Chapter 9

Chapter 9: Example: expansion work

Calculate the amount of work (kJ) done during a synthesis of ammonia in which the volume contracts from 8.6 Liters (initial) to 4.3 Liters (final) at a constant external pressure of 44 atm.

w = - P Δ V = - (44 atm)* (4.3 Liters – 8.6 Liters) wrong units

1 liter atm = 101.33 Joule ← memorize this
(use this slide instead)

w = - (44 atm)* (-4.3 liters) = 189.2 liter atm

$$w = 189.2 \text{ liter atm} * \frac{101.33 \text{ Joule}}{1 \text{ liter atm}} = 19171.636 \text{ Joule}$$

$$w = 19171.636 \text{ Joule} \times \frac{1 \text{ kJ}}{1000 \text{ J}} = 19 \text{ kJ}$$

Calorimetry and Heat Capacity

$q = c\ m\ \Delta T$ $\qquad c$ = *specific heat capacity (J/g ^oC)*

$q = C_m$ mole ΔT $\qquad C_m$ = *molar heat capacity J/mol^{o}C*

Watch units, put equations on to 3” x 5”,
get chart with c or C_m values

HW 9.5: Hess's Law

Want enthalpy for the following reaction:

$CO\ (g) + \frac{1}{2}\ O_2\ (g) \rightarrow CO_2\ (g) \qquad \Delta H^o = ?$

Have the following equations and enthalpies:

(a) $2\ CO\ (g) \rightarrow 2\ C\ (s) + O_2\ (g) \qquad \Delta H^o = 221.08\ kJ$

(b) $CO_2\ (g) \rightarrow C\ (s) + O_2(g) \qquad \Delta H^o = +393.51\ kJ$

Use **Hess's Law** to calculate the enthalpy for the reaction above.

(a) ½ of rxn (a)

$CO\ (g) \rightarrow \cancel{C}\ (s) + \cancel{\frac{1}{2}}\ O_2\ (g) \qquad \Delta H^o = +\ 110.54\ kJ$

(b) Reverse direction of reaction

$\cancel{C}\ (s) + \cancel{O_2}(g) \rightarrow CO_2\ (g) \qquad \Delta H^o = -\ 393.51\ kJ$

$CO\ (g) + \frac{1}{2}\ O_2\ (g) \rightarrow CO_2\ (g) \qquad \Delta H^o = -\ 282.97\ kJ$

HW 9.6: Standard Heats of Formation

$$\boxed{\Delta H^\circ = \Delta H^\circ_{\mathrm{f}}\ \text{(Products)} - \Delta H^\circ_{\mathrm{f}}\ \text{(Reactants)}}$$

$$aA + bB \longrightarrow cC + dD$$

$$\underbrace{2\,Cl_2\,(g) + 2\,H_2O(l)}_{\text{Reactants}} \rightarrow \underbrace{4\,HCl\,(g) + O_2\,(g)}_{\text{Products}}$$

$$\left.\begin{array}{l} \Delta H^\circ_{\mathrm{f}}\,[HCl\,(g)] = -92.31\ \text{kJ/mol} \\ \Delta H^\circ_{\mathrm{f}}\,[H_2O\,(l)] = -285.8\ \text{kJ/mol} \end{array}\right\} \begin{array}{l} \text{from chart Appendix B} \\ \text{from your text} \end{array}$$

HW 9.6: Standard Heats of Formation – generalized equation

$$\boxed{\Delta H^\circ = \Delta H^\circ_f\,\text{(Products)} - \Delta H^\circ_f\,\text{(Reactants)}}$$

$$aA + bB \longrightarrow cC + dD$$

$$\underbrace{2\,\mathrm{Cl_2}(g) + 2\,\mathrm{H_2O}(l)}_{\text{Reactants}} \rightarrow \underbrace{4\,\mathrm{HCl}(g) + \mathrm{O_2}(g)}_{\text{Products}}$$

$$\Delta H^\circ = \{[4\,\Delta H^\circ_f\,(\mathrm{HCl}(g))] + [\Delta H^\circ_f\,(\mathrm{O_2}(g))]\} - \{[2\,\Delta H^\circ_f\,(\mathrm{Cl_2}(g))] + [2\,\Delta H^\circ_f\,(\mathrm{H_2O}(l))]\}$$

$$\Delta H^\circ = \{[4\,(-92.31)] + [(0.00)]\} - \{[2\,(0.00)] + [2\,(-285.8)]\}$$

$$\Delta H^\circ = \{4\,(-92.31)\} - \{[2\,(-285.8)]\}$$

$$\Delta H^\circ = \{4\,(-92.31)\} - \{[2\,(-285.8)]\} =$$

$$-369.24 + 571.6 = 202.4\,\mathrm{kJ}$$

An Introduction to Free Energy

Gibbs Free-Energy Change (ΔG)

$$\boxed{\Delta G = \Delta H - T\,\Delta S}$$

Enthalpy of Reaction
kJ/mol
(- ΔH spontaneous)
reaction releases heat

Temperature
(in kelvin)

Entropy Change
kJ/mol
(+ΔS spontaneous)
reaction more disordered

<u>Spontaneous</u>
(rxn goes →)
(← nonspontaneous)

ΔG = negative
ΔH = negative
ΔS = positive

End 1/15 Wed
C section

An Introduction to Free Energy

Gibbs Free-Energy Change (ΔG)

$\Delta G = \Delta H - T\,\Delta S$

high T: ΔS decides
low T: ΔH decides

$\Delta G < 0$ Process is spontaneous. (rxn goes →)

$\Delta G = 0$ Process is at equilibrium
(neither spontaneous nor nonspontaneous).

$\Delta G > 0$ Process is nonspontaneous. (rxn goes ←)

Chapter 11

HW 11.2: Phase Changes between Solids, Liquids, & Gases

ΔH_{fusion} solid to a liquid ΔH_{vap} liquid to a gas

At phase change (melting, boiling, etc) :
$\Delta G = \Delta H - T\Delta S$ & ΔG = zero (bc 2 phases in equilibrium)
$\Delta H = T\Delta S$

a. BP of ethanol is 78.4°C $\Delta H_{vap} = 38.56$ kJ/mol. What is the entropy change for the vaporization (ΔS_{vap} in J/(Kmol) ?

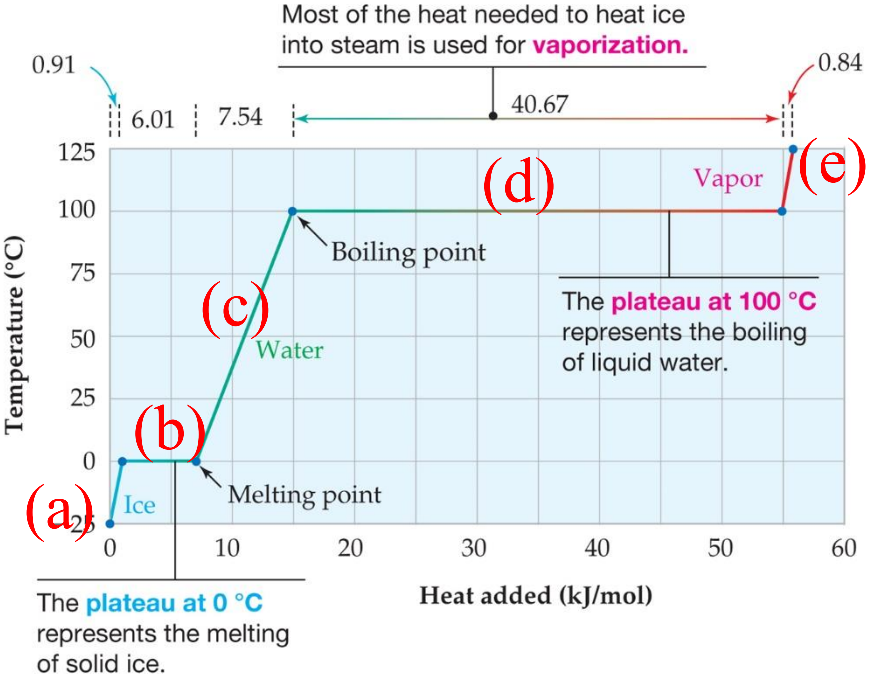
$\Delta S = \Delta H/T = 38.56$ kJ/mol / (78.4°C + 273.15) K =
$\Delta S = 38.56$ kJ/351.55 (mol K) = 0.10969 kJ/(mol K) = 110 J/(mol K)

b. $CHCl_3$ has $\Delta H_{vap} = 29.2$ kJ/mol and $\Delta S_{vap} = 87.5$ J/(K mol). What is the BP of the $CHCl_3$ in Kelvin ?

End C sect 1/21 Tuesday

Phase Changes between Solids, Liquids, and Gases

Heating Curve for Water



Most of the heat needed to heat ice into steam is used for **vaporization.**

The **plateau at 100 °C** represents the boiling of liquid water.

The **plateau at 0 °C** represents the melting of solid ice.

(a) Heat ice from -25 to 0^{o}C: molar heat capacity ice = 36.57 J/(mol ^{o}C)

(36.57 J/mol ^{o}C)($T_f – T_i$) = 0.914 kJ/mol

(b) *Melting ice: ΔH_{fusion} = +6.01 kJ / mol

(c) Heating liquid water from 0^{o}C to 100 ^{o}C : molar heat capacity water = 75.4 J/(mol ^{o}C)

(75.4 J/mol oC)* (100^{o}C) = 7.54 x 10^3 J/mol = 7.54 kJ /mol

(d) *Vaporizing liquid water: ΔH_{vap} = +40.67 kJ/mol largest energy for process

(e) Heating water vapor from 100^{o}C to 125^{o}C:
Molar heat capacity of water vapor = 33.6 J/mol ^{o}C
33.6 J/mol ^{o}C * 25^{o}C = 0.840 kJ/mol

$$q = n * C_m * \Delta T$$
$$q = n \Delta H_{fus}$$
$$q = n \Delta H_{vap}$$

For the same molecule:

Largest ΔS for gaseous form of compound
Medium ΔS for liquid form of compound
Smallest ΔS for solid form of compound

For a Reaction:

Largest ΔS for the side of the reaction with larger number of gaseous molecules.

ex: $2\,H_2\,(g) + O_2\,(g) \rightarrow 2\,H_2O\,(g)$

$$\Delta S_{\text{reactant}} > \Delta S_{\text{product}}$$

Kinds of Solids

Amorphous Solids: Particles are randomly arranged and have no ordered long-range structure.
examples: rubber, window glass, plastic, butter

Crystalline Solids: Particles have an ordered arrangement extending over a long range.

- ionic solids (ex: NaCl, BaF_2)
- molecular (covalent molecule) solids (ex: $C_{12}H_{26}$, I_2, naphthalene, sucrose)
- covalent network solids (ex: carbon, silicon dioxide)
- metallic solids (ex: metallic Fe, metallic Na, metallic Pt, etc)

End 1/27 Monday A section

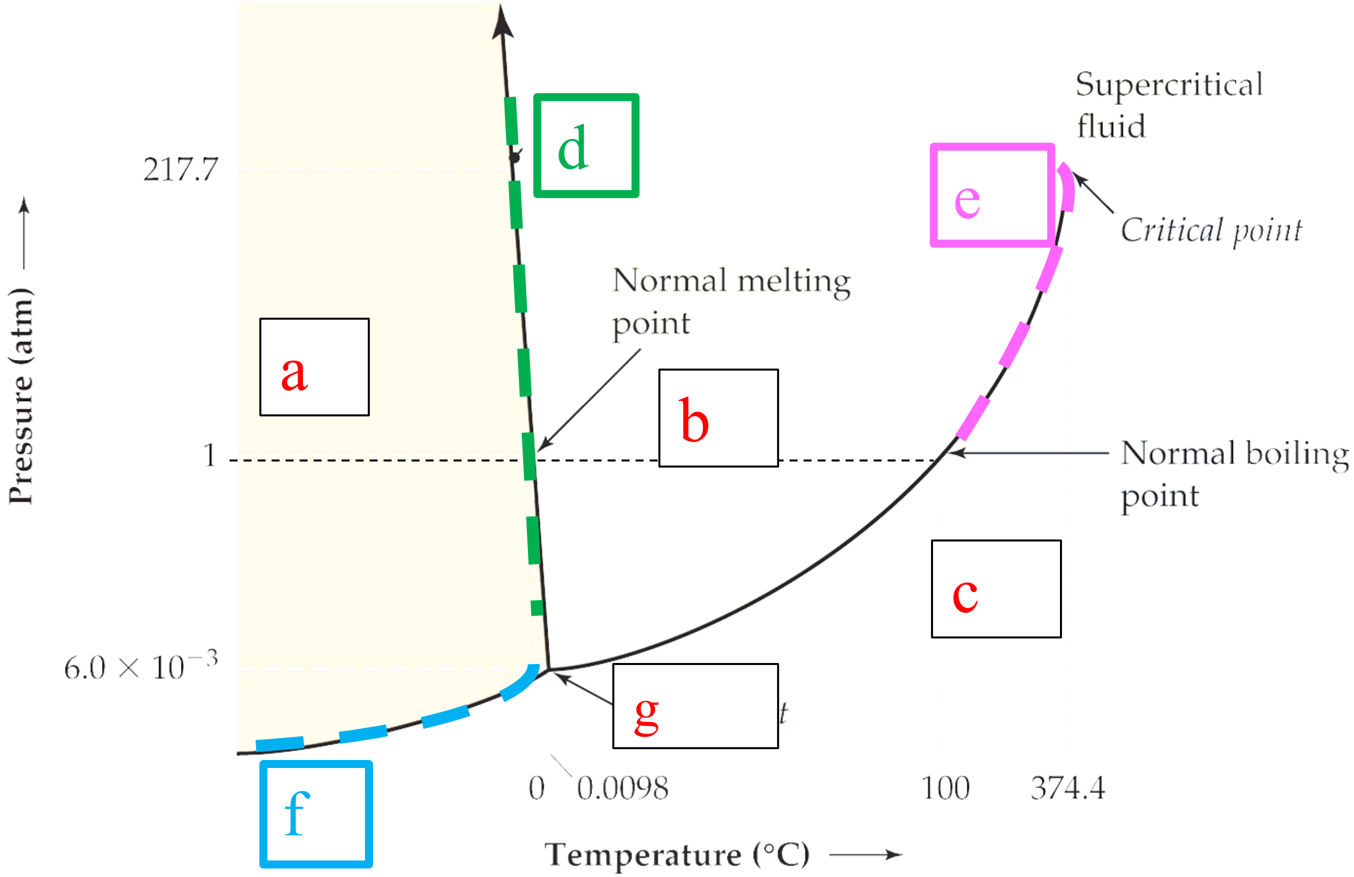
HW 11.5: Phase Diagrams **Fill in the blank with a letter**

(a) Solid
(b) Liquid
(c) Gas

(d) line transition s → l
(e) line transition l → g
(f) line transition s → g

(g) Triple Point

Water



Supercritical Fluid: A state of matter beyond the critical point that is neither liquid nor gas

Chapter 12

HW 12.1: Energy Changes and the Solution Process

Like dissolves Like

Match the following for solutes which will BEST dissolve in solvents. Fill in the blank with letters. Blank may have more than one letter. (asking for **most soluble** – may be soluble in more than one solvent)

solute

(a) CH_3CH_2OH (b) $CH_3CH_2CH_2CH_2CH_3$

(c) NaCl (d) CH_2Cl_2

solvents

(1) $CH_3CH_2CH_2CH_2CH_2CH_3$ b______

(2) H_2O a, c______

(3) CH_3Cl d______

Concentration Units for Solutions

$$\textbf{Molarity (M)} = \frac{\text{Moles of solute}}{\text{Liters of solution}}$$

$$\textbf{Mole fraction (}\boldsymbol{X}\textbf{)} = \frac{\text{Moles of component}}{\text{Total moles making up solution}}$$

$$\textbf{Mass percent} = \frac{\text{Mass of component}}{\text{Total mass of solution}} \times 100\%$$

$$\textbf{Molality (}\boldsymbol{m}\textbf{)} = \frac{\text{Moles of solute}}{\text{Mass of solvent (kg)}}$$

Physical Behavior of Solutions: Colligative Properties

Colligative Properties: Properties that depend on the amount of a dissolved solute but not on its chemical identity (usually non volatile solute)

- Vapor-pressure lowering ($P_{solution} = i * P_{solvent} * X_{solvent}$) (Raoult’s Law)
- Boiling-point elevation ($\Delta T_b = i * K_b * m$)
- Freezing-point depression ($\Delta T_f = - i * K_f * m$)
- Osmotic pressure ($\Pi = i * M * R * T$)

X = mole fraction

i = van’t Hoff factor (=1 for nonelectrolytes) (= # of ions in solution per formula)

m = molality

K_b molal BP constant, K_f molal FP constant

M= molarity

R = gas constant = 0.08206 L atm /(mol K)

End 2/7 Friday A section, C section

Chapter 13

Reaction Rates

General Reaction:

$$a\,\mathrm{A} + b\,\mathrm{B} \longrightarrow d\,\mathrm{D} + e\,\mathrm{E}$$

$$\text{rate} = -\frac{1}{a}\frac{\Delta[\mathrm{A}]}{\Delta t} = -\frac{1}{b}\frac{\Delta[\mathrm{B}]}{\Delta t} = \frac{1}{d}\frac{\Delta[\mathrm{D}]}{\Delta t} = \frac{1}{e}\frac{\Delta[\mathrm{E}]}{\Delta t}$$

disappearing reactant
(negative sign)

appearing product
(positive sign)

reaction:

$$2\,\mathrm{N_2O_5}(g) \longrightarrow 4\,\mathrm{NO_2}(g) + \mathrm{O_2}(g)$$

$$\text{rate} = -\frac{1}{2}\frac{\Delta[\mathrm{N_2O_5}]}{\Delta t} = \frac{1}{4}\frac{\Delta[\mathrm{NO_2}]}{\Delta t} = \frac{\Delta[\mathrm{O_2}]}{\Delta t}$$

disappearing
reactant

appearing product

Rate Laws and Reaction Order

Rate Law: An equation that shows the dependence of the reaction rate on the concentration of each reactant

$$aA + bB \longrightarrow \text{products}$$

$$\text{rate} \propto [A]^m[B]^n$$

$$\text{rate} = k[A]^m[B]^n$$

k is the **rate constant**.

$$m,n \neq a,b$$

HW 13.3: Experimental Determination of a Rate Law

Given the following experimentally determined rate law: what is the overall order of the reaction ?

(a) Rate = $k[A][B]^2$ 3

(b) Rate = $k\ [BrO_3^-][Br^-][H^+]^2$ 4

(c) Rate = $k[A]^2$ 2

End 2/14
Friday C
section

Note additional questions not in HW:
Order of [B] in (a) is ? 2
Order of [A] in (a) is ? 1

Order of [A] in (c) is ? 2

(integrated rate law – summary of last few slides)

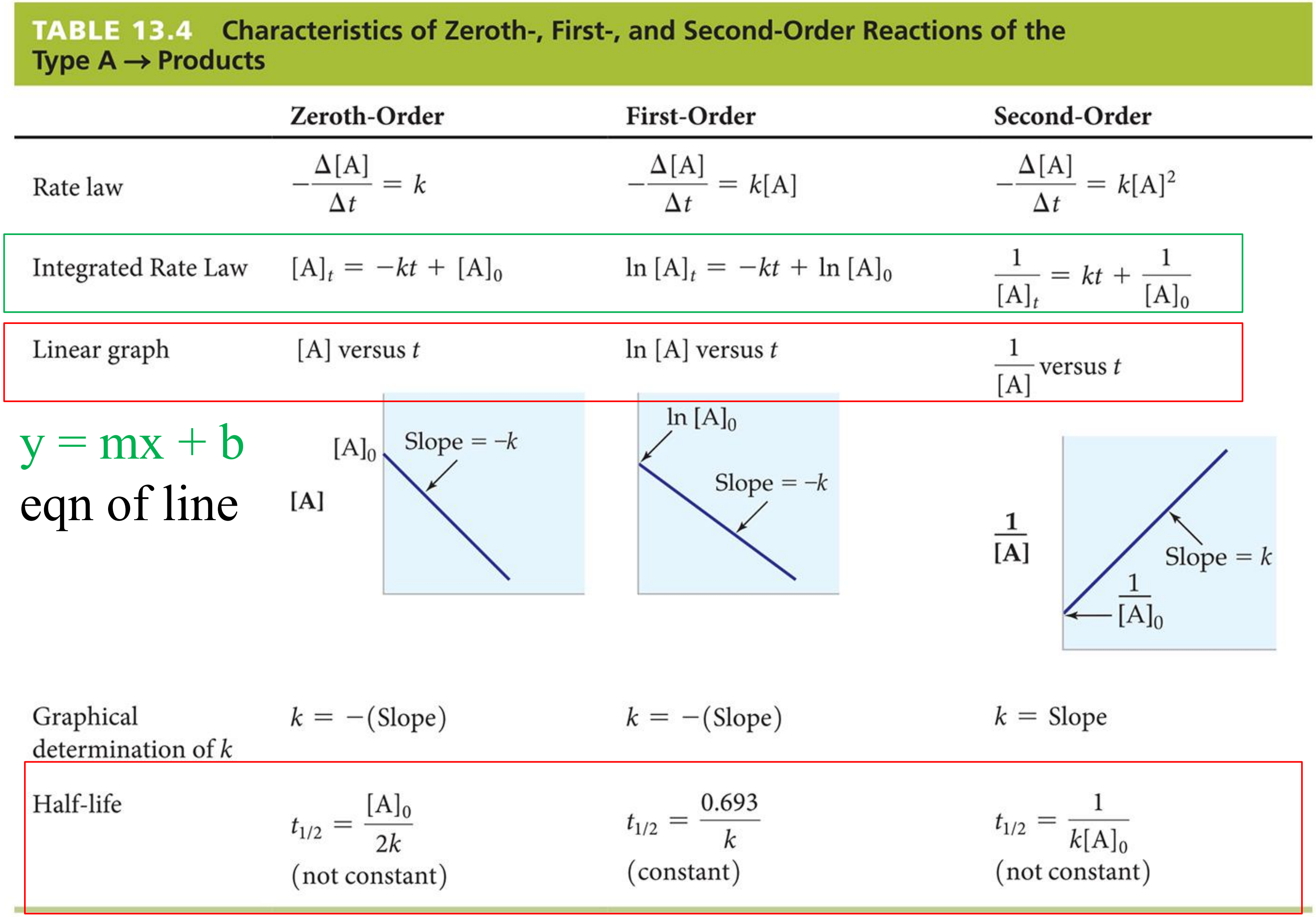


TABLE 13.4 Characteristics of Zeroth-, First-, and Second-Order Reactions of the Type A → Products

	Zeroth-Order	First-Order	Second-Order
Rate law	$-\frac{\Delta[A]}{\Delta t} = k$	$-\frac{\Delta[A]}{\Delta t} = k[A]$	$-\frac{\Delta[A]}{\Delta t} = k[A]^2$
Integrated Rate Law	$[A]_t = -kt + [A]_0$	$\ln [A]_t = -kt + \ln [A]_0$	$\frac{1}{[A]_t} = kt + \frac{1}{[A]_0}$
Linear graph	[A] versus t	ln [A] versus t	$\frac{1}{[A]}$ versus t
	[A] $[A]_0$ Slope = $-k$	$\ln [A]_0$ Slope = $-k$	$\frac{1}{[A]}$ Slope = k $\frac{1}{[A]_0}$
Graphical determination of k	$k = -(\text{Slope})$	$k = -(\text{Slope})$	$k = \text{Slope}$
Half-life	$t_{1/2} = \frac{[A]_0}{2k}$ (not constant)	$t_{1/2} = \frac{0.693}{k}$ (constant)	$t_{1/2} = \frac{1}{k[A]_0}$ (not constant)

$y = mx + b$

eqn of line

HW 13.7: Rate Laws for Elementary Reactions (reaction mechanism steps)

For the following elementary reactions (reaction mechanism steps) give the rate law and the molecularity of the reaction mechanism step (unimolecular, bimolecular, trimolecular, etc.)

(a) $CH_3Br + OH^- \rightarrow Br- + CH_3OH$ — rate = $k[CH_3Br][OH^-]$ bimolecular

(b) $O_3 \rightarrow O_2 + O$ — rate = $k[O_3]$ unimolecular

(c) $H_2O_2 + OH^- \rightarrow H_2O + HO_2$ — rate = $k[H_2O_2][OH^-]$, bimolecular

(d) $NO_2 + F \rightarrow NO_2F$ — rate = $k[NO_2][F]$ bimolecular

(e) $2\ ClO \rightarrow Cl_2O_2$ — rate = $k\ [ClO]^2$ bimolecular

End 2/24 Monday A section

Chapter 14

HW 14.1: The Equilibrium Constant K_c

Write the equilibrium constant expression for the following reactions. (K_p = equilibrium constants with amounts are given by pressure of each component)(if gas is in solution use K_c, if gas is in gas state use K_p)

$$aA + bB \rightleftharpoons cC + dD \qquad K_c = \frac{[C]^c[D]^d}{[A]^a[B]^b}$$

(s) or (l) leave out of K expression

(a) $2\ BrCl\ (soln) \rightleftarrows Br_2\ (soln) + Cl_2\ (soln)$

$$K_c = \frac{[Br_2]\ [Cl_2]}{[BrCl]^2}$$

(b) $2\ N_2\ (g) + 2\ O_2\ (g) \rightleftarrows 4\ NO\ (g)$ (give K_p)

$$K_p = \frac{(P_{NO})^4}{(P_{N2})^2(P_{O2})^2}$$

(c) $BaSO_4\ (s) \rightleftarrows Ba^{2+}\ (aq) + SO_4^{\ 2-}(aq)$

$$K_c = [Ba^{2+}]\,[SO_4^{\ 2-}]$$

(d) $Fe_2O_3\ (s) + 3\ CO(g) \rightleftarrows 2\ Fe\ (s) + 3\ CO_2(g)$

$$K_p = \frac{(P_{CO2})^3}{(P_{CO})^3}$$

2/28 Friday
C section

The Equilibrium Constant K_c

The equilibrium constant and the equilibrium constant expression are for the chemical equation *as written*.

$$aA + bB \rightleftharpoons cC + dD \qquad K_c = \frac{[C]^c[D]^d}{[A]^a[B]^b}$$

End 2/28F A section

(a) $N_2(g) + 3\,H_2(g) \rightleftharpoons 2\,NH_3(g)$ $\qquad K_c = \frac{[NH_3]^2}{[N_2][H_2]^3}$

reverse rxn

double rxn

(b) $2\,NH_3(g) \rightleftharpoons N_2(g) + 3\,H_2(g)$ $\qquad K_c' = \frac{[N_2][H_2]^3}{[NH_3]^2} = \frac{1}{K_c}$

reversed rxn (a) - turn K upside down

(c) $2\,N_2(g) + 6\,H_2(g) \rightleftharpoons 4\,NH_3(g)$ $\qquad K_c'' = \frac{[NH_3]^4}{[N_2]^2[H_2]^6} = K_c^{\,2}$

doubled rxn (a) – square all terms

HW 14.4: The Equilibrium Constant K_c

$K_p = K_c(RT)^{\Delta n}$ **R** is the gas constant, $0.08206 \dfrac{\text{L atm}}{\text{K mol}}$.

$\Delta n = (c+d) - (a+b)$

End 3/2 Monday
A section

$H_2O\ (g) + CH_4(g) \rightleftarrows CO\ (g) + 3H_2\ (g)$

If $K_c = 3.8 \times 10^{-3}$ at 1000 K, what is the value of K_p ?

$K_p = K_c\ (RT)^{\Delta n}$

$\Delta n = (1+3) - (1+1) = 2$

$K_p = (3.8 \times 10^{-3})\ [(0.08206\ \text{L atm/mol K})(1000\ \text{K})]^2$

$K_p = 25.588 = 26$

HW 14.5: Using the Equilibrium Constant

- If $Q_c < K_c$ net reaction goes from left to right (*reactants to products,* →).
- If $Q_c > K_c$ net reaction goes from right to left (*products to reactants,* ←).
- If $Q_c = K_c$ no net reaction occurs. (at equilibrium)

If the equilibrium constant $K_c = 2.78 \text{ x } 10^{-10}$ and you have a reaction $A + 2\,B \rightarrow 3\,C + D$ and $[A] = 1.78 \text{ x } 10^{-4}$, $[B] = 1.1 \text{ x } 10^{-5}$ and $[C] = 1.23 \text{ x } 10^{-6}$, $[D] = 7.8 \text{ x } 10^{-8}$ does the reaction go forward to produce more products or backwards to produce more reactant or stay in equilibrium ?

$$Kc = \frac{[C]^3[D]}{[A]\,[B]^2} \qquad Q = \frac{[1.23 \text{ x } 10^{-6}]^3[7.8 \text{ x } 10^{-8}]}{[1.78 \text{ x } 10^{-4}]\,[1.1 \text{ x } 10^{-5}]^2}$$

$Q = 6.74 \text{ x } 10^{-12}$ & $6.74 \text{ x } 10^{-12} < 2.78 \text{ x } 10^{-10}$

Reaction goes forward until equilibrium is reached.

End 3/2/20
Monday
section A

HW 14.6: Using the Equilibrium Constant

For $H_2(g) + I_2(g) \rightleftharpoons 2HI(g)$, for initial concentrations of $H_2(0.00623\ M)$, I_2 $(0.00414\ M)$ and HI $(0.0224\ M)$. Calculate the concentrations at equilibrium given K_c (54.3) assuming forward rxn.

(a) Set up the ICE table. (where x = change of H_2)

(b) show the expression for Kc

(c) If your solution to Kc gives x = 0.00156, calculate the concentration of HI.

$H_2(g) + I_2(g) \rightleftharpoons 2\ HI(g)$

0.00623	0.00414	0.0224	Initial
$-x$	$-x$	$+2x$	Change
0.00623-x	0.00414-x	*0.0224 + 2x*	Equilibrium

$$Kc = \frac{[0.0224+2x]^2}{[0.00623-x][0.00414-x]} = 54.3$$

Solution x = 0.00156 $[HI] = \mathbf{0.0224 + 2(0.00156) = 0.0255}\ \boldsymbol{M}$

HW 14.7: Le Châtelier's Principle: If a stress is applied to a reaction mixture *at equilibrium*, net reaction occurs in the direction that relieves the stress. (equilibrium shifts to undo stress applied to system)

$2\,A(g) + B(g) \rightarrow 3\,C(g) + 2\,D(g)$ ΔH = negative (circle one parenthesis under each letter)

which direction will equilibrium shift

hint: ΔH negative is exothermic = (+ heat, heat is product)

(a) add A rxn goes [[✓ (→)] or (←)]

(b) remove B rxn goes [(→) or [✓ (←)]]

(c) add C rxn goes [(→) or [✓ (←)]]

(d) remove C rxn goes [[✓ (→)] or (←)]

(e) higher T rxn goes [(→) or [✓ (←)]] end Test 3 (e) & below

(higher T = add heat)

(f) higher P rxn goes [(→) or [✓ (←)]]

(assume all reactants & products in reaction are gases) (moves to fewer moles gas)

(g) higher V rxn goes [[✓ (→)] or (←)]

(assume all reactant & products in reaction are gases) (moves to fewer moles gas)

Chapter 15

HW 15.1: Acid-Base Concepts: Arrhenius & Brønsted-Lowry Theory (conjugate acid/base)

Given the following equation of acid/base dissociation. Give the Acid/Conjugate Base, Base/Conjugate Acid pairs.

$H\ NO_3 + H_2O \rightarrow NO_3^- + H_3O^+$

Acid $H\ NO_3$ Conjugate Base NO_3^-
Base H_2O Conjugate Acid H_3O^+

$NH_3 + H_2O \rightarrow NH_4^+ + OH^-$

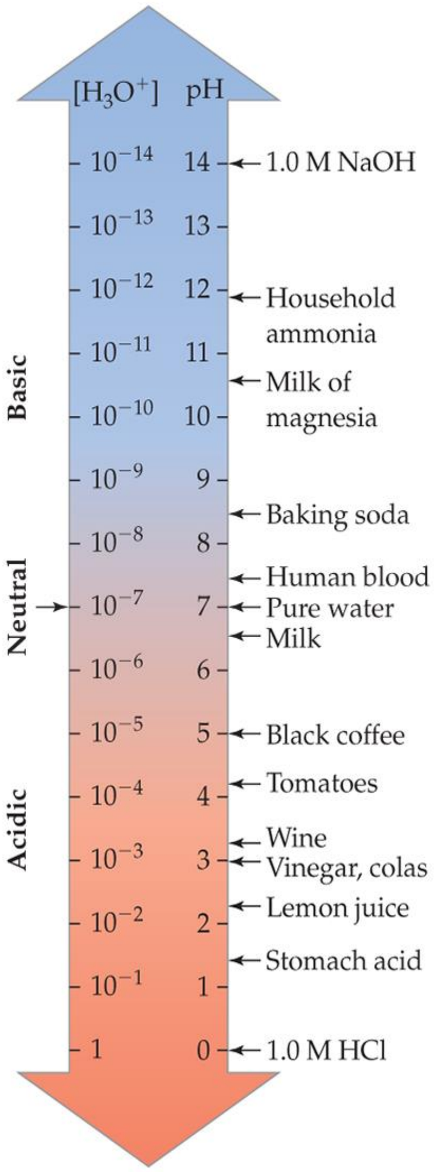
Base NH_3 Conjugate Acid NH_4^+
Acid H_2O Conjugate Base OH^-

$H\ NO_3 + NaOH \rightarrow Na\ NO_3 + H_2O$

Acid $H\ NO_3$ Conjugate Base NO_3^-
Base NaOH Conjugate Acid H_2O

Do this HW by emailing the answer in the text of your email instead of taking a photo of your answer on paper.

The pH Scale



$pH = -\log[H_3O^+]$ $[H_3O^+] = 10^{-pH}$

antilog

Acidic: pH < 7

Neutral: pH = 7

Basic: pH > 7

The pH Scale Equations

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

$$pOH = -\log [OH^-]$$

$$pK_w = -\log K_w$$

$$K_w = [H^+][OH^-] = 1 \times 10^{-14}$$

$$pK_w = pH + pOH = 14$$

HW 15.2: The pH Scale

(a) What is the pH of a solution of $[H^+] = 2.3 \times 10^{-2}$

pH = – log $[H^+]$ = - log(2.3×10^{-2}) = -(-1.638) = 1.64

(b) What is the pOH of a solution of [OH-] = 7.7×10^{-3}

pOH = - log $[OH^-]$ = -log(7.7×10^{-3}) = -(-2.11) – 2.11

(c) What is the pH of the solution in (b) above ?

pH + pOH = 14 pH + 2.11 = 14 pH= 11.9

(d) What is the $[H^+]$ of the solution in (c) above ?

$[H^+]$ = antilog (-pH) = antilog(-11.9) = 1.26×10^{-12}

Answer this HW by writing the text of the answer into your email instead of uploading photo.

HW 15.3: The pH in Solutions of Strong Acids and Strong Bases

Just use concentration of strong acid or base bc dissociates completely.

a. What is the pH of a 0.053 M solution of HCl ?

pH = -log(0.053) = (1.28)

a. What is the pH of a solution of 0.150 M solution of NaOH ? pOH= -log (0.150M) = - (-0.823)

pH= 14 - 0.823 = 13.2

Answer this HW by writing the text of the answer into your email instead of uploading photo.

Equilibria in Solutions of Weak Acids

$$HA(aq) + H_2O(l) \rightleftharpoons H_3O^+(aq) + A^-(aq)$$

Acid-Dissociation Constant: $K_a = \dfrac{[H_3O^+][A^-]}{[HA]}$

For Weak Acids and Weak Bases –
MUST USE K_a / K_b

HW 15.4: Calculating Equilibrium Concentrations of Weak Acids

Calculate the pH of a 0.10 M HCN solution. At 25 °C, $K_a = 4.9 \times 10^{-10}$.

$$HCN(aq) + H_2O(l) \rightleftharpoons H_3O^+(aq) + CN^-(aq)$$

0.10		≈0	0
$-x$		$+x$	$+x$
$0.10 - x$		x	x

$$K_a = \frac{[H_3O^+][CN^-]}{[HCN]}$$

Chapter 16

HW 16.1: Neutralization Reactions

(a) Give the neutralization reaction for the following acid / base reaction

$HF + KOH \rightarrow H_2O + KF$ molecular rxn

$HF + \cancel{K^+} + OH^- \rightarrow H_2O + \cancel{K^+} + F^-$ total ionic

$HF + OH^- \rightarrow H_2O + F^-$ net ionic rxn

(b) Is the product of the neutralization.
[(acidic) [✓ (basic)] (neutral)] (choose one)

K + is neutral, F- is weak base so weak base

The Common-Ion Effect

Calculate the pH of a solution that is prepared by dissolved 0.10 mol of acetic acid and 0.10 mol sodium acetate in enough water to make 1.00 L of solution. $Ka=1.8 \times 10^{-5}$

$$CH_3CO_2H(aq) + H_2O(l) \rightleftharpoons H_3O^+(aq) + CH_3CO_2^-(aq)$$

0.10		≈0	0.10
$-x$		$+x$	$+x$
$0.10 - x$		x	$0.10 + x$

$$K_a = \frac{[H_3O^+][CH_3CO_2^-]}{[CH_3CO_2H]}$$

Buffer Solutions

Buffer Solution: A solution that contains a weak acid and its conjugate base and resists drastic changes in pH

Weak acid

+

Conjugate base

For example:

- CH_3CO_2H + $CH_3CO_2^-$
- HF + F^-
- NH_4^+ + NH_3
- $H_2PO_4^-$ + HPO_4^{2-}

HW 16.3: Buffer Solutions

What is the pH of a buffer solution with 0.25 M HF and 0.50 M NaF. Ka = 3.5 x 10^{-4}. What is the pH after the addition of 0.002 mol HNO_3 ?

$$HF\ (aq) + H_2O(l) \rightleftharpoons H_3O^+(aq) + F^-(aq)$$

Using Henderson – Hasselbalch:

$$pH = pKa + \log \frac{[base]}{[acid]} = -\log(3.5 \times 10^{-4}) + \log \frac{0.50}{0.25}$$

pH = 3.456 + 0.30 = 3.757

HW 16.4: Buffer Solutions

Buffer Solution: A solution that contains a weak acid and its conjugate base and resists drastic changes in pH

Given the following reagents, fill the blank with the other molecule needed to make a buffer solution and label each as (A) for acid or (B) for base.

1. CH_3CO_2H (A) $CH_3CO_2^-$ (B)

2. CN^- (B) H CN (A)

3. NH_3 (B) NH_4^+ (A)

Weak Acid–Strong Base Titrations

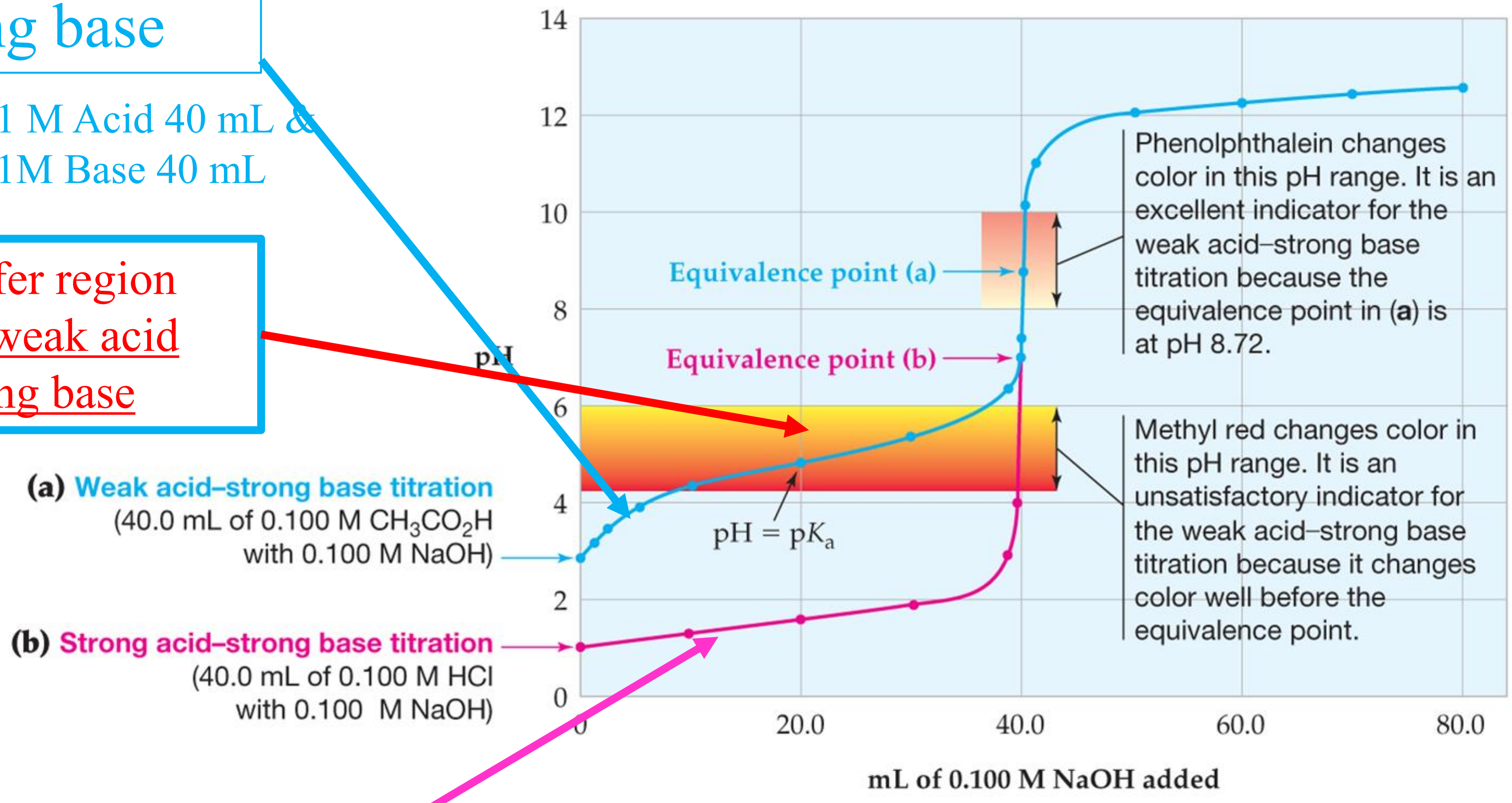
$$CH_3CO_2H(aq) + OH^-(aq) \rightleftharpoons H_2O(l) + CH_3CO_2^-(aq)$$

pH = basic at equivalence pt for weak acid/strong base titration

weak acid vs strong base

0.1 M Acid 40 mL &
0.1M Base 40 mL

Buffer region for weak acid strong base



(a) Weak acid–strong base titration (40.0 mL of 0.100 M CH_3CO_2H with 0.100 M NaOH)

(b) Strong acid–strong base titration (40.0 mL of 0.100 M HCl with 0.100 M NaOH)

Phenolphthalein changes color in this pH range. It is an excellent indicator for the weak acid–strong base titration because the equivalence point in **(a)** is at pH 8.72.

Methyl red changes color in this pH range. It is an unsatisfactory indicator for the weak acid–strong base titration because it changes color well before the equivalence point.

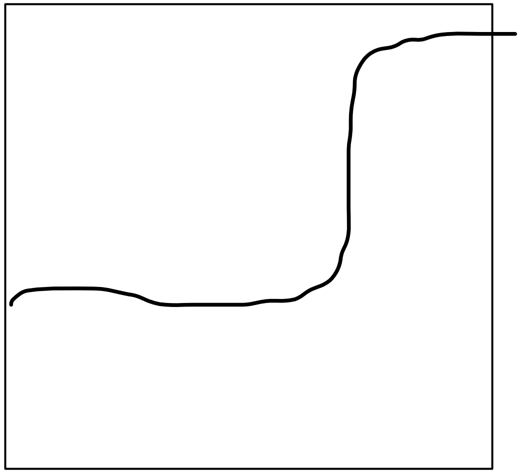
Either phenolphthalein or methyl red can be used for the strong acid–strong base titration because the curve rises very steeply in the region of the equivalence point in **(b)** at pH 7.00.

strong acid vs strong base

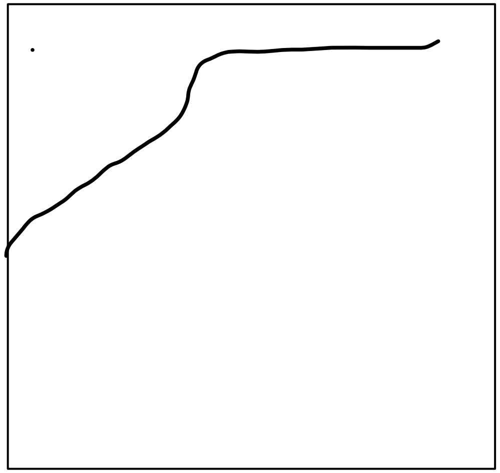
0.1 M acid 40 mL &
0.1 M base 40 mL

HW 16.4: pH Titration Curves

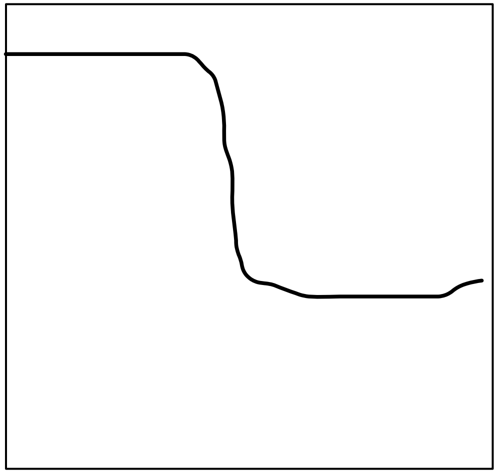
Label the following titration curve as (A) strong acid added to strong base (B) strong base added to strong acid (C) strong base added to weak acid (choose one letter under each #)



1.(A)[✓ (B)](C)



2.(A)(B)[✓ (C)]



3.[✓ (A)](B)(C)

Measuring K_{sp} and Calculating Solubility from K_{sp}

Calculate the molar solubility of MgF_2 in water at 25 °C.

$$MgF_2(s) \rightleftharpoons Mg^{2+}(aq) + 2F^-(aq)$$

	x	$2x$

$$K_{sp} = 7.4 \times 10^{-11} = [Mg^{2+}][F^-]^2 = (x)(2x)^2$$

$$4x^3 = 7.4 \times 10^{-11}$$

$$x = [Mg^{2+}] = \text{Molar solubility} = \boxed{2.6 \times 10^{-4}\ \text{M}}$$

HW 16.6: Measuring K_{sp} and Calculating Solubility from K_{sp}

Calculate the molar solubility of Ag Cl in water at 25 °C.

$$\text{Ag Cl}\,(s) \rightleftharpoons \text{Ag}^{+}(aq) + \text{Cl}^{-}(aq)$$

	x	*x*

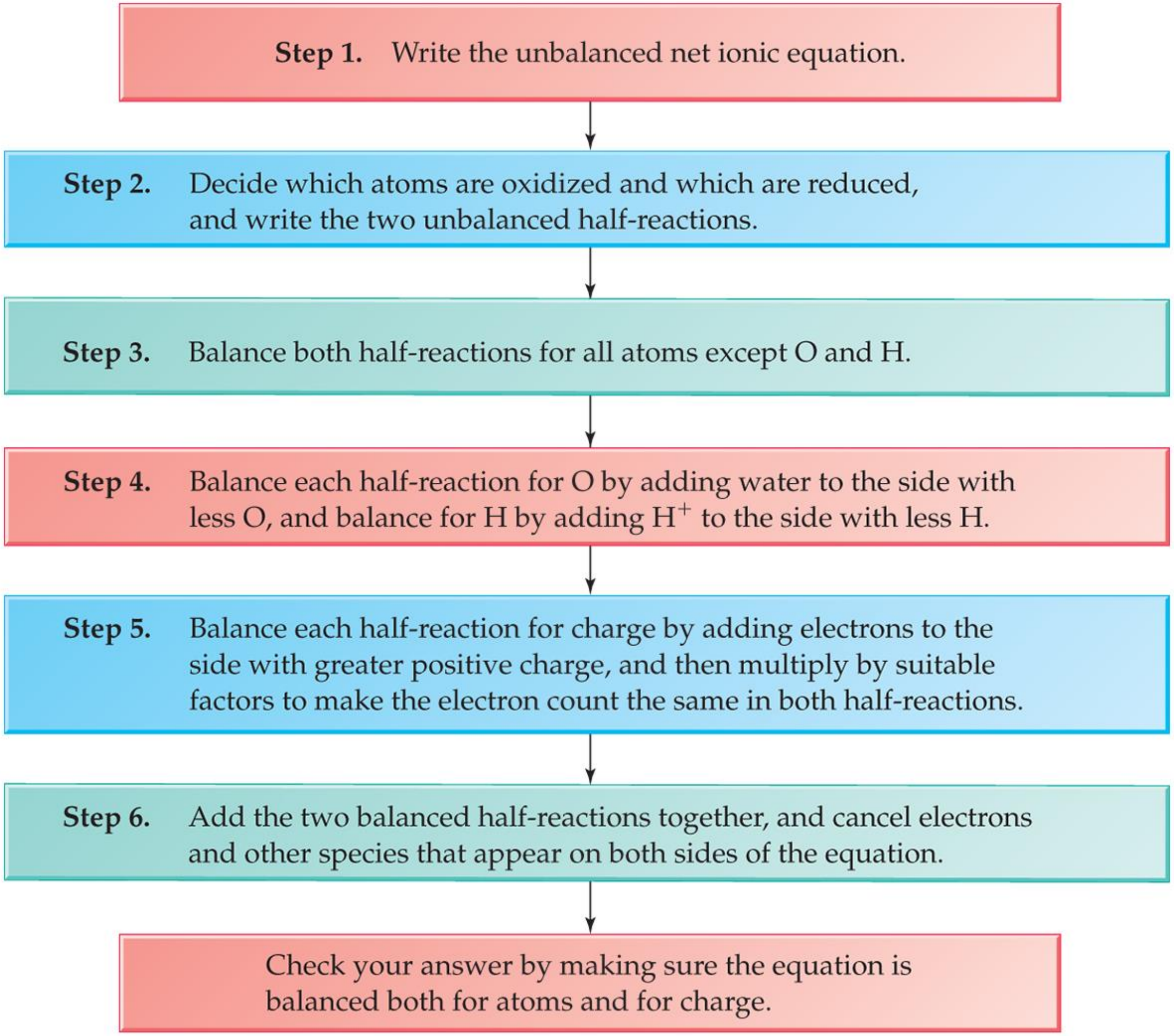
$$K_{sp} = 1.8 \times 10^{-10} = [\text{Ag}^{+}][\text{Cl}^{-}] = (x)(x)$$

$$x^{2} = 1.8 \times 10^{-10}$$

$$x = [\text{Ag}^{+}] = \text{Molar solubility} = \boxed{1.3 \times 10^{-5}\ \text{M}}$$

Chapter 18

Balancing Redox Reactions by the Half-Reaction Method (in acidic solution)



Step 1. Write the unbalanced net ionic equation.

Step 2. Decide which atoms are oxidized and which are reduced, and write the two unbalanced half-reactions.

Step 3. Balance both half-reactions for all atoms except O and H.

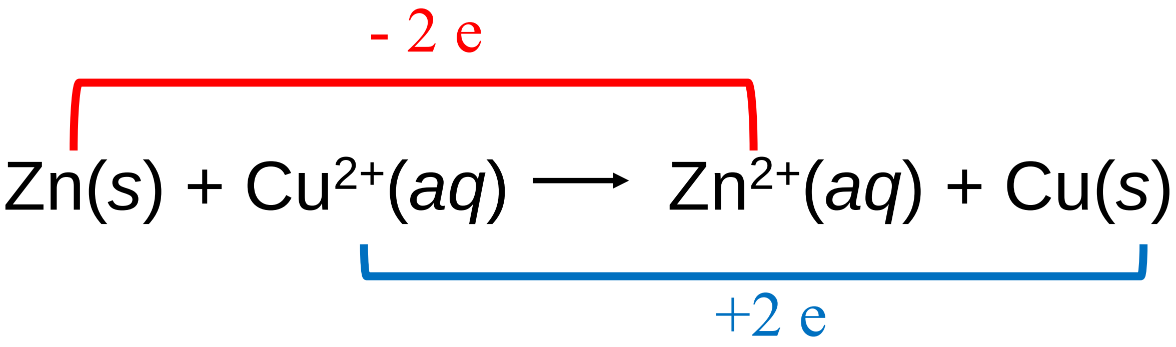
Step 4. Balance each half-reaction for O by adding water to the side with less O, and balance for H by adding H^+ to the side with less H.

Step 5. Balance each half-reaction for charge by adding electrons to the side with greater positive charge, and then multiply by suitable factors to make the electron count the same in both half-reactions.

Step 6. Add the two balanced half-reactions together, and cancel electrons and other species that appear on both sides of the equation.

Check your answer by making sure the equation is balanced both for atoms and for charge.

18.1: Balancing Redox Reactions: The Half-Reaction Method



- 2 e

$$\mathrm{Zn}(s) + \mathrm{Cu}^{2+}(aq) \longrightarrow \mathrm{Zn}^{2+}(aq) + \mathrm{Cu}(s)$$

+2 e

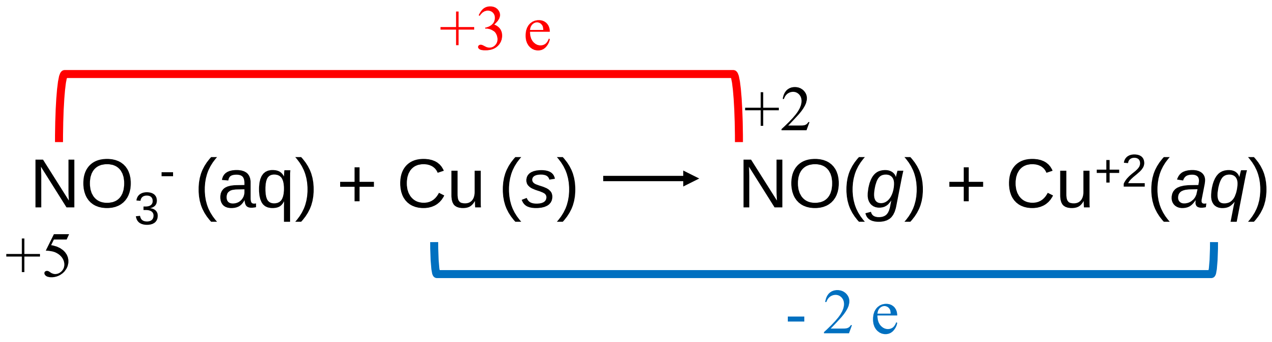
(a) Oxidation half-reaction: $\mathrm{Zn}(s) \longrightarrow \mathrm{Zn}^{2+}(aq) + 2\ e^{-}$

(b) Reduction half-reaction: $\mathrm{Cu}^{2+}(aq) + 2\ e^{-} \longrightarrow \mathrm{Cu}(s)$

Balanced redox reaction:

$$\mathrm{Zn}(s) + \mathrm{Cu}^{2+}(aq) \longrightarrow \mathrm{Zn}^{2+}(aq) + \mathrm{Cu}(s)$$

HW 18.2: Balancing Redox Reactions: The Half-Reaction Method (acidic solution)



+3 e

+2

$NO_3^-\ (aq) + Cu\ (s) \longrightarrow NO(g) + Cu^{+2}(aq)$

+5

- 2 e

(a) Oxidation half-reaction: $Cu(s) \longrightarrow Cu^{2+}(aq) + 2e^-$

(b) Reduction half-reaction: $NO_3^-(aq) + 3\ e^- \longrightarrow NO(g)$

(c) Balanced redox reaction:

$$2\ NO_3^-\ (aq) + 8\ H^+ + 3\ Cu\ (s) \rightarrow 2\ NO(g) + 3\ Cu^{+2}(aq) + 4\ H_2O\ (l)$$