and fulminic acid. Determine the value of the C≡N bond enthalpy.

rct	prod
H−N = 93.45	H−C = 98.71
N=C = 147	C≡N = ?
C=O = 178.1	N−O = 48.04

$418.55 - (146.75 + x) = 58.80$

$-146.75 - x = -359.75$

$-x = -213$

$x = 213$ kcal/mol.

Section 9.03

Topic 9.6 Worksheet

733. Explain how a reaction that is thermodynamically unfavorable can still occur as a coupled reaction.

A couple reaction can use the overall dG of both reactions to create a thermodynamically set of reactions. So long as the thermodynamically favorable reaction is negative enough to overcome the positive dG of the other reaction the coupled reactions can occur.

734. Calcium carbonate decomposes as shown below:

$$CaCO_3(s) \rightleftharpoons CaO(s) + CO_2(g) \qquad \Delta G^o = 130.40 \text{ kJ/mol}$$

a. Is the decomposition of calcium carbonate endergonic or exergonic?

Endergonic.

b. Is the decomposition of calcium carbonate thermodynamically favorable? Justify your answer.

It is not thermodynamically favorable. The dG is positive.

c. The reaction given above can be coupled with the combustion of carbon, shown below:

$$C(s) + O_2(g) \rightleftharpoons CO_2(g) \qquad \Delta G^o = -394.36 \text{ kJ/mol}$$

Determine the Gibbs free energy for the overall coupled reaction:

$$CaCO_3(s) + C(s) + O_2(g) \rightleftharpoons CaO(s) + 2\ CO_2(g) \qquad \Delta G^o = ???$$

$\Delta G^\circ = 130.40 - 394.36$

$\Delta G^\circ = -263.96$ kJ

d. Is the coupled reaction thermodynamically favorable? Justify your answer.

The coupled reaction is thermodynamically favorable. The dG of the coupled reaction is negative.

$N_2(g) + 3\,F_2(g) \rightarrow 2\,NF_3(g) \qquad \Delta H^\circ_{298} = -264\ \text{kJ mol}^{-1}$

735. Use the equation above to answer the questions that follow.

a. Determine the amount of energy released, in kJ, when 2.80 grams of N_2 reacts with 11.0 grams of F_2.

2.80 g N_2 | 1 mol / 28.02 g = 0.0999 mols N_2 | −264 kJ / 1 N_2 = ~~−26.3 kJ~~

11.0 g F_2 | 1 mol / 38 g = 0.289 mols F_2 | −264 kJ / 3 F_2 = −25.5 kJ

25.5 kJ are released

b. The energy from 735.a above is used to heat 200 g of water originally at 13 °C.

i. Will the temperature of the water increase, decrease, or remain the same? Justify your answer.

The temperature of the water will increase. The reaction is exothermic and gives off energy. ~~As~~ a result the water will absorb the temperature making the temp increase.

ii. Determine the final temperature of the water assuming the specific heat capacity of water is 4.2 J/g °C and there is no loss of energy.

$q = mc\Delta t$

$25500\ \text{J} = (200)(4.2)\Delta t$

$30.4\,^\circ\text{C} = \Delta t$

$\Delta t = t_f - t_i$

$30.4 = t_f - 13$

$43.4\,^\circ\text{C} = t_f$

$$N_2(g) + 3\,F_2(g) \rightarrow 2\,NF_3(g) \qquad \Delta H^o_{298} = -264 \text{ kJ mol}^{-1}$$

c. Which has a greater average bond enthalpy, the reactants or the products? Justify your answer.

The products. The ΔH is – which means the products have a greater bond enthalpy.

d. Use the information below to determine the bond enthalpy of N_2.

Bond	Bond Enthalpy Value (kJ/mol)
N−F	272
F−F	155

rct	prod
$N_2 = ?$	$6 \times N\text{–}F = 6 \times 272$
$3 \times F\text{–}F = 3 \times 155$	

$(x + 465) - 1632 = -264$

$x - 1167 = -264$

$x = 903$ kJ/mol $= N_2$

e. The bond enthalpy of N=N is 418 kJ/mol. Is the N_2 bond a single bond or a triple bond? Justify your answer.

The N_2 bond is a triple bond. The bond enthalpy of 903 is greater than the bond enthalpy of N=N (418). Since the bond enthalpy is greater it must be a stronger bond. $N \equiv N$ is stronger than N=N.

Topic 9.5 Worksheet

736. Is the value of K greater than 1, less than 1, or equal to 1 …
    a. for a thermodynamically favored reaction at equilibrium?

Greater than 1

    b. for a thermodynamically unfavorable reaction at equilibrium?

Less than 1

737. As the sign of ΔG becomes more and more negative, does the equilibrium constant become larger, smaller, or remain the same?

The equilibrium constant becomes larger.

738. What happens to the size of ΔG as K_{eq} …
    a. gets smaller and smaller?

dG becomes more and more positive as K gets smaller.

    b. gets bigger and bigger?

dG becomes more and more negative as K gets bigger.

Note: Don't forget that K only changes with a change in temperature.

739. The reaction for the enthalpy of formation of $FeCl_3(s)$ is shown below:

$$Fe(s) + \frac{3}{2}Cl_2(g) \rightarrow FeCl_3(s)$$

a. Determine the value for the enthalpy of formation, ΔH_f^o in kcal/mol, of $FeCl_3$ from the enthalpy changes of the two reactions shown below:

$$Fe(s) + Cl_2(g) \rightarrow FeCl_2(s) \quad \Delta H^o = -81.7\ kcal/mol_{rxn}$$

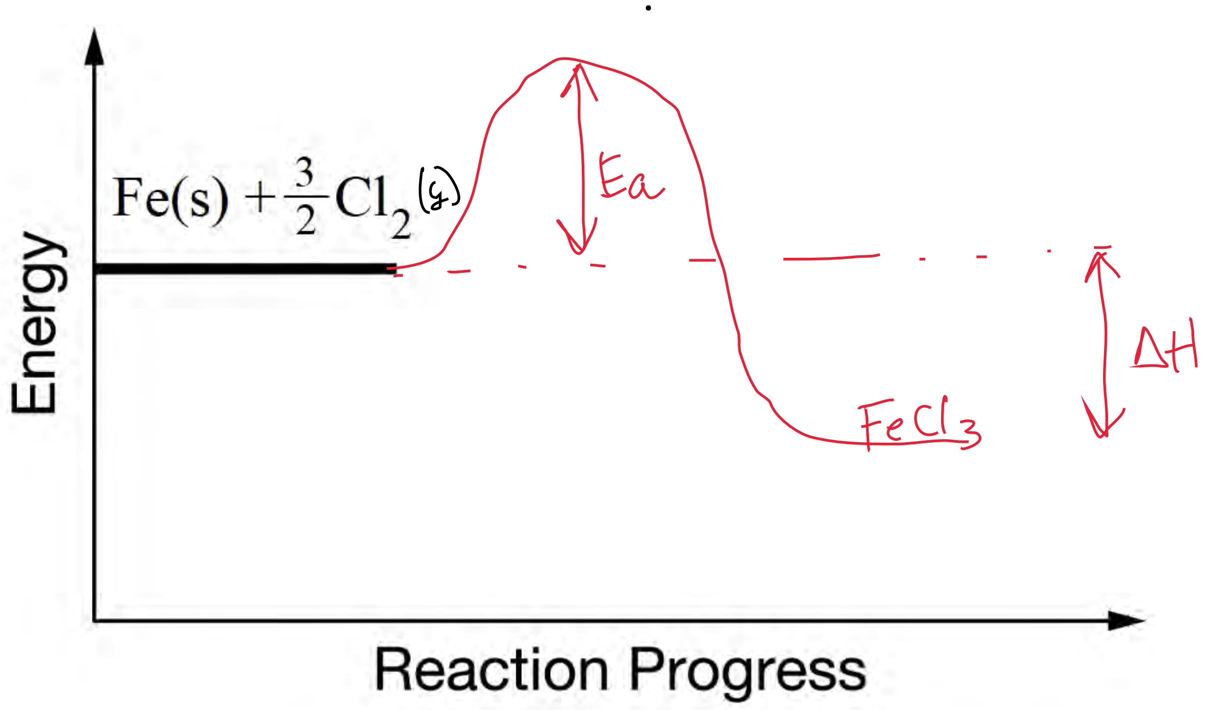
$$FeCl_3(s) \rightarrow FeCl_2(s) + \frac{1}{2}Cl_2(g) \quad \Delta H^o = 13.8\ kcal/mol_{rxn}$$

$$Fe(s) + Cl_2(g) \rightarrow \cancel{FeCl_2(s)} \quad \Delta H = -81.7\ kcal/mol_{rxn}$$

$$\cancel{FeCl_2(s)} + \frac{1}{2}Cl_2(g) \rightarrow FeCl_3(s) \quad \Delta H = -13.8\ kcal/mol_{rxn}$$

$$Fe(s) + \frac{3}{2}Cl_2(g) \rightarrow FeCl_3(s) \quad \Delta H = -95.5\ kcal/mol_{rxn}$$

b. The energy of the reactants is shown on the energy diagram below.
   i. Complete the energy diagram by showing the placement of the product.
   ii. Label the activation energy, E_a, and the ΔH of the reaction.



740. The reaction shown below is known to occur at temperatures above 1200 °C.

$$CO(g) + 2\ H_2S(g) \rightleftharpoons CS_2(g) + H_2O(g) + H_2(g) \qquad \Delta H = ???,\ K_{eq} = ???$$

The reactions given below can be used to calculate ΔH and K_{eq} for the desired reaction above.

$CO(g) + 3\ \cancel{H_2}(g) \rightleftharpoons \cancel{CH_4}(g) + H_2O(g)$ $\Delta H = -205.9$ kJ/mol$_{rxn}$ $K_1 = 9.17 \times 10^{-2}$

$\cancel{CH_4}(g) + 2\ H_2S(g) \rightleftharpoons CS_2(g) + \cancel{4}\ H_2(g)$ $\Delta H = 232.7$ kJ/mol$_{rxn}$ $K_2 = 3.3 \times 10^{4}$

a. Calculate ΔH.

$$\begin{array}{r} -205.9 \\ +232.7 \\ \hline 26.8 \text{ kJ/mol}_{rxn} \end{array}$$

b. Calculate K_{eq}.

$K_{eq} = K_1 \times K_2$

$= (9.17 \times 10^{-2})(3.3 \times 10^{4})$

$K_{eq} = 3026$

Section 9.05

Topic 9.7 Worksheet

741. Describe the role of each part of a Galvanic cell:

a. Anode (electrode)

Loses electrons to the cathode and loses mass (if it participates in the reaction).
Allows electrons to flow to cathode

b. Cathode (electrode)

Gains electrons from the anode and gains mass (if it participates in the reaction).
Gains electrons from anode.

c. Anode solution

Site of oxidation

d. Cathode solution

Site of reduction

e. Salt bridge

Balances charge by allowing anions to flow to the anode and cations to flow to the cathode.

742. For an anode with an electrode that participates in the reaction …

a. does oxidation or reduction occur?

oxidation

b. does the electrode gain or lose mass?

lose mass

c. does the salt bridge contribute cations or anions?

anions

d. do the electrons flow to or away from the electrode?

away

743. For a cathode with an electrode that participates in the reaction …

a. does oxidation or reduction occur?

reduction

b. does the electrode gain or lose mass?

gain mass

c. does the salt bridge contribute cations or anions?

cations

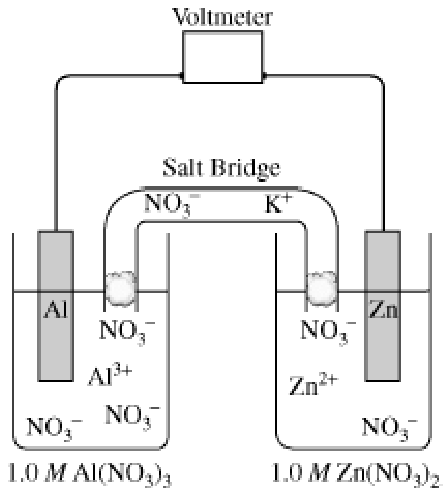
d. do the electrons flow to or away from the electrode?

flow to

744. A galvanic cell uses platinum as an inert electrode for the anode. Does the mass of the electrode increase, decrease, or remain the same as the reaction proceeds?

It remains the same. The Pt electrode will not participate in the reaction.

745. The voltaic cell shown below is thermodynamically favorable.



Reduction Reaction	Voltage
$Al^{3+}(aq) + 3\,e^- \rightarrow Al(s)$	−1.66 V
$Zn^{2+} + 2\,e^- \rightarrow Zn(s)$	−0.763

a. Determine the voltage of the voltaic cell.

Flip the Al^{3+} rxn + 1.66V

$$\begin{array}{r} +1.66V \\ + -\underline{.763V} \\ 0.897V = E^\circ \end{array}$$

b. Determine the net-ionic reaction. $2\times\ Al(s) \rightarrow Al^{3+} + 3e^-$

$$3\times\ \underline{Zn^{2+} + 2e^- \rightarrow Zn(s)}$$

$$2Al(s) + 3Zn^{2+} \rightarrow 2Al^{3+} + 3Zn(s)$$

c. How many moles of electrons are transferred during the chemical reaction?

6 moles of e^-

d. Which metal, Al(s) or Zn(s), is used in the anode?

Al(s)

e. Is the value for the standard free energy change, ΔG°, positive, negative, or zero. Justify your answer.

dG would be negative. A + voltage is thermodynamically favorable. A - dG is thermodynamically favorable.

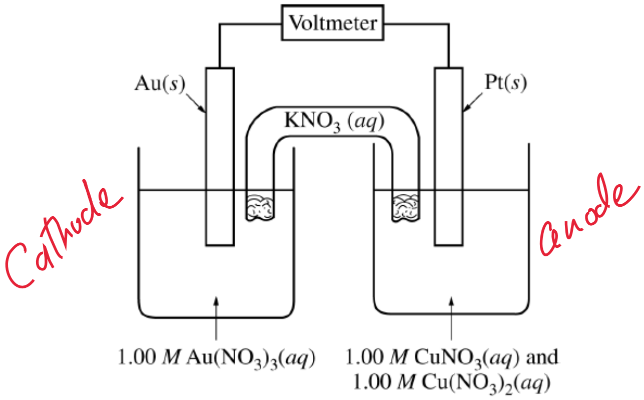
f. If the concentration of Zn^{2+} was changed from 1.0 M to 0.01 M, would the cell voltage increase, decrease, or remain the same? Explain your answer.

The cell voltage would decrease as Q becomes larger than K.

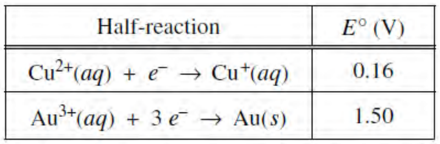
g. What would happen to the voltage if the salt bridge was removed? Explain why.

The voltage would drop to zero. There are no cations/anions to balance out the buildup of charge due to the flow of electrons.

746. The galvanic cell shown below is constructed and analyzed. The standard free-energy change of the cell is negative. Answer the questions that follow.



The standard reduction potentials associated with the cell are given in the following table.



Half-reaction	$E°$ (V)
$Cu^{2+}(aq) + e^- \rightarrow Cu^+(aq)$	0.16
$Au^{3+}(aq) + 3e^- \rightarrow Au(s)$	1.50

a. Determine the value of the standard cell potential, $E°$.

Flip the Cu^{2+} rxn $\quad -0.16V$

$$\begin{array}{r} -0.16V \\ +\ 1.50V \\ \hline 1.34V = E° \end{array}$$

b. Give the net-ionic reaction for the cell. $3\times$ $Cu^+ \rightarrow Cu^{2+} + e^-$

$$\begin{array}{l} Au^{3+} + 3e^- \rightarrow Au(s) \\ \hline 3Cu^+ + Au^{3+} \rightarrow 3Cu^{2+} + Au(s) \end{array}$$

c. Label which half-cell is the cathode and which is the anode.

d. How many moles of electrons are transferred as the reaction proceeds?

3 moles of e^-

e. Does the mass of the Pt(s) electrode increase, decrease, or remain the same as the cell operates? Justify your answer.

Remain the same. The Pt electrode does not participate in the reaction.

f. Would the voltage of the cell increase, decrease, or remain the same if the mass of the Au(s) electrode was increased?

It would remain the same. The mass/size of the electrode does not change the overall voltage of the cell.

g. Calculate ΔG°, in kJ/mol$_{rxn}$, for the cell.

$$\Delta G^\circ = -nFE^\circ$$

$$= -(3)(96485)(1.34\,\overset{V = \frac{J}{C}}{V})$$

$$\Delta G^\circ = -387870\ \text{J}/\text{m}_{\text{rxn}} = -388\ \text{kJ}/\text{mol}_{\text{rxn}}$$

747. In an electrolytic cell …

a. will the more positive or more negative reduction reaction occur at the cathode?

The more positive

b. will the more positive or more negative oxidation reaction occur at the anode?

The more positive

748. In each electrolyte cell below, determine which reaction will occur at the anode and which reaction will occur at the cathode by placing an “X” in the appropriate box.

a. Electrolysis of NaCl(aq)

Reaction	Voltage	Occurs at the cathode	Occurs at the anode
$Na^{+}(aq) + e^{-} \rightarrow Na(l)$	$E^{o} = -2.71$ volts	☐	☐
$2H_2O(l) + 2\ e^{-} \rightarrow H_2(g) + 2\ OH^{-}(aq)$	$E^{o} = -0.83$ volts	☒	☐
$Cl^{-}(aq) \rightarrow \frac{1}{2}\ Cl_2(g) + e^{-}$	$E^{o} = -1.359$ volts	☐	☐
$H_2O(l) \rightarrow \frac{1}{2}\ O_2(g) + 2\ H^{+}(aq) + 2\ e^{-}$	$E^{o} = -1.23$ volts	☐	☒

b. Electrolysis of LiCl(aq)

Reaction	Voltage	Occurs at the cathode	Occurs at the anode
$Li^{+}(aq) + e^{-} \rightarrow Li(s)$	$E^{o} = -3.05$ volts	☐	☐
$2H_2O(l) + 2\ e^{-} \rightarrow H_2(g) + 2\ OH^{-}(aq)$	$E^{o} = -0.83$ volts	☒	☐
$Cl^{-}(aq) \rightarrow \frac{1}{2}\ Cl_2(g) + e^{-}$	$E^{o} = -1.359$ volts	☐	☐
$H_2O(l) \rightarrow \frac{1}{2}\ O_2(g) + 2\ H^{+}(aq) + 2\ e^{-}$	$E^{o} = -1.23$ volts	☐	☒

c. Electrolysis of $CuBr_2$(aq)

Reaction	Voltage	Occurs at the cathode	Occurs at the anode
$Cu^{2+}(aq) + 2e^{-} \rightarrow Cu(s)$	$E^{o} = 0.34$ volts	☒	☐
$2H_2O(l) + 2\ e^{-} \rightarrow H_2(g) + 2\ OH^{-}(aq)$	$E^{o} = -0.83$ volts	☐	☐
$2\ Br^{-}(aq) \rightarrow Br_2(g) + 2\ e^{-}$	$E^{o} = -1.07$ volts	☐	☒
$H_2O(l) \rightarrow \frac{1}{2}\ O_2(g) + 2\ H^{+}(aq) + 2\ e^{-}$	$E^{o} = -1.23$ volts	☐	☐

749. Will a higher molarity voltaic cell operate for a longer time, a shorter time, or the same amount of time as a lower molarity voltaic cell? Explain your reasoning.

It will operate for a longer period of time. A higher molarity means there are more particles to undergo a redox reaction which means the cell will operate a longer period of time.

750. Give the line notation of …

a. #745

$Al_{(s)} | Al^{3+}_{(aq)} || Zn^{2+}_{(aq)} | Zn_{(s)}$

b. #746

$Pt_{(s)} | Cu^{+}, Cu^{2+} || Au^{3+} | Au_{(s)}$

Review Topic 6.8

751. Propane and butane are two different organic compounds that can be used for heating. The combustion of each is shown below.

Combustion of Propane	$C_3H_8(g) + 5\ O_2(g) \rightarrow 3\ CO_2(g) + 4\ H_2O(g)$
Combustion of Butane	$2\ C_4H_{10}(g) + 13\ O_2(g) \rightarrow 8\ CO_2(g) + 10\ H_2O(g)$

a. Determine the enthalpy of reaction, ΔH_{rxn}, in kJ/mol$_{rxn}$, for the combustion of propane using the heats of formation given below.

Substance	$C_3H_8(g)$	$CO_2(g)$	$H_2O(g)$
Standard Enthalpy of Formation (kJ/mol)	−103.85	−393.5	−241.82

$\Delta H = (3(CO_2) + 4(H_2O)) - ((C_3H_8) + 5(O_2))$

$= (3(-393.5) + 4(-241.82)) - (-103.85)$

$= -2147.78 + 103.85$

$\Delta H = -2044\ kJ/mol_{rxn}$

b. Determine the enthalpy of reaction, ΔH_{rxn}, in kJ/mol$_{rxn}$, for the combustion of butane using the heats of formation given below.

Substance	$C_4H_{10}(g)$	$CO_2(g)$	$H_2O(g)$
Standard Enthalpy of Formation (kJ/mol)	−124.73	−393.5	−241.82

$\Delta H = (8(CO_2) + 10(H_2O)) - (2(C_4H_{10}) + 13(O_2))$

$= (8(-393.5) + 10(-241.82)) - (2(-124.73) + 13(0))$

$= -5566.2 + 249.46$

$= -5317\ kJ/mol_{rxn}$

c. Determine which gas, propane or butane, delivers more energy per gram. Justify your answer with a calculation.

propane

$$\frac{-2044\text{ kJ}}{1\text{ mol}_{rxn}}\left|\frac{1\text{ mol}_{rxn}}{1\ C_3H_8}\right|\frac{1\text{ mol}}{44.094\text{ g}} = -46.4\text{ kJ/g}$$

butane

$$\frac{-5317\text{ kJ}}{1\text{ mol}_{rxn}}\left|\frac{1\text{ mol}_{rxn}}{2\ C_4H_{10}}\right|\frac{1\text{ mol}}{58.12\text{ g}} = -45.7\text{ kJ/g}$$

Propane delivers more energy per gram. Propane gives off 46.4 kJ while butane gives off 45.7 kJ

d. If not enough oxygen is present propane will undergo incomplete combustion, forming CO(g) in addition to $CO_2(g)$ and $H_2O(g)$, as shown below. The standard enthalpy of formation of CO(g) is -99.0 kJ/mol.

$$2\ C_3H_8(g) + 9\ O_2(g) \rightarrow 4\ CO_2(g) + 8\ H_2O(g) + 2\ CO(g)$$

Determine the ΔH_{rxn} for the incomplete combustion of propane, in kJ/mol$_{rxn}$,

$$\Delta H = \left(4(CO_2) + 8(H_2O) + 2(CO)\right) - \left(2(C_3H_8) + 9(O_2)\right)$$

$$= \left(4(-393.5) + 8(-241.82) + 2(-99)\right) - \left(2(-103.85) + 9(0)\right)$$

$$= -3706.56 + 207.7$$

$$\Delta H = -3499\text{ kJ/mol}_{rxn}$$

e. Determine the amount of energy per gram of propane for the incomplete combustion.

$$\frac{-3499\text{ kJ}}{1\text{ mol}_{rxn}}\left|\frac{1\text{ mol}_{rxn}}{2\ C_3H_8}\right|\frac{1\text{ mol}}{44.094\text{ g}} = -39.7\text{ kJ/g}$$

Section 9.06

Topic 9.8 Worksheet

752. What must be true for a thermodynamically favorable electrochemical cell …

a. In terms of voltage?

Voltage must be +

b. In terms of Gibbs Free Energy?

ΔG must be −

753. What must be done to run an electrochemical cell that is thermodynamically unfavorable?

Enegy must be put into the cell

754. Consider the reduction potentials given below.

Standard Reduction Half-Reaction	Standard Reduction Potential, E° (volts)
$Al^{3+} + 3\ e^{-} \rightarrow Al(s)$	−1.68
$ZnO_2(s) + 4\ H^{+} + 4\ e^{-} \rightarrow Zn(s) + 2\ H_2O(l)$	−1.47
$V^{2+} + 2\ e^{-} \rightarrow V(s)$	−1.13
$PtCl_4^{2-} + 2\ e^{-} \rightarrow Pt(s) + 4\ Cl^{-}$	0.73

Determine the balanced chemical reaction and the value of E° for a thermodynamically favorable standard galvanic cell made with …

a. Al/Al^{3+} and V/V^{2+} half cells.

2× $Al(s) \rightarrow Al^{3+} + 3e^{-}$ 1.68 V

3× $V^{2+} + 2e^{-} \rightarrow V(s)$ −1.13V

0.55V

$2\ Al(s) + 3V^{2+} \rightarrow 2Al^{3+} + 3V(s)$

b. $Pt/PtCl_4^{2-}$ and ZnO_2/Zn half cells.

$Zn(s) + 2\ H_2O(l) \rightarrow ZnO_2(s) + 4H^{+} + 4e^{-}$ +1.47V

2× $PtCl_4^{2-} + 2e^{-} \rightarrow Pt(s) + 4Cl^{-}$ +0.73V

$Zn(s) + 2H_2O(l) + 2PtCl_4^{2-} \rightarrow ZnO_2(s) + 4H^{+} + 2Pt(s) + 8Cl^{-}$ 2.20V

755. For the cell in 754.a above, does the mass of Al(s) increase, decrease, or remain the same. Justify your answer by referring to the chemical reaction.

The mass of Al(s) would decrease. As the reaction proceeds Al(s) is conveted to Al^{3+} therefore Al(s) mass is lost.

756. For the cell in 754.b above, does the mass of Pt(s) increase, decrease, or remain the same. Justify your answer by referring to the chemical reaction.

The mass of Pt(s) increases. As the reaction proceeds more and more Pt(s) is formed which increases the mass of Pt(s)

757. Determine Gibbs Free Energy, ΔGº, for the cell made in

a. 754.a

$$\Delta G^\circ = -nFE^\circ$$

$$= -(6)(96485)(.55)$$

$$\Delta G^\circ = -318400\,J/mol_{rxn} \quad \text{or} \quad -318\,kJ/mol_{rxn}$$

g. 754.b

$$\Delta G^\circ = -nFE^\circ$$

$$= -(4)(96485)(2.20)$$

$$\Delta G^\circ = -849068\,J/ml_{rxn} \quad \text{or} \quad -849\,kJ/mol_{rxn}$$

758. Which cell would have a greater value of K at 25°C , the cell made in 754.a or the cell made in 754.b? Support your answer with a calculation.

The cell in 754b.

754a

$$\Delta G^\circ = -RT\ln K$$

$$-318 = -\left(\frac{8.314}{1000}\right)(298)\ln K$$

$$\frac{-318}{-2.48} = \frac{-(2.48)\ln K}{-2.48}$$

$$e^{128.4} = e^{\ln K}$$

$$K = 5.53 \times 10^{55}$$

754b

$$\Delta G^\circ = -RT\ln K$$

$$-849 = -\left(\frac{8.314}{1000}\right)(298)\ln K$$

$$-849 = -2.48\ln K$$

$$e^{342.7} = e^{\ln K}$$

$$K = 6.80 \times 10^{148}$$

759. As Gibbs Free Energy, ΔG, decreases (becomes more negative), does the value of K increase, decrease, or remain the same?

K increases

760. Explain the relationship between thermodynamic favorability and the size of K.

As K increases the thermodynamic favorability increases

$$CaCO_3(s) + 2\,H^+(aq) \rightarrow Ca^{2+}(aq) + H_2O(l) + CO_2(g)$$

761. Refer to the equation above. A student uses several small pieces of $CaCO_3(s)$ with a total mass of 5.00 g and excess 1.0 M $HCl(aq)$ and measures the mass loss over time.

a. Refer to the chemical reaction and explain why the beaker loses mass over time?

$CO_2(g)$ is created in the reaction. Since the $CO_2(g)$ isn't captured mass is lost.

b. Would the following conditions increase, decrease, or have no effect on the initial rate of production of $CO_2(g)$? In each case explain your reasoning.

i. Using a large, single 5.00 g chunk of $CaCO_3(s)$ instead of several lumps.

Decrease the rate. A large chunk would have less surface area than several lumps. Less surface area means fewer collisions.

ii. Using 5.00 g of $CaCO_3$ powder instead of several lumps.

Increase the rate. Reasoning is opposite of bi.

iii. Using excess 0.50 M $HCl(aq)$ instead of 1.0 M $HCl(aq)$.

Decrease the rate. A lower molarity will have fewer reactants which means fewer collisions.

iv. Increasing the temperature of the 1.0 M $HCl(aq)$.

Increase the rate. As temperature increases the number of particles w/ the required energy to start the reaction increases.

v. Carrying out the reaction in an ice-water bath.

Decrease the rate. Reasoning is opposite of b iv.

c. In each case from 761.b above, the mass loss remains the same. Explain why the mass loss remained the same.

The rate of the reaction (how fast or slow) has no effect on how much product is formed.

d. A student makes the claim that a catalyst would increase the mass loss in the experiment. Do you agree or disagree with the student? Explain your reasoning.

A catalyst only speeds up the rate of a reaction. A catalyst has no effect on how much product is formed.

762. Use the equation below to answer the questions that follow:

$$5\ H_2O_2(aq) + 2\ MnO_4^-(aq) + 6\ H^+(aq) \rightarrow 2\ Mn^{2+}(aq) + 8\ H_2O(l) + 5\ O_2(g)$$

a. Is the initial rate of disappearance of $MnO_4^-(aq)$ greater than, less than, or equal to the rate of the disappearance of $H_2O_2(aq)$? Justify your answer by referring to the balanced chemical equation.

The rate of disappearance of MnO_4^- is less than the rate of disappearance of H_2O_2. For every 2 moles of MnO_4^- that disappears 5 moles of H_2O_2 disappear.

b. The initial rate of appearance of Mn^{2+} is 3.00×10^{-4} M/s.

i. Determine the initial rate of appearance of $O_2(g)$ over the same time period.

$\frac{1}{2} Rate_{Mn^{2+}} = \frac{1}{5} Rate_{O_2}$

$Rate_{O_2} = 7.5 \times 10^{-4}$ M/s

$\frac{1}{2}(3.00 \times 10^{-4}) = \frac{1}{5} Rate_{O_2}$

ii. Determine the initial rate of disappearance of $H^+(aq)$ over the same time period.

$\frac{1}{2} Rate_{Mn^{2+}} = \frac{1}{6} Rate_{H^+}$

$Rate_{H^+} = 9 \times 10^{-4}$ M/s

$\frac{1}{2}(3.00 \times 10^{-4}) = \frac{1}{6} Rate_{H^+}$

iii. Would the initial rate of disappearance of $H_2O_2(aq)$ be greater than, less than, or equal to the rate of appearance of $O_2(g)$?

The rates would be the same

$E = E^{\circ}_{cell} - \frac{RT}{nF} \ln Q$

763. For the following questions compare the standard voltage of the cell to the new voltage under the conditions specified.

a. Does the voltage of a cell increase, decrease, or remain the same when the concentration of a reactant is increased? Justify your answer in terms of Q.

$E = E^{\circ}_{cell} - \frac{RT}{nF} \ln Q$

$E = 2 - (1)\ln(5) = .39V$

$E = 2 - (1)\ln(2) = 1.3V$

Q would be smaller. As Q gets smaller the ln Q gets smaller and eventually becomes negative. As a result the voltage increases.

b. Does the voltage of a cell increase, decrease, or remain the same when the concentration of a product is increased? Justify your answer in terms of Q.

Q would be larger. As Q gets bigger the ln Q gets bigger. As a result the voltage decreases.

c. Does the voltage of a cell increase, decrease, or remain the same when the volume of a solid electrode is increased? Justify your answer in terms of Q.

The voltage would remain the same. Solids do not appear in Q and therefore do not effect the voltage

d. Does the voltage of a cell increase, decrease, or remain the same when the molarity of reactants and products is increased to 2.0 M from 1.0 M? Justify your answer in terms of Q.

Remains the same. The value of Q remains the same.

$$E_{cell} = E^{o}_{cell} - \left(\frac{RT}{nF}\right) lnQ$$

$2Al(s) + 3Zn^{2+} \rightarrow 2Al^{3+} + 3Zn \quad E^{o} = .897\ V$

764. Use the Nernst equation, given above, to calculate the new voltage at 298 K of the cell made in 745.a when …

a. the [Al^{3+}] = 0.500 M and the [Zn^{2+}] = 1.00 M.

$E_{cell} = .897 - \left(\frac{(8.314)(298)}{(6)(96485)}\right) ln\left(\frac{.5}{1}\right)$

$Q = .5$; voltage increased

$E_{cell} = .897 - (-2.97 \times 10^{-3}) = 0.90\ V$

b. the [Al^{3}] = 1.00 M and the [Zn^{2+}] = 0.500 M.

$E_{cell} = .897 - \left(\frac{(8.314)(298)}{(6)(96485)}\right) ln\left(\frac{1}{.5}\right)$

$Q = 2$; voltage decreased

$E_{cell} = .897 - 2.9 \times 10^{-3} = 0.894\ V$

$3Cu^{+} + Au^{3+} \rightarrow 3Cu^{2+} + Au(s) \quad E^{o} = 1.34\ V$

765. Use the Nernst equation, given above, to calculate the new voltage at 298 K of the cell made in 746.a when …

a. the [Au^{3+}] = 2.00 M, [Cu^{+}] = 1.00 M, and the [Cu^{2+}] = 2.00 M

$E_{cell} = 1.34 - \left(\frac{(8.314)(298)}{(3)(96485)}\right) ln\left(\frac{2}{(1)(2)}\right)$

$Q = 1$

$E_{cell} = 1.34 - 0 = 1.34\ V$

b. the [Au^{3+}] = 2.00 M, [Cu^{+}] = 1.00 M, and the [Cu^{2+}] = 0.500 M

$E_{cell} = 1.34 - \left(\frac{(8.314)(298)}{(3)(96485)}\right) ln\left(\frac{.5}{(1)(2)}\right)$

$Q = .25$; voltage increased

$E_{cell} = 1.34 - (-0.0119) = 1.35\ V$

766. Does the voltage of a cell increase, decrease, or remain the same as the size of Q increases? Support your answer by referring to the answers from 764 and 765.

The voltage of a cell decreases as the size of Q increases. In 764 Q went from 0.5 to 2 and the voltage dropped from 0.90 V to 0.894 V. The same can be said for 765.

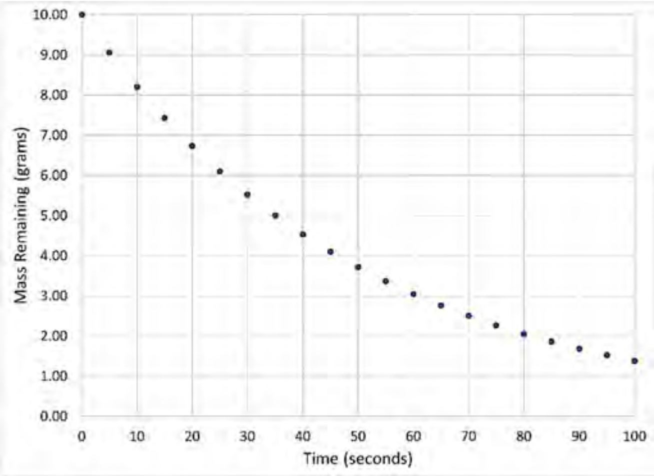
767. Does the voltage of the cell constructed in #764a increase, decrease, or remain the same as temperature is increased? Support your answer by referring to the Nernst equation.

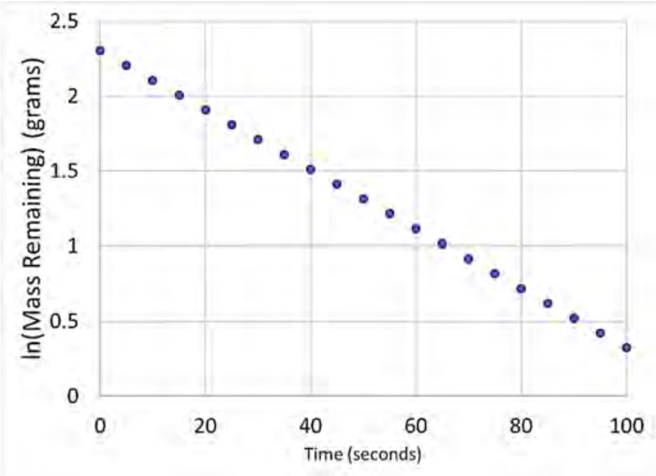
The voltage of the cell will increase. As T increases the value of (RT/nF) also increases which increases the value we are adding to the original cell voltage.

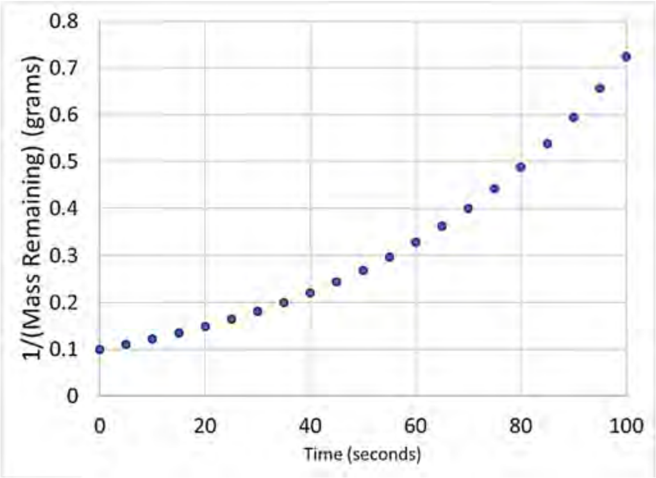
Review Topic 5.3

768. A certain gaseous reactant forms a single solid product, as shown below:

$$A_2(g) \rightarrow 2A(s)$$







a. What is the order of the reaction with respect to A_2? Justify your answer by referring to one of the three graphs.

The reaction is first order. The graph of ln[A_2] vs. time gives a straight line.

b. Use a different graph than the one you identified in 768.a to determine the order of the reactant. Justify your answer by referring to the graph.

The graph of [A_2] vs. time gives a consistent half-life. Half-life is a first order reaction. Also, the graph is a smooth curve which is also a sign of a first order reaction.

c. What is the half-life of $A_2(g)$? Explain your reasoning.

The half-life of A_2 is 35s. Every 35 s the concentration drops by 50%.

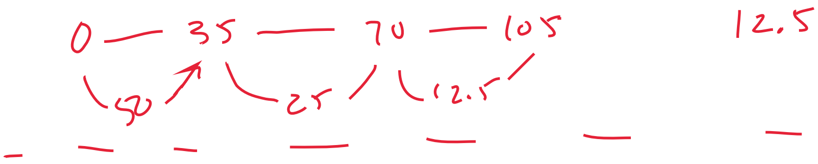
d. Determine the value of k.

$t_{1/2} = \frac{0.693}{k}$

$k = \frac{.693}{35} = 0.0198\ s^{-1} = k$

$k = \frac{.693}{t_{1/2}}$

e. What percentage of A_2 is remaining after 105 seconds?



12.5%

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

$\ln(x) - \ln(1) = -(.0198)(105)$

$e^{\wedge}\ln(x) - 0 = e^{\wedge} -2.079$

$x = .125 = [A_t]$

$\frac{.125\ M}{1\ M} \times 100 = \%$

12.5% remains

769. A student is attempting to plate out Ni(s) from a $Ni(NO_3)_2(aq)$ solution. The reaction is shown below.

$$Ni^{2+} + \underline{???}\ e^- \rightarrow Ni(s)$$

a. How many moles of electrons are used in the balanced equation?

2 moles

b. How many grams of Ni(s) will be plated out if the student uses a 10A circuit for 5 minutes?

$10A = 10\frac{C}{s}$

$5 min = 300 s$

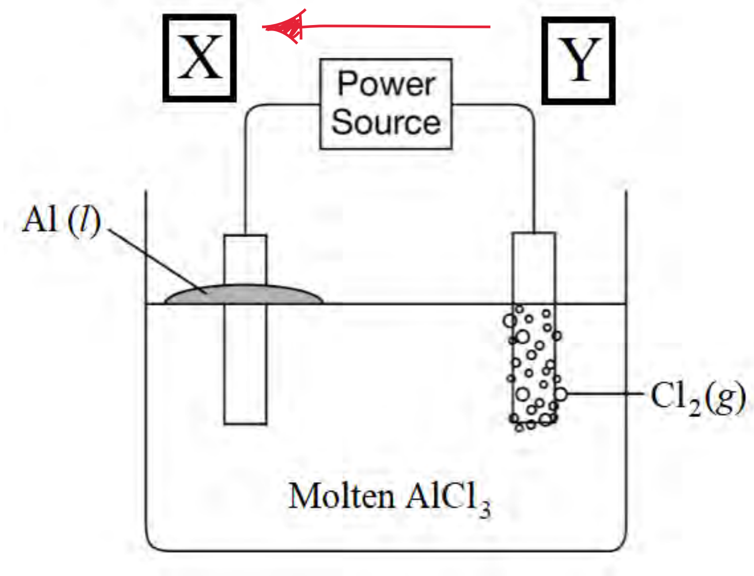
$$\frac{300 s \mid 10 C}{\mid 1 s} = \frac{3000 C \mid 1 mole\ e^-}{\mid 96485 C} = 0.0311 moles\ e^-$$

$$\frac{0.0311 mols\ e^- \mid 1 Ni(s)}{\mid 2 mols\ e^-} = \frac{0.0155 mols\ Ni \mid 58.69 g}{\mid 1 mol} = 0.912 g\ Ni$$

c. Determine the mass of Ni(s) that will be plated out if 0.4 faraday of electrical charge is used.

$$0.4\ faraday = \frac{0.4 moles\ e^- \mid 1 Ni(s)}{\mid 2 mols\ e^-} = \frac{0.2 mols\ Ni(s) \mid 58.69 g}{\mid 1 mol} = 11.7 g\ Ni$$

770. Molten $AlCl_3$ can be decomposed into its elements if a voltage is applied using inert electrodes. A representation of the cell is shown below.



Half-reaction	E° (V)
$Al^{3+} + 3e^- \rightarrow Al(s)$	− 1.66
~~$Cl_2 + 2e^- \rightarrow 2Cl^-$~~	~~+ 1.36~~
$2Cl^- \rightarrow Cl_2 + 2e^-$	−1.36

a. Draw an arrow between X and Y indicating the flow of electrons.

b. What is the minimum voltage required for this reaction to occur? Justify your answer with a calculation.

3.02 V is the minimum voltage required.

$$\begin{array}{r} -1.66V \\ + -1.36V \\ \hline -3.02V \end{array}$$

c. Determine the minimum current required to produce 3.00 g Al(l) at the cathode if the cell is run for 3000 s.

$$\frac{3.00g\ Al}{} \bigg| \frac{1 mol}{26.98g} = \frac{0.111 mol\ Al}{} \bigg| \frac{3e^-}{1 Al} = \frac{0.333 mol\ e^-}{} \bigg| \frac{96485 C}{1 mol\ e^-} = 32186 C$$

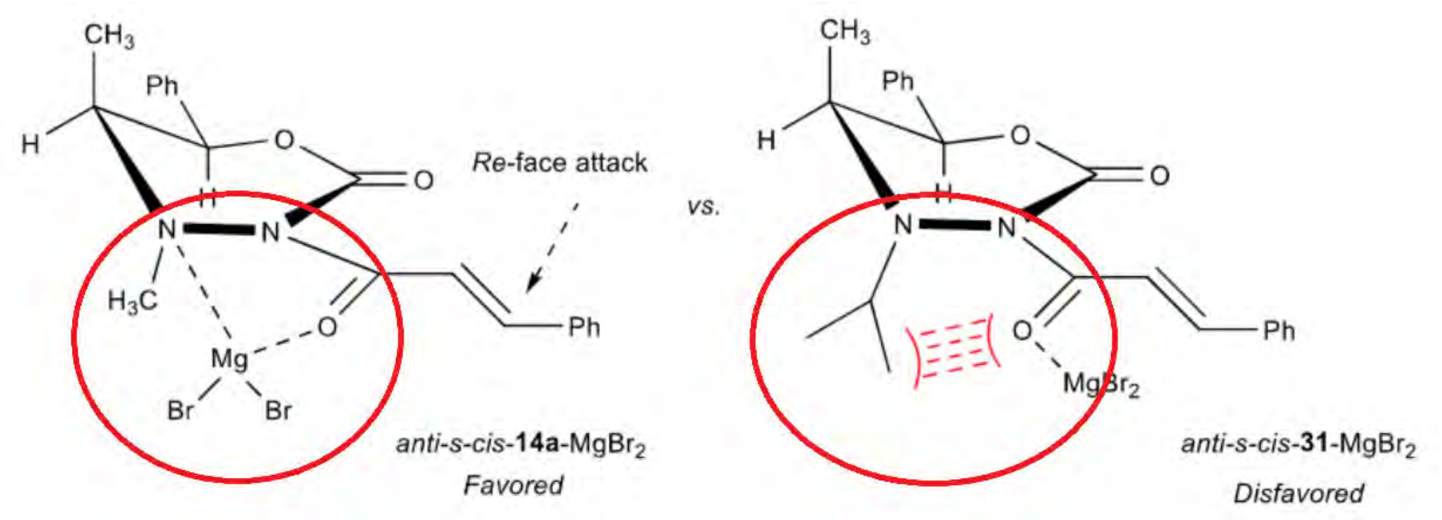
$$A = \frac{C}{s} = \frac{32186 C}{3000 s} = 10.73 A$$

d. A student claims that the same process occurs when producing Al(l) from $AlCl_3(aq)$. Do you agree or disagree with the student? Explain your reasoning.

I disagree with the student. When the same process is carried out with $AlCl_3(aq)$ there is a possibility of H_2O being oxidized or reduced instead of the reactions given here.

Review Topic 5.5, 5.4, 5.6, 5.10

771. Magnesium bromide, $MgBr_2$, is used as a reagent with oxadiazinones, as shown below. The shape of the oxadiazinones changes depending on how the $MgBr_2$ attaches. And how the $MgBr_2$ attaches is determined by the size of the side chain, shown in the circle below. On the left, the side chain coming off the N is a smaller methyl group, CH_3, while on the right it is a larger isopropyl group, $CH(CH_3)_2$. Notice the location of the $MgBr_2$ as a result of the different side chain. The next reaction occurs at the double bond labeled “Re-face attack”. The oxadiazinone on the left is favored for the re-face attack while the oxadiazinone on the right is not favored. Use the collision model to explain why the oxadiazinone on the right is not favored.



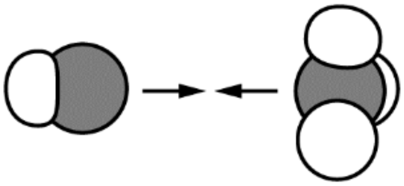
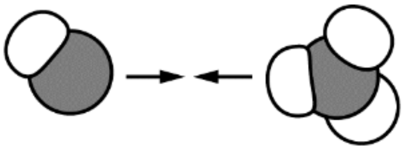
anti-s-cis-**14a**-$MgBr_2$
Favored

anti-s-cis-**31**-$MgBr_2$
Disfavored

The $MgBr_2$ doesn't allow a collision to occur at the Re-face attack in the Disfavored molecule because the $MgBr_2$ is blocking the path.

$$NO(g) + NO_3(g) \rightarrow 2\ NO_2(g)$$

772. The reaction between NO(g) and NO_3(g) is shown above. For the molecular orientations shown below, determine if a reaction will occur. If not, explain why not.

Molecular Orientation	Will a reaction occur?	If no, why?
	Yes or [✓ No]	For a reaction to occur the O in NO must collide with an O from NO_3.
	[✓ Yes] or No	

773. An elementary step is expected to be fast. However, during experimentation the step hardly proceeds at all and is very slow. Using particle-level reasoning, explain how each of the possible reasons below could make the elementary step slower than expected.

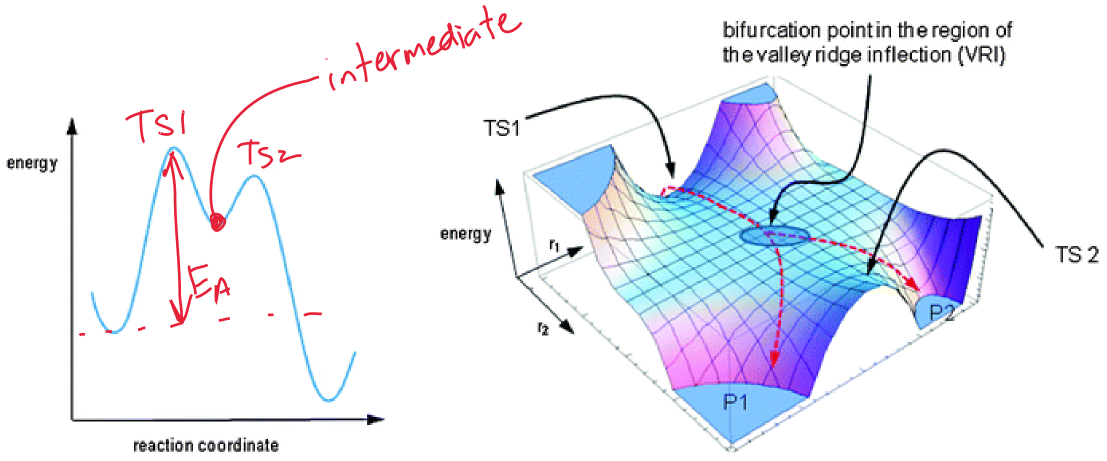
a. Molecular orientation

If the orientation isn't correct and is difficult to get the correct orientation the reaction will be slower because there will be fewer correct collisions.

b. Activation energy

A high activation energy creates an energy barrier that slows down a reaction w/ fewer particles w/ enough energy to overcome E_A.

774. Reaction energy profiles, the figure on the left, are constructed from potential energy surfaces, the figure on the right.



a. Is this forward reaction endothermic or exothermic? What about the reverse reaction?

Exothermic, Endothermic

b. How many elementary steps are in the reaction mechanism according to the reaction energy profile?

2

c. Use "TS1" and label the transition state 1 and "TS2" to label the transition state 2 on the reaction energy profile.

d. Label the intermediate on the reaction energy profile.

e. Label the activation energy, E_a, for step 1.