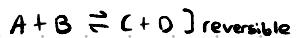
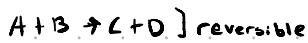


Reversible Reactions and Equilibrium

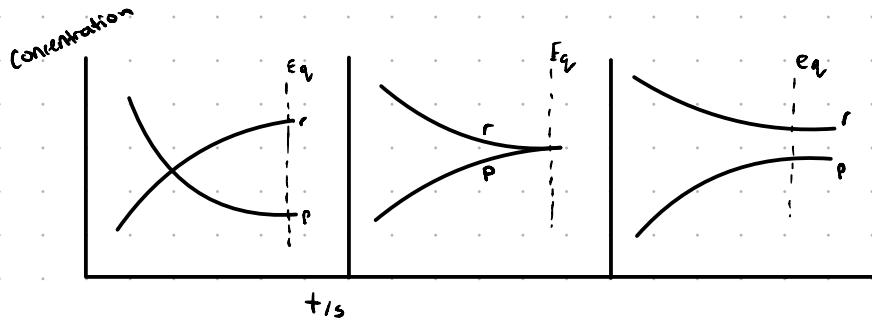


At Start:

- $[A]$ and $[B]$ are decreasing
- $[C]$ and $[D]$ are increasing

After time:

- non-reversible: $[A]$ and $[B]$ decrease, rate \downarrow . Almost all $A + B$ is converted.
- reversible: After enough $C + D$ formed, reaction can also go reverse.
Eventually, concentrations will remain constant.
- Some chemical reactions are irreversible such as combustion.
- Many reactions can occur in both directions at the same time. They are said to be reversible and will eventually reach equilibrium.
- The relative concentrations of reactants and products in equilibrium varies by reaction.



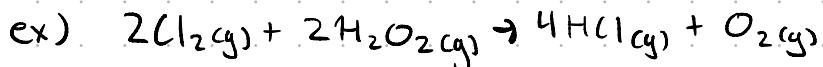
- For a closed chemical equilibrium system, the same equilibrium concentrations are reached regardless of the direction you start from (reactants vs. products).
- Chemical equilibrium is when opposing reactions are occurring at the same rate. For equilibrium to occur, the conditions must be met:
 - ↳ The process is reversible and is in constant flux at a microscopic level.
 - ↳ Observable chemical properties are constant (color, viscosity, etc.).
 - ↳ The system is closed (temp const).
 - ↳ It can be reached from either direction.

Equilibrium Constant

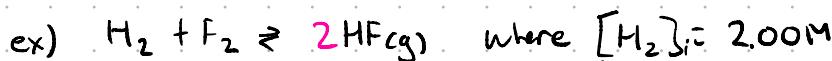
- At equilibrium, there is a constant ratio between concentrations of the products over the reactants because the forward rate = the backwards rate.
 - K is known as the equilibrium constant. It is only valid at a specific temperature. The bigger K is, the more products have formed compared to reactants.

$K = \frac{[C][D]}{[A][B]}$

- $K < 10^{-4}$ → mostly reactants
- $10^{-4} < K_u < 10^4$ → about the same
- $K > 10^4$ → mostly products



$$K = \frac{[HCl]^4 [O_2]}{[Cl_2]^2 [H_2O_2]^2} \quad \begin{array}{l} \text{(Products)} \\ \text{(Reactants)} \end{array}$$



				$[F_2]_i = 2.00M$
I	2.00M	2.00M	0M	$[HF]_i = 0M$
C	-x	-x	+2x	$[F_2]_e = 0.48M$
E	$2.00M - x$	$2.00M - x$	$2x$ $(0.48M)$	

$$\text{Solve for } x: 2.00\text{M} - x = 0.48\text{M}$$

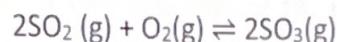
$$x = 1.52\text{M}$$

$$[F_2]_c = 2.00\text{M} - x = 0.48\text{M}$$

$$[HF]_c = 2(x) = 2 \cdot 1.52\text{M} = 3.04\text{M}$$

Now you try:

Consider the reaction:



If 2.5 mol of sulfur dioxide gas and 2.0 mol of oxygen gas are placed in a sealed 1.0L container and allowed to reach equilibrium, 0.75 mol of sulfur dioxide remains. Use an ICE table to determine the concentration of the other gases at equilibrium.

$$\begin{array}{c}
 2\text{SO}_2 + \text{O}_2 \rightleftharpoons 2\text{SO}_3 \\
 \text{I} \quad 2.5 \text{M} \quad 2.0 \text{M} \quad 0 \\
 (\quad -2x \quad -x \quad +2x \\
 \text{E} \quad 0.75 \text{M} \quad 2.0 \text{M} - x \quad 2x \\
 \quad \quad \quad (2.5 \text{M} - 2x)
 \end{array}
 \quad
 \begin{aligned}
 0.75 \text{M} &= 2.5 \text{M} - 2x \\
 1.75 \text{M} &= 2x \\
 x &= 0.875 \text{M} \\
 [\text{O}_2] &= 2.0 \text{M} - x \\
 &= 2.0 \text{M} - 0.875 \text{M} \\
 &= 1.125 \text{M} \\
 [\text{SO}_3] &= 2x \\
 &= 2(0.875) \\
 &= 1.75 \text{M}
 \end{aligned}$$

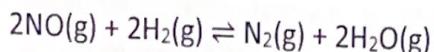
Now combine what you've learned about the equilibrium law and ICE tables

Given the equation $\text{H}_2 + \text{I}_2 \rightleftharpoons 2\text{HI}$, determine the equilibrium concentrations when the starting concentrations of H_2 and I_2 are both 0.200 mol/L and $K=64$

$$\begin{array}{c}
 \text{H}_2 + \text{I}_2 \rightleftharpoons 2\text{HI} \\
 \text{I} \quad 0.2 \text{M} \quad 0.2 \text{M} \quad 0 \\
 (\quad -x \quad -x \quad +2x \\
 \text{E} \quad 0.2 \text{M} - x \quad 0.2 \text{M} - x \quad 2x
 \end{array}
 \quad
 \begin{aligned}
 K &= \frac{[\text{HI}]^2}{[\text{H}_2][\text{I}_2]} \\
 64 &= \frac{(2x)^2}{(0.2 \text{M} - x)(0.2 \text{M} - x)} \\
 \dots \text{graph:}
 \end{aligned}$$

$$\begin{aligned}
 [\text{H}_2]_e &= 0.2 \text{M} - 0.1699 \\
 &= 0.0301 \text{M} \\
 [\text{I}_2]_e &= 0.0301 \text{M} \\
 [\text{2HI}] &= 2x = 2(0.1699) \\
 &= 0.3398
 \end{aligned}
 \quad
 \begin{aligned}
 x &= 0.1699 \text{ and } 0.2429
 \end{aligned}$$

Initially, a mixture of 0.100 M NO, 0.050 M H₂, 0.100 M H₂O was allowed to reach equilibrium (initially there was no N₂). At equilibrium the concentration of NO was found to be 0.062 M. Determine the value of the equilibrium constant, K, for the reaction:



$$\begin{array}{cccc} 0.1\text{M} & 0.05\text{M} & 0\text{M} & 0.1\text{M} \\ -2x & -2x & +x & +2x \\ 0.1\text{M}-2x & 0.05-2x & x & 0.1+2x \\ (0.062) & 0.012 & 0.019 & 0.138 \end{array}$$

$$0.1\text{M}-2x = 0.062\text{M}$$

$$-2x = -0.038$$

$$x = 0.019$$

$$\begin{aligned} K &= \frac{[\text{H}_2\text{O}]^2 [\text{N}_2]}{[\text{H}_2]^2 [\text{NO}]^2} \\ &= \frac{(0.138)^2 (0.019)}{(0.012)^2 (0.062)^2} \end{aligned}$$

$$= 653.68$$

$$\therefore K = 653 \text{ (3sf)}$$

Practice Questions:

Page 436 #1, #2, #3

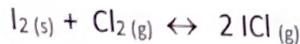
Define "chemical equilibrium"

Equilibrium Practice Problems

1. Write the equilibrium law equations for each of the following reactions



$$K = \frac{[\text{NH}_3]^2}{[\text{N}_2][\text{H}_2]^3}$$

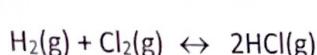


$$K = \frac{[\text{ICl}]^2}{[\text{Cl}_2][\text{I}_2]}$$



$$K = \frac{[\text{NO}][\text{O}_2]^{\frac{1}{2}}}{[\text{NO}_2]} =$$

2. Consider the following reaction



$$K = \frac{[\text{HCl}]^2}{[\text{H}_2][\text{Cl}_2]}$$

a) What is the equilibrium law equation?

b) If we start with 1.0M H₂ and 1.0M Cl₂, and we know the value of K is 4.0, what are the concentrations of all entities at equilibrium?

I	1.0M	1.0M	0M
(-x	-x	+2x
E	1-x	1-x	2x

$$K = \frac{(2x)^2}{(1-x)(1-x)}$$

$$4.0 = \frac{4x^2}{1-2x+x^2}$$

$$4x^2 = 4.0(1-2x+x^2)$$

$$4x^2 = 4.0 - 8x + 4x^2$$

$$x^2 = 1 - 2x + x^2$$

$$0 = 1 - 2x$$

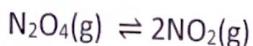
$$\frac{1}{2} = \frac{2x}{2}$$

$$x = \frac{1}{2}M$$

$$[\text{Cl}_2] = [\text{H}_2] = 1-x \\ = 1-0.5 \\ = 0.5M$$

$$[\text{HCl}] = 2x \\ = \frac{2}{1} \cdot \frac{1}{2} \\ = 1M$$

3. Consider the following reaction



a) What is the equilibrium law equation?

$$K = \frac{[\text{NO}_2]^2}{[\text{N}_2\text{O}_4]}$$

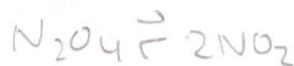
b) If 0.5 mol of N_2O_4 is placed in a 1L closed container, what will be the concentrations at equilibrium? The equilibrium constant is 4.50.

$$\text{N}_2\text{O}_4 = 0.5 \text{ mol}$$

$$K_e [\text{N}_2\text{O}_4]_e = ?$$

$$K = 4.50$$

$$K_e [\text{NO}_2]_e = ?$$



$$K = 4.50$$

$$I \quad 0.5 \text{ M} \quad 0$$

$$= \frac{[\text{NO}_2]^2}{[\text{N}_2\text{O}_4]}$$

$$C \quad -x \quad +2x$$

$$E \quad 0.5 - x \quad 2x$$

$$4.50 = \frac{(2x)^2}{0.5 - x}$$

$$x = \frac{-b \pm \sqrt{b^2 - 4ac}}{2a}$$

$$4.50(0.5 - x) = 4x^2$$

$$= \frac{-4.50 \pm \sqrt{4.50^2 - 4(4)(-2.25)}}{2(4)} = 8$$

$$2.25 - 4.50x = 4x^2$$

$$4x^2 + 4.50x - 2.25 = 0$$

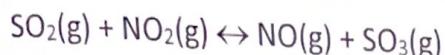
$$= [0.375] \text{ or } x = -1.5$$

$$K \text{ quadratic}$$

$$[\text{N}_2\text{O}_4]_e = 0.5 - 0.375 = 0.125 \text{ M}$$

$$[\text{NO}_2]_e = 2x = 0.75 \text{ M}$$

6. A chemist is studying the reaction



a) What is the equilibrium law equation?

$$K = \frac{[\text{NO}][\text{SO}_3]}{[\text{SO}_2][\text{NO}_2]} = \frac{[\text{NO}][\text{SO}_3]}{[\text{SO}_2][\text{NO}_2]}$$

b) In a 1L container, the chemist added 0.17 mol of $\text{SO}_2(\text{g})$ to 0.11 mol of NO_2 . The equilibrium constant, K, for a certain temperature is 4.8. What is the equilibrium concentration of SO_3 ?

$$[\text{SO}_2] = 0.17 \text{ M}$$

$$K = 4.8$$

$$[\text{NO}_2] = 0.11 \text{ M}$$

	SO_2	NO_2	\rightleftharpoons	NO	SO_3
I	0.17M	0.11M		0	0
C	-x	-x		+x	+x
E	0.17-x	0.11-x		x	x

$$4.8 = \frac{(x)(x)}{(0.17-x)(0.11-x)}$$

$$4.8 = \frac{x^2}{0.0187 - 0.17x - 0.11x + x^2}$$

$$= \frac{x^2}{x^2 - 0.28x + 0.0187}$$

$$4.8x^2 = -1.344x + 0.08976 =$$

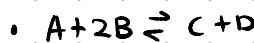
... quadratic formula ...

$$x = \boxed{0.089366}, x = 0.2643183$$

Le Chatelier's Principle

- If dynamic equilibrium is disturbed, the position of equilibrium shifts to counteract and reestablish equilibrium.

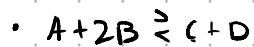
(ex:)



- ↳ Add A, rightward.
- ↳ Add C, leftward.
- ↳ Decrease A, leftward.



- ↳ Exothermic reaction, heat as product.
- ↳ Heat up, leftward
- ↳ Cool down, rightward
- Heating reaction favors side where heat is a reactant.

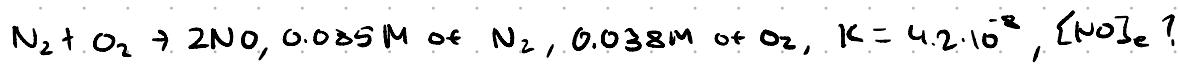


- ↳ Increase pressure, rightward, (less molecules on right)
- ↳ Decrease pressure, leftward.



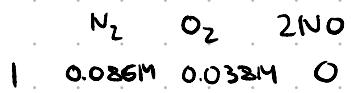
- ↳ Catalysts.
- ↳ Presence of inert gas.
- ↳ State of reactants,

Approximations



$$K = \frac{[NO]^2}{[N_2][O_2]}$$

$$K = \frac{(2x)^2}{(0.085-x)(0.038-x)}$$



When $K \leq 10^{-5}$, ignore $+/- x$. Negligible.

$$4.2 \cdot 10^{-8} = \frac{4x^2}{0.085 \cdot 0.038}$$
$$x = 5.324 \cdot 10^{-6}$$

$$\begin{aligned}[NO]_e &= 2x \\ &= 2(5.324 \cdot 10^{-6}) \\ \therefore [NO]_e &= 1.16 \cdot 10^{-5} M \end{aligned}$$

Reaction Quotient (Q)

- K applies when all components are at equilibrium. Q applies anywhere.

$$Q = \frac{[C][D]}{[A][B]}$$

- Q compared with K determines where the reaction will go.

↳ $K < Q \rightarrow$ Backwards to reach equilibrium.

↳ $K = Q \rightarrow$ Rxn does not shift. At equilibrium.

↳ $K > Q \rightarrow$ Forwards to reach equilibrium.

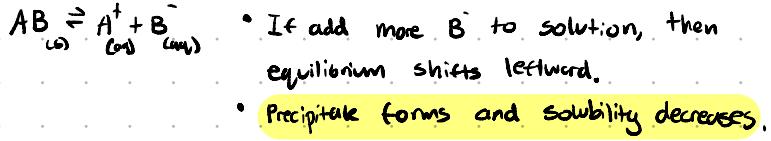
Solubility

- Solubility is the quantity of solute that dissolves in a solvent under specific conditions.
- Similar to reversible reactions, ionic compounds reach a state of dynamic equilibrium when dissolved in water.

$$K_{sp} = \frac{[C][D]}{[A][B]^3}$$
, ignoring solids.
At this point, equilibrium, no more ions are dissolving.

- Molar solubility is how many moles can be dissolved before it becomes saturated. When it is saturated, it is at equilibrium.
 - ↳ [] at equilibrium basically (E in ICE)
- K_{sp} is value from multiplying concentrations of dissolved ions in solution at equilibrium.
 - ↳ Large - concentration of ions are high, salt is very soluble.
 - ↳ Small - concentration of ions are low, salt isn't very soluble.
- Q_{sp} represents a system not at equilibrium, $Q_{sp} + K_{sp}$ can predict where the reaction will go and if precipitate forms or not.
 - ↳ $Q_{sp} < K_{sp}$: Rightward to equilibrium, no precipitate
 - ↳ $Q_{sp} = K_{sp}$: Equilibrium (saturated), no precipitate
 - ↳ $Q_{sp} > K_{sp}$: Leftward to equilibrium, precipitate forms.

Common Ion Effect



Predicting Solubilities

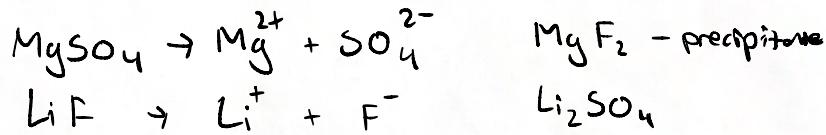
Read pages 460 – 470 in your textbook.

1. Using Table 3 (provided below), determine if the following are likely to be soluble or not (circle one):

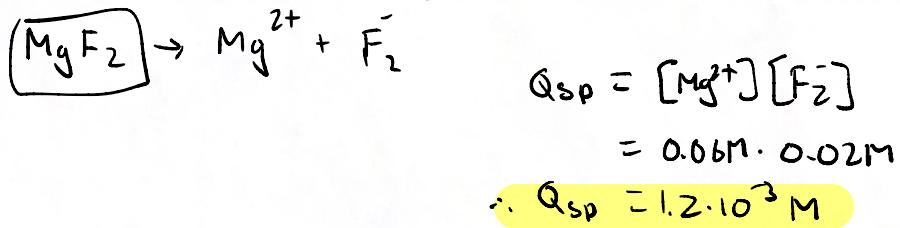
- | | | |
|------------------------------------|---------|-------------|
| a) NaNO ₃ | soluble | not soluble |
| b) Ag ₂ SO ₄ | soluble | not soluble |
| c) MgSO ₄ | soluble | not soluble |
| d) CaOH ₂ | soluble | not soluble |

2. You combine 0.06M MgSO₄ aqueous solution with 0.02M LiF aqueous solution. (Psst, if you're having issues with this question, it's very similar to tutorial 3 sample problem 1 in your textbook)

a) Could a precipitate form when the resulting ions combine? What would the precipitate be?



b) What is Q_{sp} for the precipitate? Keep in mind the concentrations used originally.



c) According to Table 1 in appendix B4 of your textbook, what is the K_{sp} value for this precipitate?

$$K_{sp} = 6.4 \cdot 10^{-9} \text{ M}$$

d) Compare the K_{sp} and Q_{sp} values. Will a precipitate form?

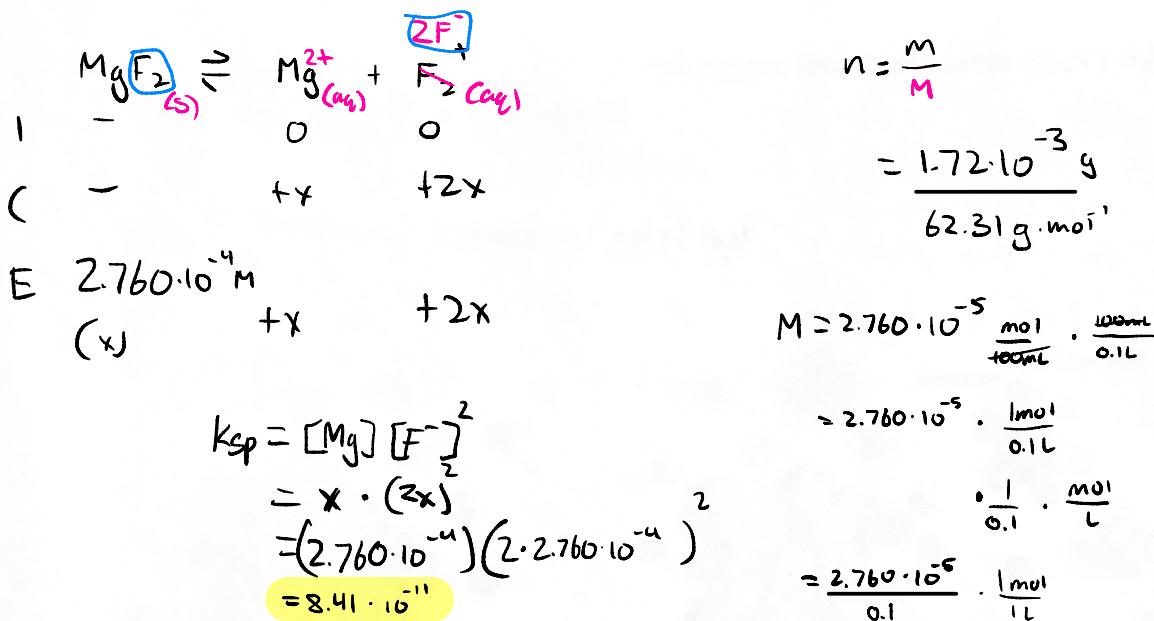
Yes, e.g. ↗

Table 3 Solubility of Some Ionic Compounds at SATP

Anions	Cations	
	high solubility ≥ 0.1 mol/L at SATP	low solubility < 0.1 mol/L at SATP
F ⁻	most	Li ⁺ , Mg ²⁺ , Ca ²⁺ , Sr ²⁺ , Ba ²⁺ , Fe ²⁺ , Hg ₂ ²⁺ , Pb ²⁺
Cl ⁻ , Br ⁻ , I ⁻	most	Ag ⁺ , Pb ²⁺ , Tl ⁺ , Hg ₂ ²⁺ , Hg ⁺ , Cu ⁺
S ²⁻	Group 1, Group 2, NH ₄ ⁺	most
OH ⁻	Group 1, NH ₄ ⁺ , Sr ²⁺ , Ba ²⁺ , Tl ⁺	most
SO ₄ ²⁻	most	Ag ⁺ , Pb ²⁺ , Ca ²⁺ , Ba ²⁺ , Sr ²⁺ , Ra ²⁺
CO ₃ ²⁻ , PO ₄ ³⁻ , SO ₃ ²⁻	Group 1, NH ₄ ⁺	most
C ₂ H ₃ O ₂ ⁻	most	Ag ⁺
NO ₃ ⁻	all	none
IO ₃ ⁻	NH ₄ ⁺ , K ⁺ , Na ⁺	most

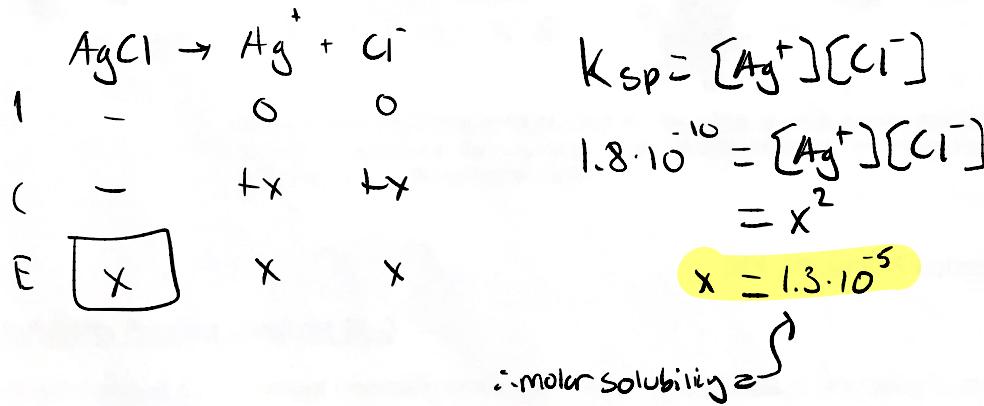
Homework:

1. The solubility of magnesium fluoride (MgF_2) is 1.72×10^{-3} g/100ml. What is the K_{sp} value of MgF_2 ? (Hint: start with converting 1.72×10^{-3} g/100ml into mol/L!)



2. The K_{sp} of silver chloride ($AgCl(s)$) is 1.8×10^{-10} at 25C.

a) What is the molar solubility of solid silver chloride in pure water?



b) What is the molar solubility of silver chloride in a 0.1M aqueous solution of sodium chloride (NaCl)?

$$K_{sp} = 1.8 \cdot 10^{-10}$$

	$\text{AgCl} \rightleftharpoons$	Ag^+	Cl^-
I	-	0	0.1M
C	-	+x	+x
E	x	x	0.1M+x

$$\begin{aligned} K_{sp} &= [\text{Cl}^-][\text{Ag}^+] \\ 1.8 \cdot 10^{-10} &= (0.1M + x)(x) \\ 1.8 \cdot 10^{-10} &= 0.1x + x^2 \\ \frac{1.8 \cdot 10^{-10}}{0.1} &= \frac{0.1x}{0.1} \\ \therefore \text{molar solubility} &\quad \curvearrowleft x = 1.8 \cdot 10^{-9} \end{aligned}$$

Acids and Bases

Arrhenius Theory of Acids and Bases

- Acids produce H^+ ions in water.
- Bases produce OH^- ions in water.
- ↳ e.g. HCl is an acid because $\text{HCl} \rightarrow \text{H}^+ + \text{Cl}^-$
- This assumes acid/base chemistry can only be aqueous.
- Does not consider basic molecules that don't have an OH^- group such as bleach.

Bronsted-Lowry Theory of Acids and Bases

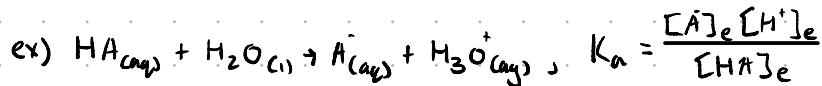
- An acid is a H^+ donor.
- A base is a H^+ acceptor.
- ↳ $\text{NH}_3 + \text{HCl} \rightleftharpoons \text{NH}_4\text{Cl}$
$$\begin{array}{c} \text{H} \\ | \\ \text{N} \\ || \\ \text{H} \end{array} + \text{H}-\text{Cl} \rightleftharpoons \left[\begin{array}{c} \text{H}^+ \\ | \\ \text{N}-\text{H}-\text{Cl}^- \end{array} \right]$$
- Amphiprotic substances both accepts and donates H^+ .
↳ For example H_2O can become H_3O^+ or OH^- .

Conjugate Acid-Base pairs

- Acid-base reactions are reversible since H^+ can happen forward and backward.
- ↳ $\text{HA}(\text{aq}) + \text{H}_2\text{O}(\text{l}) \rightleftharpoons \text{A}^-(\text{aq}) + \text{H}_3\text{O}^{(\text{aq})}$
acid base conj. base conj. acid
- Conjugate acids are acids formed when a base accepts H^+ .
- Conjugate bases are bases formed when an acid donates H^+ .

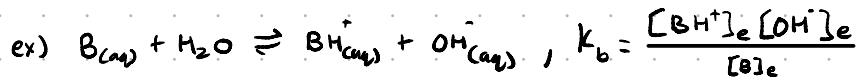
Acids

- Strong acids will dissociate completely into their respective ions.
- Weak acid and water will form dynamic equilibrium with H_3O^+ and conj. base, A^- .
- Equilibrium law equation (K_a) ignores pure H_2O .



Bases

- Most strong bases contain OH^- which will dissociate completely in water.
- Weak bases don't contain hydroxide ion directly but will produce OH^- in solution.
- Dissociation of bases represented by K_b similar to acids by K_a .



Water

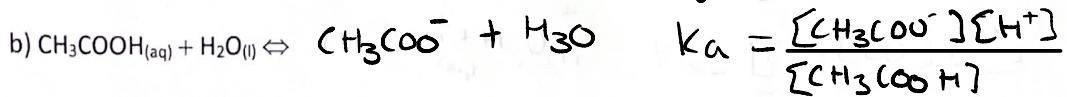
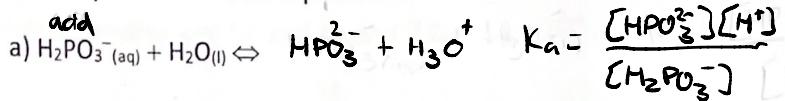
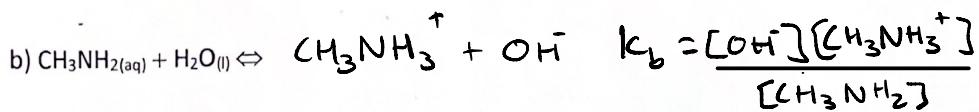
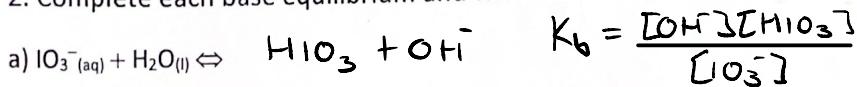
- Water has a unique equilibrium expression because it switches from OH^- to H_2O to H_3O^+ .
- $2\text{H}_2\text{O} \rightleftharpoons \text{H}_3\text{O}^+ + \text{OH}^-$, $K_w = [\text{H}^+] [\text{OH}^-]$
- @ 25°C , water has K_w or $(1.0 \cdot 10^{-7})^2 = 1.0 \cdot 10^{-14}$
- $K_w = K_A \cdot K_B$

Summary

- Strong acid/base has a very weak conjugate.
- Weak acid/base has a weak conjugate.
- Very weak acid/base has a strong conjugate.
- $K_a = \frac{[\text{H}^+] [\text{A}^-]}{[\text{HA}]}$
- $K_b = \frac{[\text{BH}^+] [\text{OH}^-]}{[\text{B}]}$
- $K_w = [\text{H}^+] [\text{OH}^-]$
- $K_w = K_A \cdot K_B$



Acid-Base Worksheet

1. Complete each acid equilibrium and write the K_a expression2. Complete each base equilibrium and write the K_b expression3. Perchloric acid, HClO_4 , is a strong acid. Calculate the pOH of a 0.025M solution of perchloric acid.

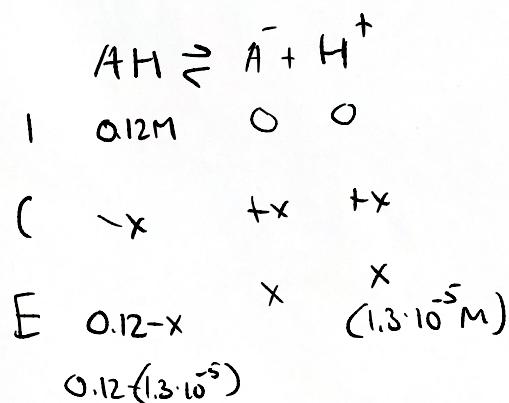
$$\begin{aligned} & -\log(0.025\text{M}) \\ & = 1.6 \\ & 14 - 1.6 \\ & = 12.4 \end{aligned}$$

4. a) The $[\text{H}^+]$ concentration of a 0.12 M solution of a weak acid is 1.3×10^{-5} M. What is the pH of the solution?

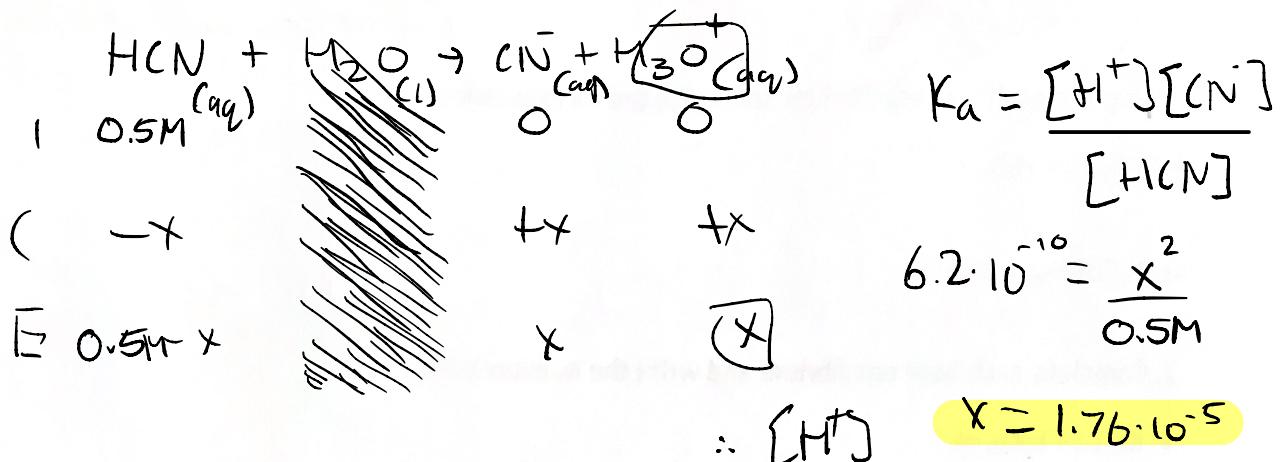
$$\begin{aligned} \text{pH} &= -\log(\text{H}^+) \\ &= -\log(1.3 \cdot 10^{-5} \text{M}) \\ &= 4.9 \quad (2 \text{ s.f.}) \end{aligned}$$

b) What is the K_a for this ~~base~~
~~acid~~

$$\begin{aligned} K_a &= \frac{[\text{H}^+][\text{A}^-]}{[\text{HA}]} \\ &= \frac{(1.3 \cdot 10^{-5})^2}{0.12 - (1.3 \cdot 10^{-5})} \end{aligned}$$

 \approx $\approx 1.41 \cdot 10^{-4}$ 

5. What is the $[H^+]$ and the pH of a 0.5M solution of HCN if the K_a value of HCN is 6.2×10^{-10} .



$$\therefore \text{pH} = -\log(1.76 \cdot 10^{-5}) \\ = 4.75$$

6. 500mL of water is used to dissolve 10g of NaOH.

a) What is the $[OH^-]$ and pOH ?
 $(NaOH)_{0.5M}$

$$\text{Na} \geq \text{Na}^+ \text{OH}^- \quad -\log(0.5) \\ \therefore 0.5M \approx \text{pOH} 0.3$$

$$n = \frac{m}{M} \\ = \frac{14.8}{40 \text{ g/mol}} \\ = 0.25 \text{ mol} \\ = 0.5 \text{ mol} \cdot L^{-1}$$

b) What is the $[H^+]$ and pH ?

$$\text{pH} = 14 - 0.3$$

$$\therefore \text{pH} = 13.7$$

$$-\log_{10}([H^+]) = 13.7$$

$$\log_b a = x \\ b^x = a$$

$$\frac{-13.7}{10} = [H^+]$$

$$[H^+] = 2 \cdot 10^{-14}$$

pH and pOH of Acidic and Basic Solutions

$$\text{pH} = -\log [\text{H}^+] \quad \text{pOH} = -\log [\text{OH}^-] \quad [\text{H}^+][\text{OH}^-] = 1.0 \cdot 10^{-14} \quad \text{pH} + \text{pOH} = 14 \quad 10^{-\text{pH}} = [\text{H}^+]$$
$$10^{-\text{pOH}} = [\text{OH}^-]$$

- Acids increase the concentration of H^+ ions while bases increase concentration of OH^- ions.
- More acid to pure water will make $\text{pH} < 7$.



$$\begin{aligned}\text{pH} &= -\log [\text{H}^{\text{(aq)}}] \\ &= -\log (0.010) \\ &= 2\end{aligned}$$
$$\begin{aligned}\text{pOH} &= 14 - \text{pH} \\ &= 12\end{aligned}$$
$$\begin{aligned}[\text{OH}^-] &= 10^{-\text{pOH}} \\ &= 10^{-12}\end{aligned}$$

- In neutral conditions:

$$4[\text{H}^+] = 1.0 \cdot 10^{-7} \text{ M}, \text{pH} = 7$$

$$4[\text{OH}^-] = 1.0 \cdot 10^{-7} \text{ M}, \text{pOH} = 7$$

- In acidic conditions:

$$\begin{aligned}4[\text{H}^+] &> 1.0 \cdot 10^{-7} \text{ M}, \text{pH} < 7 \\ 4[\text{OH}^-] &< 1.0 \cdot 10^{-7} \text{ M}, \text{pOH} > 7\end{aligned}$$

- In basic conditions:

$$\begin{aligned}4[\text{H}^+] &< 1.0 \cdot 10^{-7} \text{ M}, \text{pH} > 7 \\ 4[\text{OH}^-] &> 1.0 \cdot 10^{-7} \text{ M}, \text{pOH} < 7\end{aligned}$$

Titration

- In titration, sample placed in receiving flask and titrant in burette.
- Titrations help determine concentration of a substance.
- It does this by using an indicator that will change colors at equivalence point.

Equivalence point: - When neutralization occurs, $[H^+]$ and $[OH^-]$ is equal. When this happens, we say we have reached the endpoint and the indicator will change colors.

Strong acids/bases: The equivalence point is always at pH 7.

Weak acids/bases: The equivalence point is not always at pH 7. This is because weak acids conjugate strong bases strong enough to affect pH and vice versa for bases and their conjugate acids.

- When a weak acid is titrated with a strong base, pH will be above 7. Weaker acid \rightarrow higher equivalence point.
- When a weak base is titrated with a strong acid, pH will be below 7. Weaker base \rightarrow lower equivalence point.

Buffers

- Buffers exist to keep pH of solutions constant (or within a narrow range), and to limit fluctuations in pH when small amounts of H^+ or OH^- are added.
- Blood has a bicarbonate buffer system to resist changes in pH.
- Buffer solutions contain a weak acid/base and its conjugate acid/base.

Now you try:

Consider a 30mL 0.5M NaOH sample, and 0.3M HCl titrant.

What is the pH of the sample before we start titrating?

$$\text{NaOH} \rightleftharpoons \text{Na}^+ \text{OH}^-$$

$$0.5\text{M}$$

$$\text{pOH} = -\log(0.5)$$

$$= 0.3$$

$$\text{pH} = 14 - 0.3$$

$$\therefore \text{pH} = 13.7$$

What is the pH of the sample after we add 12mL of titrant?

$$n_{\text{HCl}} = 0.0036\text{ mol}$$

$$n_{\text{NaOH}} = 0.5\text{M} \cdot 0.03\text{L} = 0.015\text{mol}$$

$$n_{\text{excess}} = 0.0114$$

$$\xrightarrow{\frac{0.0114\text{ mol NaOH}}{0.042\text{ L}}} \text{pOH} = -\log(\text{---})$$

$$= 0.2714285714\text{ M}$$

$$= 0.5663$$

$$\text{pH} = 13.4383$$

If we know the concentrations of our sample and our titrant, we can determine when the equivalence point will be reached.

What volume of 3M HI (hydroiodic acid) must be added to a 15mL sample of a 2.5M KOH (potassium hydroxide) solution to reach the equivalence point?