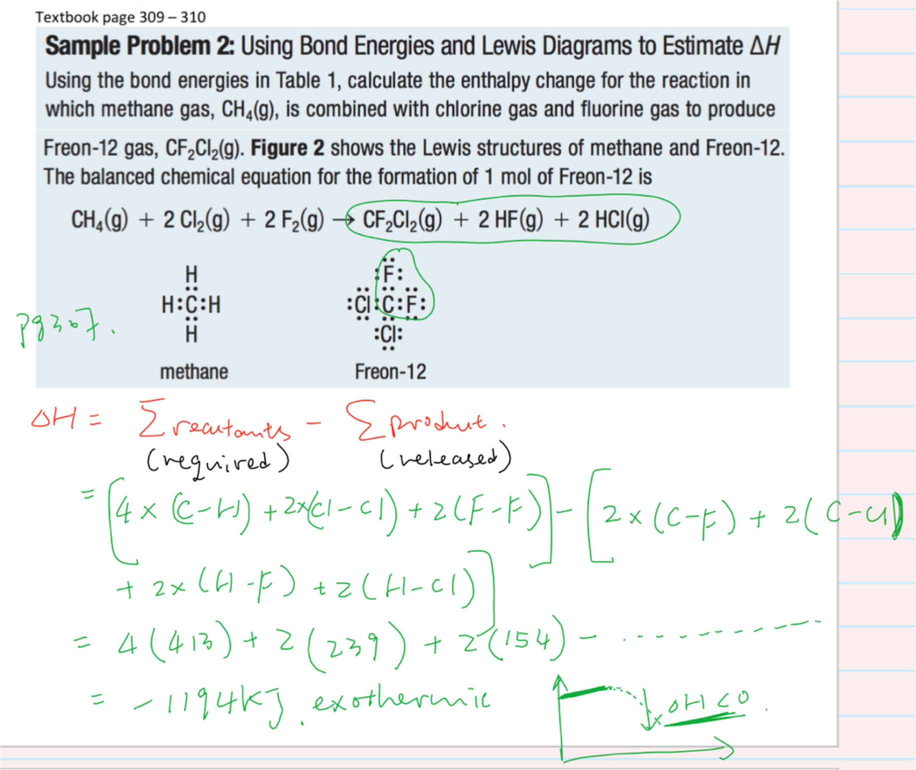
Sample questions from Notes…

Chapter 5:

1. $q = mc\Delta T \ and \ H_{system} = -H_{surrounding}$

2. Bond energy



Textbook page 309 – 310

Sample Problem 2: Using Bond Energies and Lewis Diagrams to Estimate ΔH

Using the bond energies in Table 1, calculate the enthalpy change for the reaction in which methane gas, $CH_4(g)$, is combined with chlorine gas and fluorine gas to produce Freon-12 gas, $CF_2Cl_2(g)$. **Figure 2** shows the Lewis structures of methane and Freon-12. The balanced chemical equation for the formation of 1 mol of Freon-12 is

$$CH_4(g) + 2\,Cl_2(g) + 2\,F_2(g) \to CF_2Cl_2(g) + 2\,HF(g) + 2\,HCl(g)$$

pg 307.

methane

Freon-12

$\Delta H = \sum reactants - \sum product.$

(required) (released)

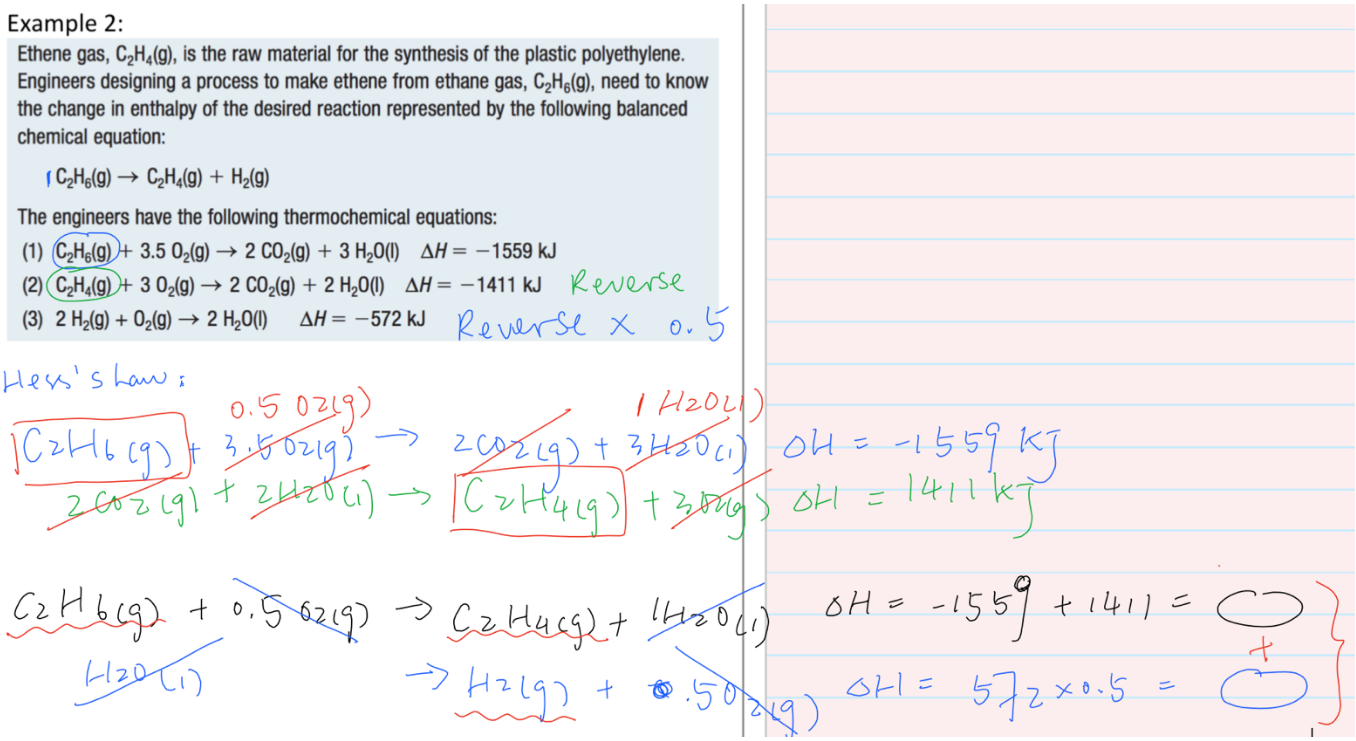
$= \left[4 \times (C-H) + 2\times(Cl-Cl) + 2(F-F)\right] - \left[2 \times (C-F) + 2(C-Cl) + 2\times(H-F) + 2(H-Cl)\right]$

$= 4(413) + 2(239) + 2(154) - \ldots\ldots$

$= -1194\,kJ$. exothermic

$\Delta H < 0$.

3. Hess's Law



Example 2:

Ethene gas, $C_2H_4(g)$, is the raw material for the synthesis of the plastic polyethylene. Engineers designing a process to make ethene from ethane gas, $C_2H_6(g)$, need to know the change in enthalpy of the desired reaction represented by the following balanced chemical equation:

$$C_2H_6(g) \to C_2H_4(g) + H_2(g)$$

The engineers have the following thermochemical equations:

(1) $C_2H_6(g) + 3.5\,O_2(g) \to 2\,CO_2(g) + 3\,H_2O(l) \quad \Delta H = -1559\ kJ$

(2) $C_2H_4(g) + 3\,O_2(g) \to 2\,CO_2(g) + 2\,H_2O(l) \quad \Delta H = -1411\ kJ$ Reverse

(3) $2\,H_2(g) + O_2(g) \to 2\,H_2O(l) \quad \Delta H = -572\ kJ$ Reverse × 0.5

Hess's Law:

$\boxed{C_2H_6(g)} + \cancel{3.5\,O_2(g)}\ 0.5\,O_2(g) \to \cancel{2CO_2(g)} + \cancel{3H_2O(l)}\ 1\,H_2O(l) \quad \Delta H = -1559\,kJ$

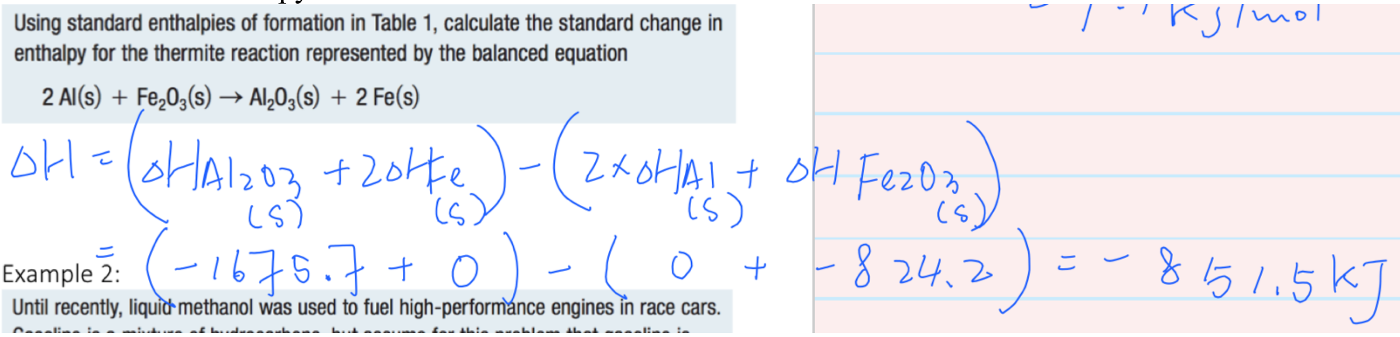
$\cancel{2CO_2(g)} + \cancel{2H_2O(l)} \to \boxed{C_2H_4(g)} + \cancel{3O_2(g)} \quad \Delta H = 1411\,kJ$

$C_2H_6(g) + \cancel{0.5\,O_2(g)} \to C_2H_4(g) + \cancel{1H_2O(l)} \quad \Delta H = -1559 + 1411 =$

$\cancel{H_2O(l)} \to H_2(g) + \cancel{0.5\,O_2(g)} \quad \Delta H = 572 \times 0.5 =$

+

4. Standard enthalpy of formation



Using standard enthalpies of formation in Table 1, calculate the standard change in enthalpy for the thermite reaction represented by the balanced equation

$2\,Al(s) + Fe_2O_3(s) \rightarrow Al_2O_3(s) + 2\,Fe(s)$

kJ/mol

$\Delta H = (\Delta H\,Al_2O_{3(s)} + 2\Delta H\,Fe_{(s)}) - (2 \times \Delta H\,Al_{(s)} + \Delta H\,Fe_2O_{3(s)})$

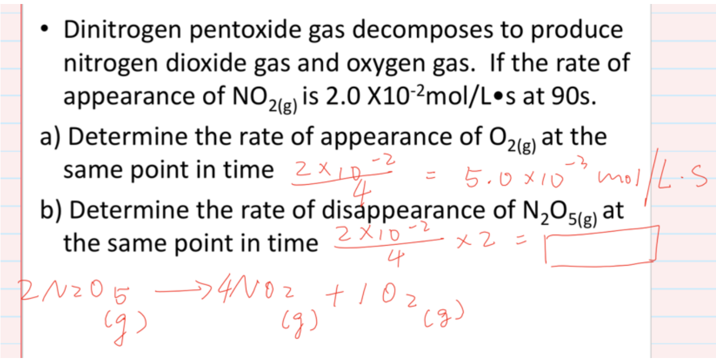
$= (-1675.7 + 0) - (0 + -824.2) = -851.5\,kJ$

Example 2:

Until recently, liquid methanol was used to fuel high-performance engines in race cars.

Chapter 6:

5. Reaction rate



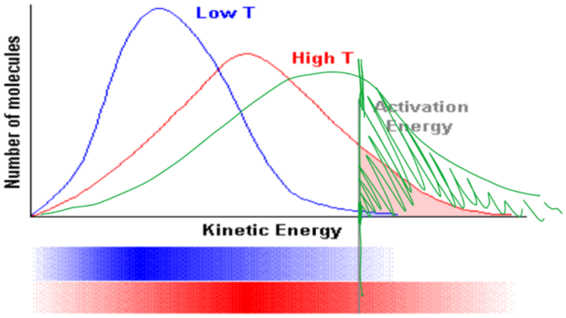
- Dinitrogen pentoxide gas decomposes to produce nitrogen dioxide gas and oxygen gas. If the rate of appearance of $NO_{2(g)}$ is 2.0 X10^{-2}mol/L•s at 90s.

a) Determine the rate of appearance of $O_{2(g)}$ at the same point in time $\frac{2 \times 10^{-2}}{4} = 5.0 \times 10^{-3}$ mol/L·s

b) Determine the rate of disappearance of $N_2O_{5(g)}$ at the same point in time $\frac{2 \times 10^{-2}}{4} \times 2 = \boxed{}$

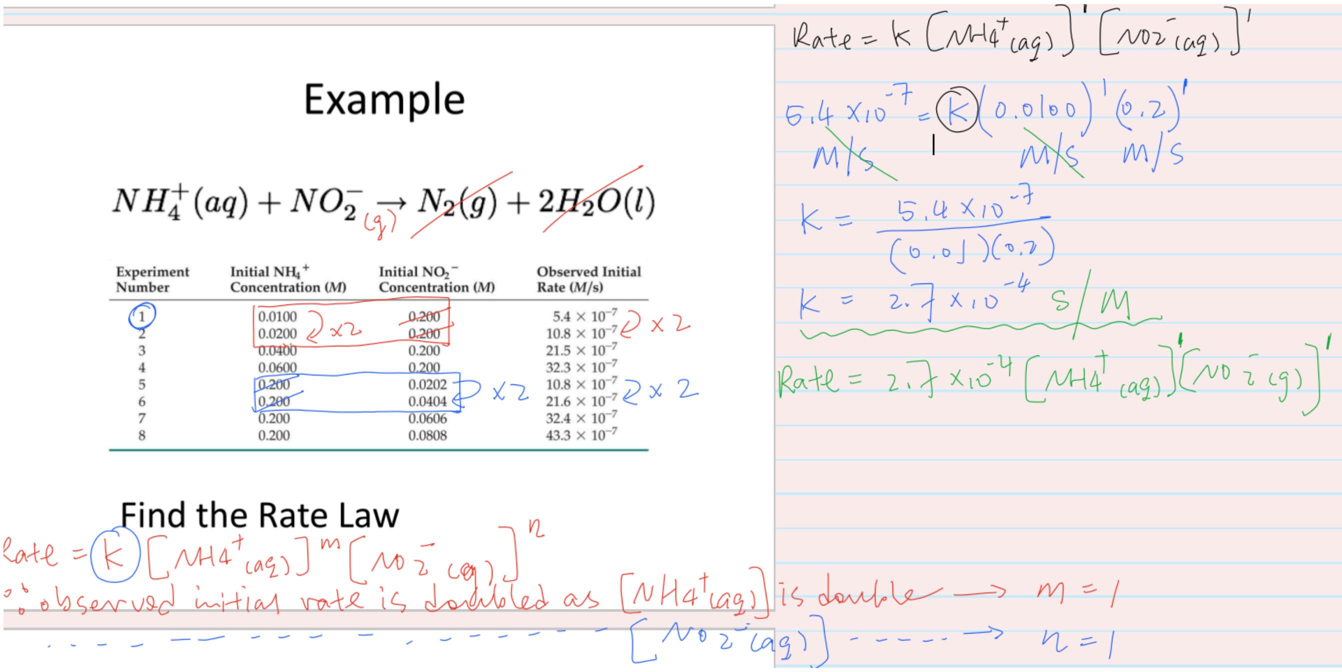
$2N_2O_{5(g)} \longrightarrow 4NO_{2(g)} + 1O_{2(g)}$

6. Factor affecting rate of reaction and collision theory (6.2 – 6.3)



as well as catalyst…

7. Rate Law (6.5)



Example

$NH_4^+(aq) + NO_{2\,(g)}^- \rightarrow \cancel{N_2(g)} + \cancel{2H_2O(l)}$

Experiment Number	Initial NH_4^+ Concentration (M)	Initial NO_2^- Concentration (M)	Observed Initial Rate (M/s)
1	0.0100	0.200	5.4×10^{-7}
2	0.0200	0.200	10.8×10^{-7}
3	0.0400	0.200	21.5×10^{-7}
4	0.0600	0.200	32.3×10^{-7}
5	0.200	0.0202	10.8×10^{-7}
6	0.200	0.0404	21.6×10^{-7}
7	0.200	0.0606	32.4×10^{-7}
8	0.200	0.0808	43.3×10^{-7}

×2 ×2

×2 ×2

Find the Rate Law

$Rate = k\,[NH_4^+{}_{(aq)}]^m\,[NO_2^-{}_{(aq)}]^n$

∴ observed initial rate is doubled as $[NH_4^+{}_{(aq)}]$ is double → m = 1

- $[NO_2^-{}_{(aq)}]$ - - - - → n = 1

$Rate = k\,[NH_4^+{}_{(aq)}]^1\,[NO_2^-{}_{(aq)}]^1$

$5.4 \times 10^{-7} = k\,(0.0100)^1\,(0.2)^1$

$\cancel{M/s}$ $\cancel{M/s}$ M/s

$k = \frac{5.4 \times 10^{-7}}{(0.01)(0.2)}$

$k = 2.7 \times 10^{-4}$ s/M

$Rate = 2.7 \times 10^{-4}\,[NH_4^+{}_{(aq)}]^1\,[NO_2^-{}_{(g)}]^1$

8. Reaction mechanism and rate determining step

$2NO_{(g)} + 2H_{2(g)}$ --- > $N_{2(g)} + 2H_2O_{(g)}$ $\boxed{rate = k[H_2]}$

$rate = k[NO]^0[H_2]^1$

The mechanism for the above reaction occurs in three steps:

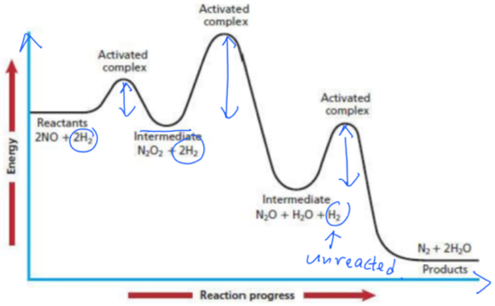
$2NO \rightarrow N_2O_2$ (fast) intermediate product

$\cancel{N_2O_2} + H_2 \rightarrow N_2O + H_2O$ $\boxed{\text{(slow)}}$

$N_2O + H_2 \rightarrow N_2 + H_2O$ (fast)

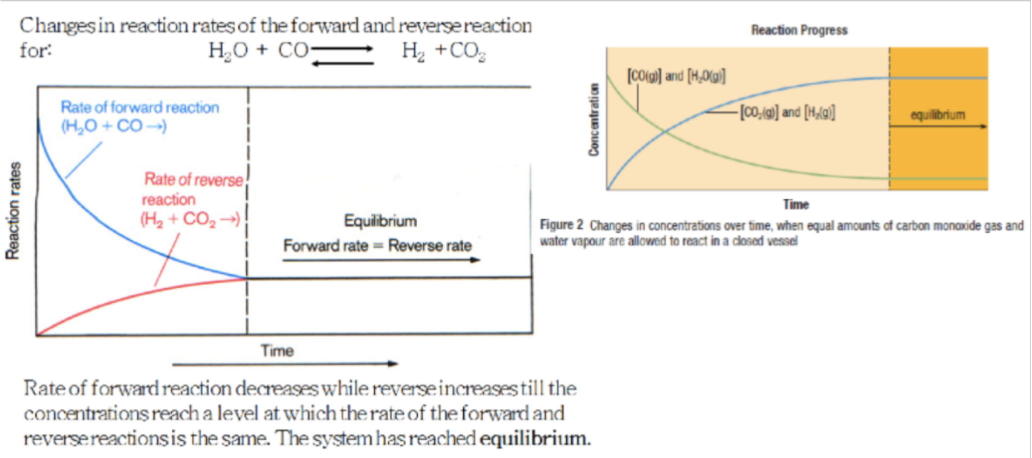
* only look at "slowest"

* NO intermediate molecule in Rate Law.



Chapter 7:

9. Reaction rate and change of concentration in dynamic equilibrium



Changes in reaction rates of the forward and reverse reaction for: $H_2O + CO \rightleftharpoons H_2 + CO_2$

Figure 2 Changes in concentrations over time, when equal amounts of carbon monoxide gas and water vapour are allowed to react in a closed vessel

Rate of forward reaction decreases while reverse increases till the concentrations reach a level at which the rate of the forward and reverse reactions is the same. The system has reached **equilibrium.**

10. Le Chatelier's Principle and drawing diagram with disturbances (7.5)

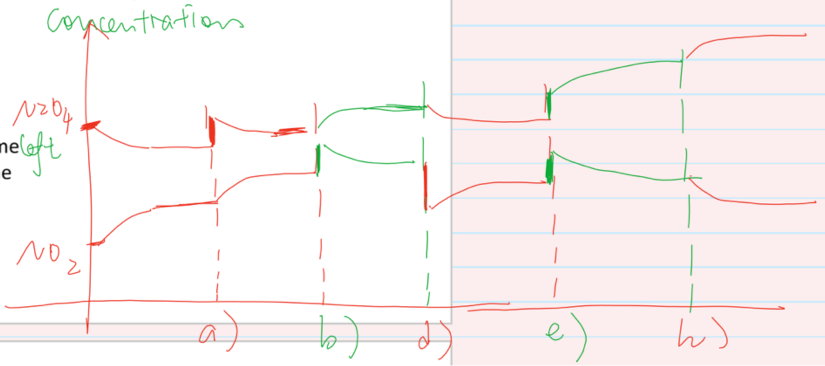
- Consider the following equilibrium system:

$$\overset{1}{N_2O_{4\,(g)}} + energy \Leftrightarrow \overset{2}{2NO_{2(g)}}$$

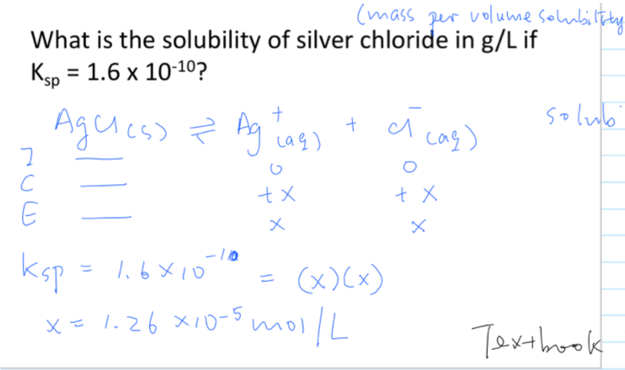
- How will the equilibrium shift if the following disturbances occur:

a) Addition of $N_2O_{4(g)}$ right
b) Addition of $NO_{2(g)}$ left
~~c)~~ Removal of $N_2O_{4(g)}$
d) Removal of $NO_{2(g)}$ right
↑P e) Decrease in container volume left
~~f)~~ Increase in container volume
~~g)~~ Increase in temperature
h) Decrease in temperature

left



11. K_{sp} (solubility product), molar solubility or mass per volume solubility (7.6)



(mass per volume solubility)

What is the solubility of silver chloride in g/L if $K_{sp} = 1.6 \times 10^{-10}$?

$AgCl_{(s)} \rightleftharpoons Ag^+_{(aq)} + Cl^-_{(aq)}$

I — 0 0

C — +x +x

E — x x

$K_{sp} = 1.6 \times 10^{-10} = (x)(x)$

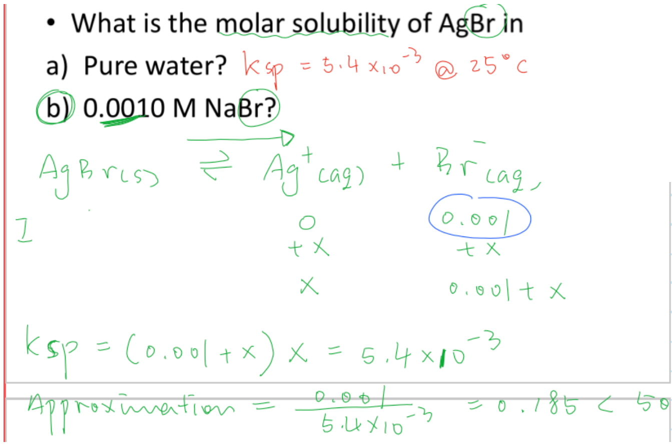
$x = 1.26 \times 10^{-5}$ mol/L

$\text{solubility} = (1.26 \times 10^{-5} \text{ mol} \times 143.32 \text{ g/mol})/L$

$= 1.81 \times 10^{-3}$ g/L

Textbook Pg 464 / sample problem 3.

12. Trial ion product, predict precipitation and common ion effect



- What is the molar solubility of AgBr in

a) Pure water? $K_{sp} = 5.4 \times 10^{-3}$ @ 25°C

b) 0.0010 M NaBr?

$AgBr_{(s)} \rightleftharpoons Ag^+_{(aq)} + Br^-_{(aq)}$

I 0 0.001

+x +x

x 0.001 + x

$K_{sp} = (0.001 + x)x = 5.4 \times 10^{-3}$

$\text{Approximation} = \frac{0.001}{5.4 \times 10^{-3}} = 0.185 < 500$

$x^2 + 0.001x - 5.4 \times 10^{-3} = 0$

$x = \frac{-0.001 \pm 0.147}{2}$

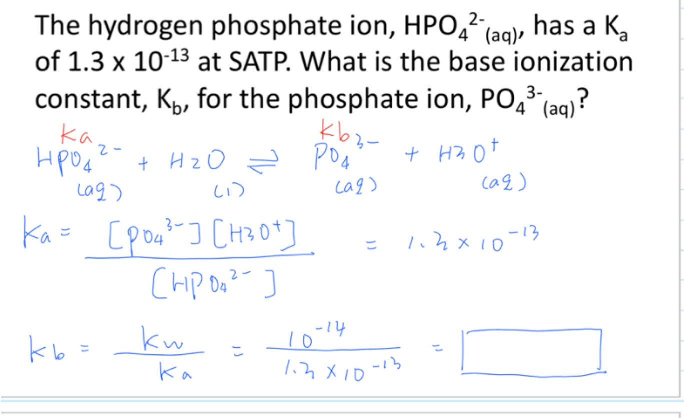
$x = 0.073$ mol/L

molar solubility of AgBr

Chapter 8:

13. Arrhenius theory of acid-base & Bronsted and Lowry; strength of conjugate acid and conjugate base (8.1)

14. Ion-product constant, PH and POH (8.2)



The hydrogen phosphate ion, $HPO_4^{2-}{}_{(aq)}$, has a K_a of 1.3×10^{-13} at SATP. What is the base ionization constant, K_b, for the phosphate ion, $PO_4^{3-}{}_{(aq)}$?

$\overset{Ka}{HPO_4^{2-}}{}_{(aq)} + H_2O_{(l)} \rightleftharpoons \overset{Kb}{PO_4^{3-}}{}_{(aq)} + H_3O^+_{(aq)}$

$Ka = \frac{[PO_4^{3-}][H_3O^+]}{[HPO_4^{2-}]} = 1.3 \times 10^{-13}$

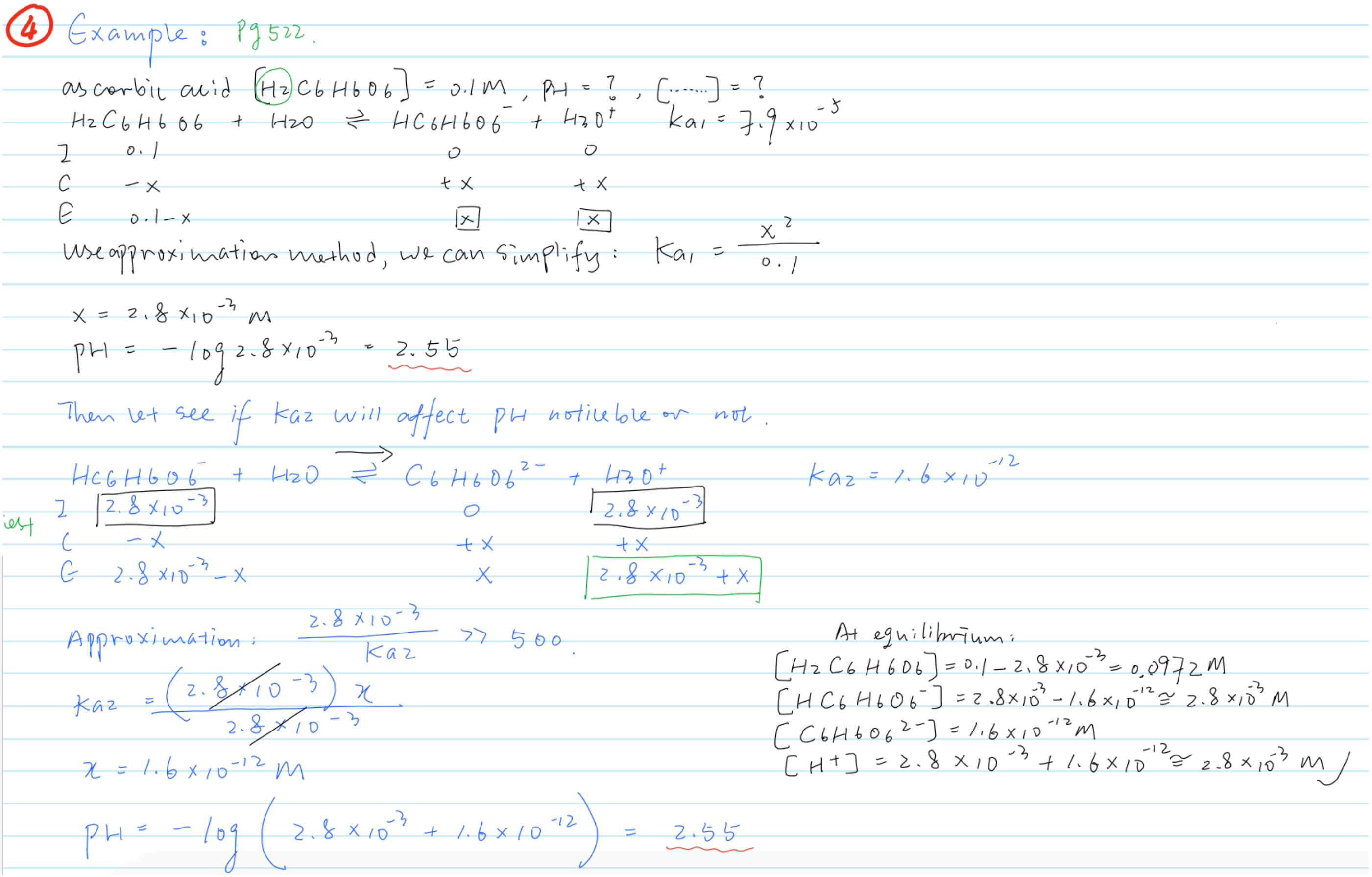
$Kb = \frac{Kw}{Ka} = \frac{10^{-14}}{1.3 \times 10^{-13}} = \boxed{}$

$\overset{Kb}{PO_4^{3-}} + H_2O \rightleftharpoons \overset{Ka}{HPO_4^{2-}} + OH^-$

$Kb = \frac{[HPO_4^{2-}][OH^-]}{[PO_4^{3-}]}$.

You could check the product of Ka of HPO_4^{2-} and Kb of PO_4^{3-} is equal to $Kw = 10^{-14}$.

15. Use % ionization to calculate Ka or Kb & use Ka or Kb to calculate pH (vice versa) and calculate related to polyprotic acids (8.4)



④ Example : pg 522.

ascorbic acid $[H_2C_6H_6O_6] = 0.1\,M$, pH = ? , [......] = ?

	$H_2C_6H_6O_6$	+ H_2O ⇌	$HC_6H_6O_6^-$	+ H_3O^+
I	0.1		0	0
C	−x		+x	+x
E	0.1−x		$\boxed{x}$	$\boxed{x}$

$Ka_1 = 7.9 \times 10^{-5}$

Use approximation method, we can simplify: $Ka_1 = \frac{x^2}{0.1}$

$x = 2.8 \times 10^{-3}\,M$

$pH = -\log 2.8 \times 10^{-3} = 2.55$

Then let see if Ka2 will affect pH noticeble or not.

	$HC_6H_6O_6^-$	+ H_2O $\xrightarrow{}$ ⇌	$C_6H_6O_6^{2-}$	+ H_3O^+
I	$\boxed{2.8 \times 10^{-3}}$		0	$\boxed{2.8 \times 10^{-3}}$
C	−x		+x	+x
E	$2.8 \times 10^{-3} - x$		x	$\boxed{2.8 \times 10^{-3} + x}$

$Ka_2 = 1.6 \times 10^{-12}$

iest

Approximation: $\frac{2.8 \times 10^{-3}}{Ka_2} >> 500.$

$Ka_2 = \frac{(\cancel{2.8 \times 10^{-3}})\,x}{\cancel{2.8 \times 10^{-3}}}$

$x = 1.6 \times 10^{-12}\,M$

$pH = -\log\left(2.8 \times 10^{-3} + 1.6 \times 10^{-12}\right) = 2.55$

At equilibrium:

$[H_2C_6H_6O_6] = 0.1 - 2.8 \times 10^{-3} = 0.0972\,M$

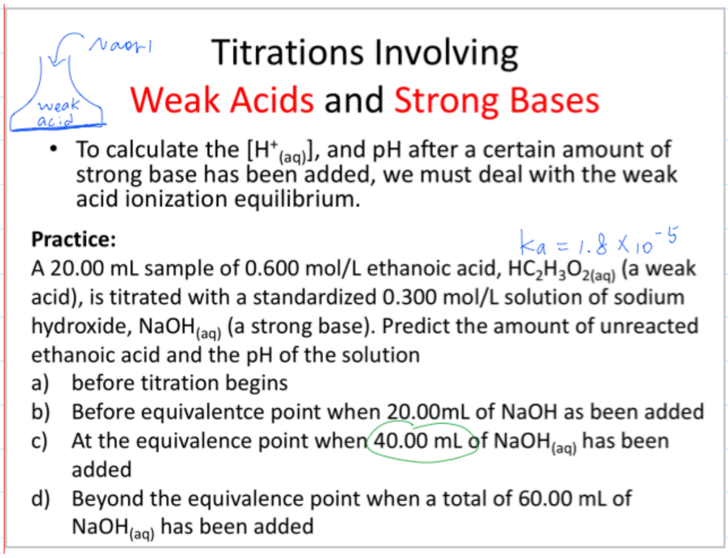
$[HC_6H_6O_6^-] = 2.8 \times 10^{-3} - 1.6 \times 10^{-12} \cong 2.8 \times 10^{-3}\,M$

$[C_6H_6O_6^{2-}] = 1.6 \times 10^{-12}\,M$

$[H^+] = 2.8 \times 10^{-3} + 1.6 \times 10^{-12} \cong 2.8 \times 10^{-3}\,M$ ✓

16. Analysis acid-base properties of salt (8.6) & hydrolysis of amphiprotic ions

17. Titration (I only put an example we did in class about weak acid/base titration, read textbook for strong acid/base titration)



Titrations Involving Weak Acids and Strong Bases

- To calculate the $[H^+_{(aq)}]$, and pH after a certain amount of strong base has been added, we must deal with the weak acid ionization equilibrium.

Practice:

$Ka = 1.8 \times 10^{-5}$

A 20.00 mL sample of 0.600 mol/L ethanoic acid, $HC_2H_3O_{2(aq)}$ (a weak acid), is titrated with a standardized 0.300 mol/L solution of sodium hydroxide, $NaOH_{(aq)}$ (a strong base). Predict the amount of unreacted ethanoic acid and the pH of the solution

a) before titration begins

b) Before equivalentce point when 20.00mL of NaOH as been added

c) At the equivalence point when 40.00 mL of $NaOH_{(aq)}$ has been added

d) Beyond the equivalence point when a total of 60.00 mL of $NaOH_{(aq)}$ has been added

a) $HC_2H_3O_2 \rightleftharpoons H^+ + C_2H_3O_2^-$

0.6	0	0
0.6 − x	+ x	+ x ....

b) $HC_2H_3O_2 + Na\cancel{O}H \rightleftharpoons NaC_2H_3O_2 + H_2O$

$n(HC_2H_3O_2) = 0.02L \times 0.6 mol/L = 0.012 mol$

$n(NaOH) = 0.02L \times 0.3 mol/L = 0.006 mol$

Before reaction	0.012	0.006	0
after reaction	0.006	0	0.006

$[HC_2H_3O_2] = \frac{0.006 mol}{0.02L + 0.02L} = 0.15 M$

$[C_2H_3O_2^-] = 0.15 M$

$HC_2H_3O_2 + H_2O \rightleftharpoons C_2H_3O_2^- + H_3O^+$

I	0.15	0.15	0
C	− x	− x	+ x
E	0.15 − x	0.15 + x	x

$Ka = 1.8 \times 10^{-5} = \frac{(0.15 + x)x}{0.15 - x} \Rightarrow x \Rightarrow pH$

$n = C \cdot V$

		$HC_2H_3O_2$	NaOH	NaC_2H_3O
c)	before reaction	0.012	0.012	0
	after reaction	0	0	0.012

$NaC_2H_3O \rightarrow Na^+$

$\searrow C_2H_3O^-$

Strong conjugate base of a weak acid.

$NaC_2H_3O \rightleftharpoons Na^+ + C_2H_3O^-$

$C_2H_3O^- + H_2O \rightleftharpoons HC_2H_3O + OH^-$

$\frac{0.012 mol}{0.02L + 0.04L}$

$= 0.2M$

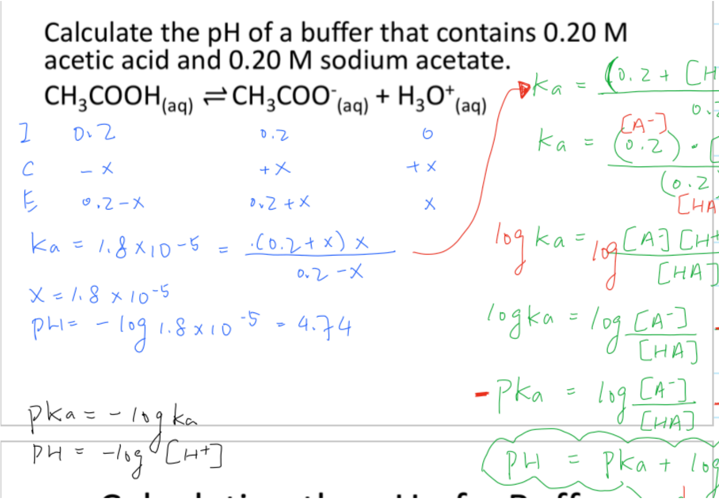
0.2	0	0
0.2 − x	x	x

$Kb = \frac{10^{-14}}{1.8 \times 10^{-5}} = \frac{x^2}{0.2 - x}$

$5.56 \times 10^{-10} = ...$

$pH = 9.1$ basic.

18. Buffer system



Calculate the pH of a buffer that contains 0.20 M acetic acid and 0.20 M sodium acetate.

$CH_3COOH_{(aq)} \rightleftharpoons CH_3COO^-_{(aq)} + H_3O^+_{(aq)}$

I	0.2	0.2	0
C	− x	+ x	+ x
E	0.2 − x	0.2 + x	x

$Ka = 1.8 \times 10^{-5} = \frac{(0.2 + x)x}{0.2 - x}$

$x = 1.8 \times 10^{-5}$

$pH = -\log 1.8 \times 10^{-5} = 4.74$

$pKa = -\log ka$

$pH = -\log [H^+]$

$Ka = \frac{(0.2 + [H^+]) \cdot ([H^+])}{0.2 - [H^+]}$

$Ka = \frac{\overset{[A^-]}{(0.2)} \cdot [H^+]}{\underset{[HA]}{(0.2)}}$

$\log ka = \log \frac{[A^-][H^+]}{[HA]} = \log\left(\frac{[A^-]}{[HA]} \cdot \frac{[H^+]}{1}\right)$

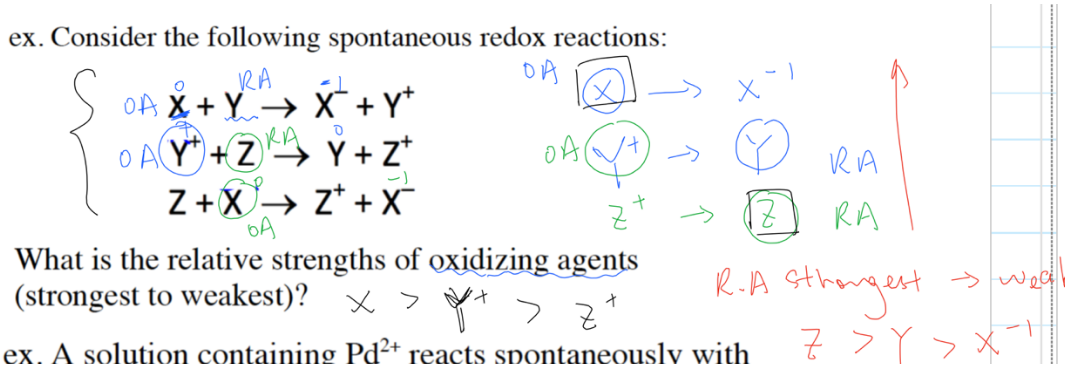
$\log ka = \log \frac{[A^-]}{[HA]} + \log [H^+]$ product rule

$-pKa = \log \frac{[A^-]}{[HA]} - pH$

$\boxed{pH = pKa + \log \frac{[A^-]}{[HA]}}$

Chapter 9 and chapter 10:

19. Identifying Oxidation reaction, reduction reaction, oxidizing agent, reducing agent, oxidation number, and strength of OA and RA. (9.1 – 9.2)



ex. Consider the following spontaneous redox reactions:

OA $\overset{0}{X} + \overset{RA}{Y} \rightarrow \overset{-1}{X^-} + Y^+$

OA $\boxed{Y^+} + \boxed{Z} \overset{RA}{\rightarrow} \overset{0}{Y} + Z^+$

$Z + \underset{OA}{X} \rightarrow Z^+ + \overset{-1}{X^-}$

OA $\boxed{X} \rightarrow X^{-1}$

OA $Y^+ \rightarrow Y$ RA

$Z^+ \rightarrow \boxed{Z}$ RA

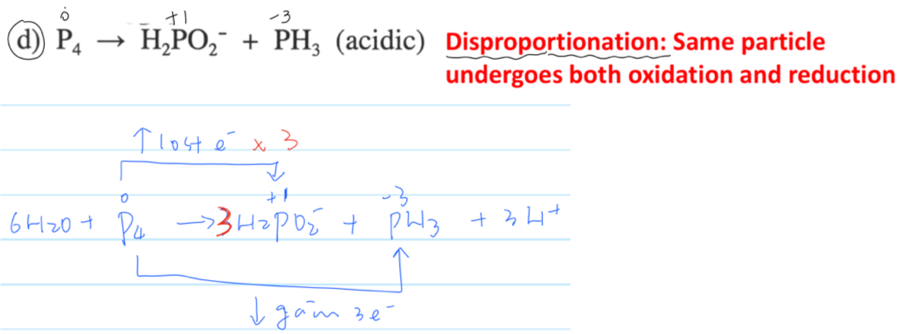
What is the relative strengths of oxidizing agents (strongest to weakest)? $X > Y^+ > Z^+$

R.A strongest → weakest

$Z > Y > X^{-1}$

ex. A solution containing Pd^{2+} reacts spontaneously with

20. Two different method of balancing equations in acidic or basic solution (9.3)



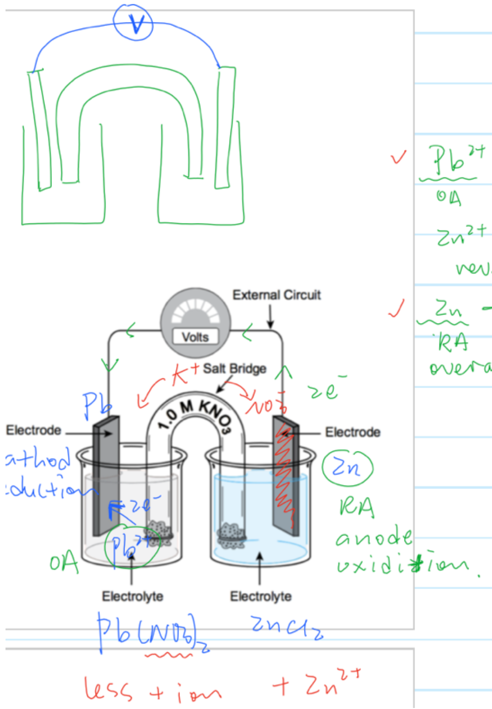
(d) $\overset{0}{P_4} \rightarrow \overset{+1}{H_2PO_2^-} + \overset{-3}{PH_3}$ (acidic) **Disproportionation: Same particle undergoes both oxidation and reduction**

↑ lost e^- x 3

$6H_2O + \overset{0}{P_4} \rightarrow 3\overset{+1}{H_2PO_2^-} + \overset{-3}{PH_3} + 3H^+$

↓ gain $3e^-$

21. Galvanic cells (10.1 – 10.2): identify name and function of different parts of galvanic cells, identify cathode, anode, direction of electron low, half-cell reaction, how salt bridge balance excess of charge accumulation, etc…Calculate cell potential and determine if it is spontaneous or non-spontaneous…



✓ $\underset{OA}{Pb^{2+}} + 2e^- \rightarrow Pb$ reduction – cathod.

$Zn^{2+} + 2e^- \rightarrow Zn$

reverse it

✓ $\underset{RA}{Zn} \rightarrow Zn^{2+} + 2e^-$ oxidation – anode

overall reaction: $Pb^{2+} + Zn \rightleftharpoons Pb + Zn^{2+}$

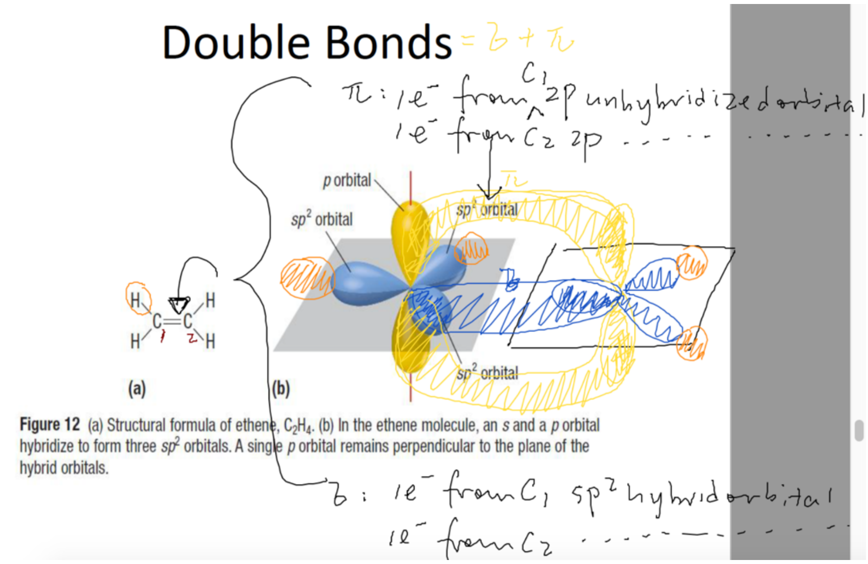
22. Preventing corrosion
23. Selecting preferred reaction in galvanic cell
24. Molten binary salt electrolysis
25. Aqueous solution electrolysis
26. Overpotential effect

Chapter 3:

27. Quantum numbers, Pauli Exclusion Principle, Aufbau principle; electron configuration, energy-level diagram, Cr and Cu exception, determine possible charges (3.4 – 3.5)

28. Lewis structure, VSEPR 3D structure of central atoms, bond and molecular polarity, hybridization

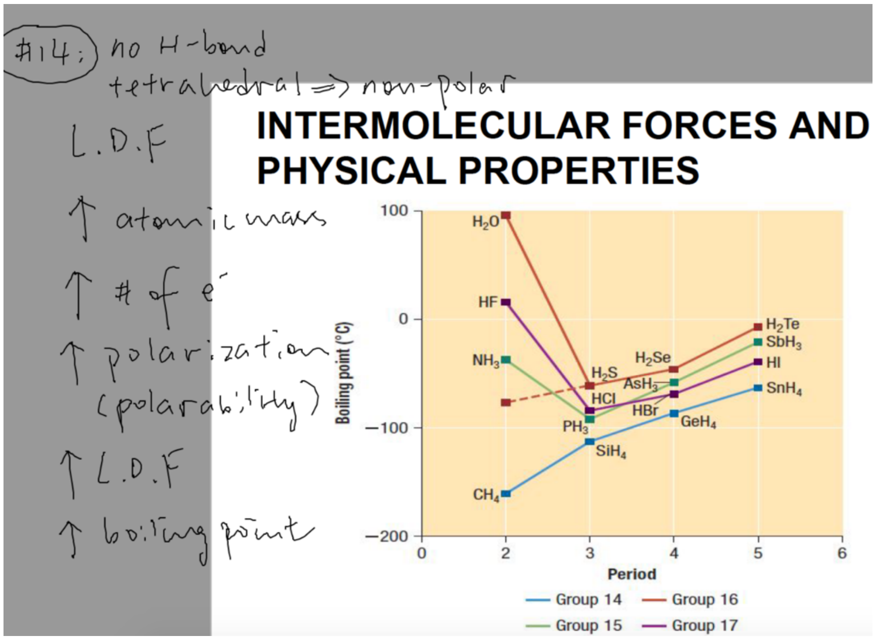
29. intermolecular forces: dipole-dipole, hydrogen bond, London dispersion forces, polarizability; physical properties (ranking melting and boiling point)…


Double Bonds = σ + π

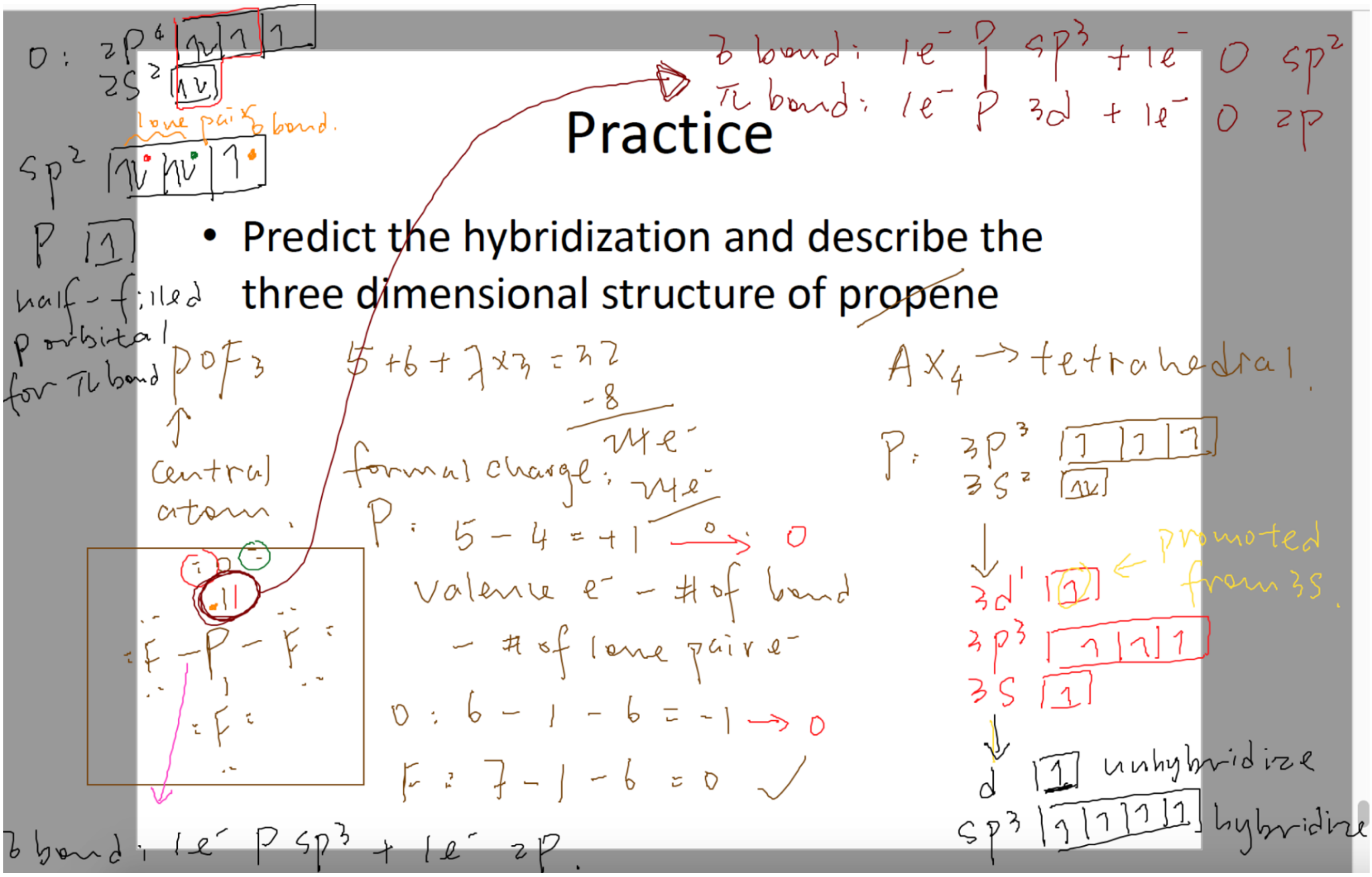
π : 1e⁻ from C1 2p unhybridized orbital
1e⁻ from C2 2p ....

σ : 1e⁻ from C1 sp² hybrid orbital
1e⁻ from C2 ....

Figure 12 (a) Structural formula of ethene, C_2H_4. (b) In the ethene molecule, an *s* and a *p* orbital hybridize to form three *sp*² orbitals. A single *p* orbital remains perpendicular to the plane of the hybrid orbitals.


#14: no H-bond
tetrahedral => non-polar
L.D.F
↑ atomic mass
↑ # of e⁻
↑ polarization (polarability)
↑ L.D.F
↑ boiling point

INTERMOLECULAR FORCES AND PHYSICAL PROPERTIES



Practice

- Predict the hybridization and describe the three dimensional structure of ~~propene~~

O: $2p^4$ $[\uparrow\downarrow|\uparrow|\uparrow]$

$2s^2$ $[\uparrow\downarrow]$

lone pairs, σ bond.

sp^2 $[\uparrow\downarrow|\uparrow\downarrow|\uparrow]$

p $[\uparrow]$

half-filled p orbital for π bond

POF_3

↑ central atom.

$5+6+7\times3=32$

-8

$24e^-$

formal charge: $24e^-$

P: $5-4=+1 \rightarrow 0$

valence e^- − # of bond − # of lone pair e^-

O: $6-1-6=-1 \rightarrow 0$

F: $7-1-6=0$ ✓

AX_4 → tetrahedral.

P: $3p^3$ $[\uparrow|\uparrow|\uparrow]$

$3s^2$ $[\uparrow\downarrow]$

↓

$3d^1$ $[\uparrow]$ ← promoted from 3s.

$3p^3$ $[\uparrow|\uparrow|\uparrow]$

$3s$ $[\uparrow]$

↓

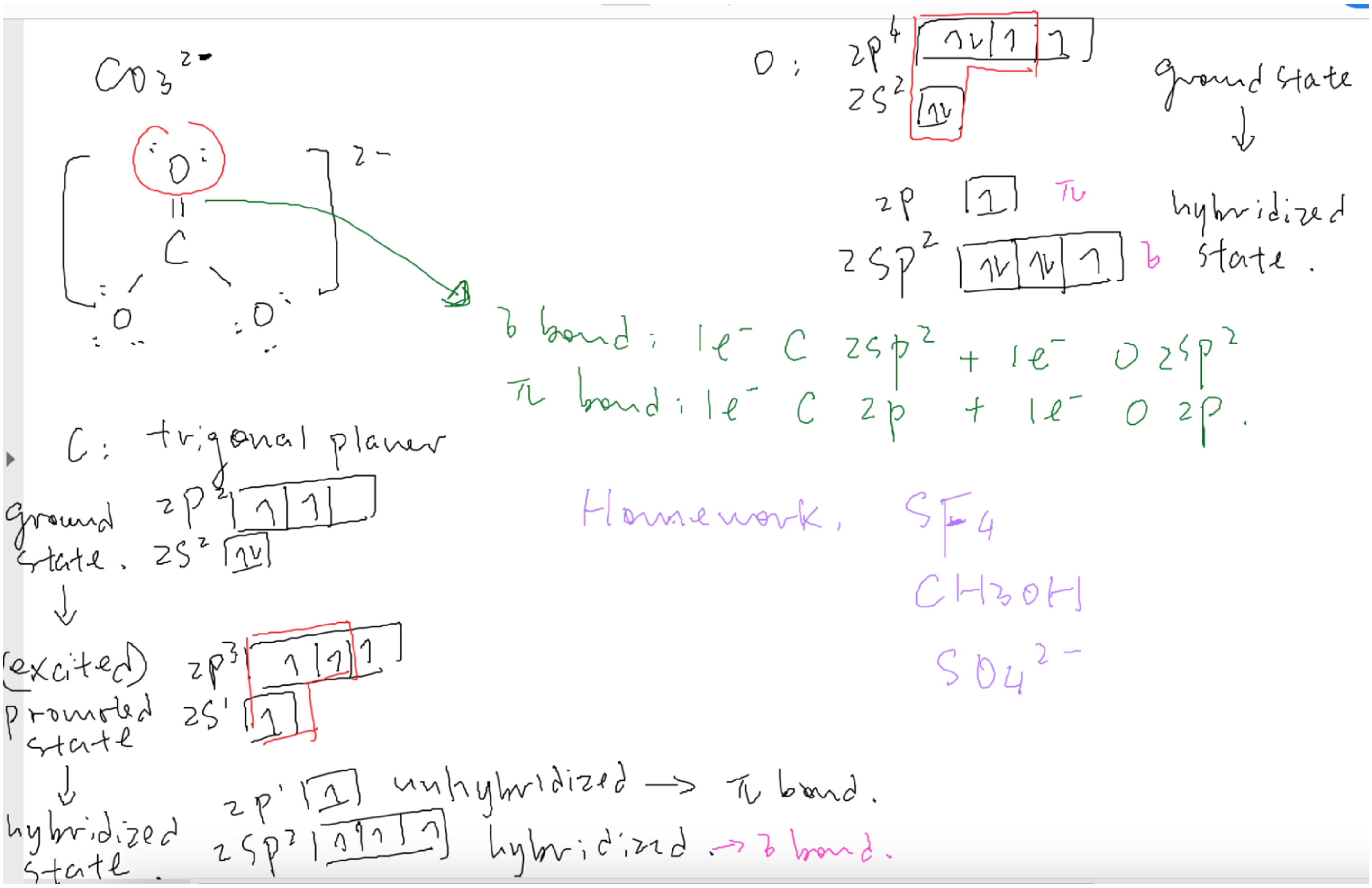
d $[\uparrow]$ unhybridize

sp^3 $[\uparrow|\uparrow|\uparrow|\uparrow]$ hybridize

σ bond: $1e^-$ P sp^3 + $1e^-$ O sp^2

π bond: $1e^-$ P 3d + $1e^-$ O 2p

σ bond: $1e^-$ P sp^3 + $1e^-$ 2p.



CO_3^{2-}

O: $2p^4$ $[\uparrow\downarrow|\uparrow|\uparrow]$

$2s^2$ $[\uparrow\downarrow]$

ground state

↓

hybridized state.

2p $[\uparrow]$ π

$2sp^2$ $[\uparrow\downarrow|\uparrow\downarrow|\uparrow]$ σ

σ bond: $1e^-$ C $2sp^2$ + $1e^-$ O $2sp^2$

π bond: $1e^-$ C 2p + $1e^-$ O 2p.

C: trigonal planer

ground state. $2p^2$ $[\uparrow|\uparrow|\ ]$

$2s^2$ $[\uparrow\downarrow]$

↓

(excited) promoted state $2p^3$ $[\uparrow|\uparrow|\uparrow]$

$2s^1$ $[\uparrow]$

↓

hybridized state. $2p^1$ $[\uparrow]$ unhybridized → π bond.

$2sp^2$ $[\uparrow|\uparrow|\uparrow]$ hybridized → σ bond.

Homework, SF_4

CH_3OH

SO_4^{2-}