

VC210 Notes

Dzsyang

November 14, 2022

Fall 2022

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1 Fundamentals

1.1 Chemistry at Three Levels

Macroscopic level: dealing with the properties of large, visible objects.

Microscopic level: dealing with the rearrangements of atoms.

Symbolic level: using terms, chemical symbols, and mathematical equations.

A chemist thinks at the microscopic level, conducts experiments at the macroscopic level, and represents both symbolically.

1.2 Nuclear Model

$$p^+ = 1.673 \times 10^{-27} \text{ kg}, n^0 = 1.675 \times 10^{-27} \text{ kg}.$$

p^+ and n^0 are 2000 times heavier than an e^- .

Isotopes of an element have the same atomic number but different mass numbers. Their nuclei have the same number of protons but different numbers of neutrons. The chemical properties of all the atom's isotopes are the same.

1.3 Properties

Extensive properties: Volume, Mass; **Intensive properties:** Density, Temperature

1.4 Energy

Kinetic is the energy that a body possesses due to motion.

Potential is the energy that a body possesses due to its position in a field of force.

Electromagnetic is the energy due to attractions and repulsions between electric charges.

1.5 Mixtures

Heterogeneous: components are visible with a microscope or an unaided eye. (Milk)

Homogeneous: well mixed into a single phase; a microscope cannot distinguish the particles. (Sugar water)

1.6 Separation

Distillation	Boiling Points
Chromatography	Absorption ability
Decanting	Density
Filtration	Particle size and solubility

2 Atomic Theory

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3 Basic Quantum Mechanics

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4 Molecular Shapes, VSEPR, VB and MO Theory

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5 Chemical Bonds

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6 Gases

6.1 The Nature of Gases

Eleven elements are gases under normal conditions: H, He, N, O, F, Ne, Cl, Ar, Kr, Xe, Rn. Low molar mass compounds such as carbon dioxide, hydrogen chloride are also gases.

6.2 The Unit of Pressure

$$1\text{Pa} = 1\text{kg} \cdot \text{m}^{-1} \cdot \text{s}^{-2}. \quad 10^5\text{Pa} = 1\text{bar}.$$
$$760\text{mmHg} = 760\text{Torr} = 1\text{atm} = 1.013 \times 10^5\text{Pa}.$$

6.3 The Gas Laws

Boyle's Pressure Experiment $PV = \text{Const.}$

Charles' Law $V/T = \text{Const.}$

No real gas has zero volume. Lowest possible temperature: -273.15°C .

6.4 Avogadro's Principle

All gases occupy the same volume under the same conditions of temperature and pressures. The molar volume of all gases is close to $22.4\text{L} \cdot \text{mol}^{-1}$ at 0°C and 1 atm.

6.5 Standard Conditions

Standard Ambient Temperature and Pressure (SATP)

SATP means exactly 25°C (298.15 K) and exactly 1 bar. The molar volume of an ideal gas is $24.79\text{L} \cdot \text{mol}^{-1}$.

Standard Temperature and Pressure (STP)

STP means 0°C (273.15 K) and 1 atm (both exactly). The molar volume of an ideal gas is $22.41\text{L} \cdot \text{mol}^{-1}$.

6.6 The Ideal Gas Law

$$pV = nRT$$

The constant $R = 8.314\text{J} \cdot \text{K}^{-1} \cdot \text{mol}^{-1}$.

The ideal gas law is a limiting law, valid only as $p \rightarrow 0$.

The differences between ideal gases and real gases are significant at high pressures and low temperatures.

6.7 Combined Gas Law

$$\frac{p_1 V_1}{n_1 T_1} = \frac{p_2 V_2}{n_2 T_2}$$

6.8 Change n for The Ideal Gas Law

Molar Concentration $M = \frac{p}{RT}$.

Gas Density $d = \frac{Mp}{RT}$.

6.9 Mixtures of Gases

A mixture of gases behaves like a single pure gas.

Dalton's Law of Partial Pressures

Dalton concluded that the total pressure is the sum of the individual pressures of each gas.

$$\chi_A = \frac{n_A}{n_A + n_B + \dots}, p_A = \chi_A p_T$$

6.10 The Kinetic Model of Gases

The kinetic model (KMT) of a gas allows us to derive the quantitative relation between pressure and the speeds of the molecules.

Root Mean Square Speed $v_{rms} = \sqrt{\frac{3RT}{M}}$.

Most Possible Speed $v_{rms} = \sqrt{\frac{2RT}{M}}$.

Effusion In effusion, molecules escape through a small hole in a barrier into a region of low pressure. And we can find $\bar{v} \propto \frac{1}{\sqrt{m}}$.

The Maxwell Distribution of Speeds $f(v) = 4\pi \left(\frac{M}{2\pi RT}\right)^{\frac{3}{2}} v^2 e^{-Mv^2/2RT}$.

For the same temperature, the greater the molar mass, the lower the speed.

For the same molar mass, the higher the temperature, the higher the average speed and the broader the spread of speeds.

6.11 Deviations from Ideality

Gases can be condensed to liquids when cooled or compressed.

A measurement of the compression factor $Z = \frac{V_{m,real}}{V_{m,ideal}}$.

At low pressures the attractive forces are dominant and $Z < 1$.

At high pressures, repulsive forces become dominant and $Z > 1$ for all gases.

6.12 van der Waals Equation

$$(p + a\frac{n^2}{V^2})(V - nb) = nRT.$$

Parameter "a" represents the attraction between molecules; the value is large for strongly attracting molecules. Parameter "b" represents the role of repulsions; it can be thought of as representing the volume.

7 Liquids and Solids

7.1 Different Types of Intermolecular Forces

Type of interaction	E_p dependence	Interacting species
ion-ion	$1/r$	ions
ion-dipole	$1/r^2$	ions and polar molecules
dipole-dipole (stationary)	$1/r^3$	stationary polar molecules
dipole-dipole (rotating)	$1/r^6$	rotating polar molecules
hydrogen bonding		special case of dipole-dipole: N-H, O-H, F-H

7.2 London Forces

Attractive forces between nonpolar molecules are London forces.

Even nonpolar noble gases can be liquefied, as well as many nonpolar compounds.

London interaction $E_p \propto \frac{\alpha_1 \alpha_2}{r^6}$.

Influence

Size: As size increases (more shells) \Rightarrow polarizability increases \Rightarrow melting and boiling points increase.

Shape: Rod-like molecules have a greater surface area, more contact points for molecules to join together (high melting and boiling points); Ball or spherical shaped molecules have fewer contact points for molecules to join together (low melting and boiling points).

7.3 Hydrogen Bonding

Very strong interaction between molecules that is specific to molecules with certain types of atoms (the second strongest only to ion-ion interaction).

A hydrogen bond is denoted by a dotted line, $X - H \dots X$, $X = N, O, F$.

7.4 Surface Tension

Viscosity is a liquid's resistance to flow: the higher the viscosity of the liquid, the more sluggish the flow. \uparrow Viscosity indicates, \uparrow intermolecular strength.

Surface tension is the reason that the surface of a liquid is smooth. Strong forces pull the molecules together, with a net inward pull.

The upward curved meniscus (concave) of water forms because both water and glass have comparable forces: Adhesion \approx Cohesion.

The downward meniscus (convex) of mercury forms because the cohesive forces in mercury is stronger than between mercury atoms and the glass: Cohesion $>$ Adhesion.

7.5 Liquid Crystals

Liquid crystals are neither a solid nor a liquid, but an intermediate called a mesophase. Here molecules have the fluidity of a liquid and some of the order of a molecular solid. They are responsive to changes in temperature and electric fields.

Anisotropic materials depend on the direction of measurement. Isotropic materials do not depend on orientation: water's viscosity is the same in all directions.

Three phases

Nematic phase: parallel molecules, staggered along their long axes.

Smectic phase: molecules are parallel and line up to form sheets.

Cholesteric phase: sheets of parallel molecules are rotated relative to their neighbors and form a helical structure.

LCD (liquid-crystal display) Light is polarized when a potential difference is applied. The molecules rotate until they are oriented with the electric field and become opaque, forming dark spots on a screen.

7.6 Solids Types

Crystalline and Amorphous Solids Crystalline solids have long-range order. Amorphous solids have short-range order.

Molecular Solids Molecular solids are held together by intermolecular forces.

Network Solids Atoms in network solids are joined to their neighbors by strong covalent bonds. Therefore, network solids are very hard, rigid materials with high melting and boiling points.

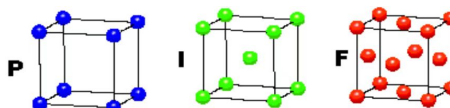
Metallic Solids Metals are cations bound tightly together by a sea of swirling electrons that the metals have lost. Many metallic structures are closed-packed (layer upon layer).

Closed-packed Structures ABABAB pattern: Hexagonal closed-pack (hcp); ABCABC pattern: Cubic closed-pack (ccp).

7.7 Unit Cells

The smallest region of the crystal lattice that repeats itself is referred to as the unit cell. The atoms in a unit cell can stack, or arrange themselves, into one of three types of cubic structures: 1. Primitive cubic; 2. Body-centered cubic; 3. Cubic closed-packed. (4. Face-centered close packed; 5. Hexagonal close packed)

CUBIC
 $a = b = c$
 $\alpha = \beta = \gamma = 90^\circ$



8 The First Law of Thermodynamics

8.1 Historical Views on Heat

Sadi Carnot Heat flowed to produce work.

James Joule Energy can be transformed into heat and/or work.

8.2 System versus Surroundings

The system and the surroundings jointly make up the universe.

System to the surroundings: Open, Close, Isolate:

Open: Exchange both matter and energy;

Close: Exchange energy;

Isolate: No contact.

8.3 Formula

$$\Delta U = Q + W. \text{ i.e., } Q = \Delta U + p\Delta V.$$

8.4 Internal Energy U

In thermodynamics, the capacity of a system to do work is called its internal energy U. Absolute internal energy is not measurable because it includes the energies of all the atoms, their electrons, and the components of their nuclei.

Internal energy is energy stored in a system as kinetic energy and potential energy.

The internal energy of an isolated system is constant.

8.5 Work

Expansion work is a change in volume of a system. $W = -p\Delta V$.

Two forms:

Isobar (Constant pressure);

Isothermal (Constant temperature) $W = -nRT \ln \frac{V_{final}}{V_{initial}}$.

Expansion against zero pressure is called free expansion.

Nonexpansion work can be the flow of electrical current.

8.6 Heat Capacity

Specific heat capacity: $Q = mC_s\Delta T$ (for water $4.18 J \cdot K^{-1} \cdot g^{-1}$);

Molar heat capacity: $Q = nC_m\Delta T$ (for water $75 J \cdot K^{-1} \cdot mol^{-1}$).

8.7 State Functions

A property that depends only on the current state of the system and is independent of how that state was prepared is called a state function.

Such as pressure, volume, temperature, mass, altitude and internal energy.

8.8 The Origin of Internal Energy

The **equipartition theorem** (not derived here) states the average value of each quadratic contribution at a temperature T is equal to $\frac{1}{2}kT$.

For monoatomic molecules: Total E_{kin} : $\frac{3}{2}RT = 3.72kJ \cdot mol^{-1}$;

For linear molecules: Total E_{kin} : $\frac{5}{2}RT = 6.02kJ \cdot mol^{-1}$;

For nonlinear molecules: Total E_{kin} : $\frac{6}{2}RT = 7.44kJ \cdot mol^{-1}$.

8.9 Enthalpy

$$H = U + pV. \text{ i.e., } \Delta H = \Delta U + p\Delta V = Q + (p - p_{ex})\Delta V.$$

For a chemical reaction open to the atmosphere, or at constant pressure, the heat released or required is the enthalpy of the system. $\Delta H = Q$. Therefore, $\Delta H < 0$ for exothermic reactions, "-"; $\Delta H > 0$ for endothermic reactions, "+".

Vaporisation $\Delta H_{vap,m} = H_m(vapor) - H_m(liquid)$, where ΔH_m is the molar heat.

Fusion $\Delta H_{fus,m} = H_m(liquid) - H_m(solid)$.

Freezing $\Delta H_{fre,m} = -\Delta H_{vap,m}$. **Condensation** $\Delta H_{con,m} = -\Delta H_{fus,m}$.

Sublimation $\Delta H_{sub,m} = H_m(vapor) - H_m(solid)$. **Deposition** $\Delta H_{dep,m} = -\Delta H_{sub,m}$.

8.10 The Relation between H and U

At constant volume, the heat transfer is ΔU ; $C_{V,m} = \frac{\Delta U}{\Delta T}$.

At constant pressure, it is ΔH ; $C_{p,m} = C_{V,m} + R$.

8.11 Hess's Law

The overall reaction enthalpy is the sum of the reaction enthalpies of each step.

8.12 The Born-Haber Cycle

In a Born-Haber cycle, we

- break apart the bulk elements into atoms,
- ionize the atoms,
- combine the gaseous ions to form the ionic solid,
- then form the elements again from the ionic solid.

9 The Second and Third Laws of Thermodynamics

9.1 Entropy and Disorder

Entropy, S , is a measure of disorder.

Low entropy means little disorder. High entropy means great disorder.

In an isolated system the entropy increases in the course of any spontaneous change.

The natural progression of a system and its surroundings is from order to disorder, from lower to higher entropy.

An entropy change in a system is calculated as: $\Delta S = \frac{q_{rev}}{T}$.

Entropy can predict the natural direction of a reaction.

9.2 Deriving a Change in Entropy

Changes in temperature: $\Delta S = C \ln \frac{T_2}{T_1}$.

Changes in volume: $\Delta S = nR \ln \frac{V_2}{V_1}$.

Changes in pressure: $\Delta S = nR \ln \frac{p_1}{p_2}$.

9.3 A Molecular Interpretation of Entropy

The entropies of all perfect crystals approach zero as the absolute temperature approaches zero.

The entropy of any substance to be greater than zero above $T = 0$.

9.4 Boltzmann Formula - Statistical Entropy

$S = k \ln W$, where $k = 1.381 \times 10^{-23} \text{ J} \cdot \text{K}^{-1}$, W is the number of positions atoms or molecules can arrange into and are called microstates.

$W = \Pi$ orientations of each molecule.

9.5 Standard Molar Entropies

$$\Delta S^\circ = \frac{\Delta H^\circ}{T}$$

Complex molecules have greater entropy values. Heavier molecules have greater entropy values.

9.6 Global Changes in Entropy: Total

$$\Delta S(\text{total}) = \Delta S(\text{system}) - \Delta S(\text{surroundings}).$$

$$\Delta S(\text{system}) = \sum n S_m(\text{products}) - \sum n S_m(\text{reactants}).$$

$$\Delta S(\text{surroundings}) = -\frac{\Delta H}{T}.$$

If $\Delta S(\text{total})$ is positive (an increase), the process is spontaneous.

If $\Delta S(\text{total})$ is negative (a decrease), the reverse process is spontaneous.

If $\Delta S(\text{total}) = 0$, the process has no tendency to proceed in either direction (phase changes are the most common examples of when $\Delta S(\text{total}) = 0$).

Clausius inequality In an isolated system $q = 0$ so $\Delta S \geq 0$, meaning the entropy cannot decrease in an isolated system, or in other words, the entropy of the universe is steadily increasing.

9.7 Equilibrium

A system at equilibrium has no tendency to change in either direction (forward or reverse). The equilibrium state is a dynamic equilibrium, where the forward and reverse processes are continually at matching rates.

9.8 Gibbs Free Energy

Gibbs free energy accomplishes the same task more simply, and it also tells us how much nonexpansion work (work under free expansion when W usually = 0) we can get from the system.

Gibbs free energy is a measure of the energy free to do nonexpansion work.

$$G = H - TS. \text{ i.e., } \Delta G = \Delta H - T\Delta S.$$

When ΔG is large "-" the process is spontaneous.

Labile versus inert

If $\Delta G_f^\circ < 0$, then elements are poised to change spontaneously into the compound.

If $\Delta G_f^\circ > 0$, then the compound is poised to change spontaneously into the pure elements (unstable).

The crossover point is where ΔG_f° goes from "+" to "-".

Therefore, we can take $\Delta G^\circ = \Delta H^\circ - T\Delta S^\circ$ set it to 0, i.e., $T = \frac{\Delta H^\circ}{\Delta S^\circ}$.

10 Physical Equilibrium

10.1 Phases and Phase Transitions

Matter exists in a single phase such as a solid, liquid, or gas. A phase change occurs when converting one phase into another.

Carbon has three distinct solid phases: diamond, graphite, and Buckminsterfullerenes (C₆₀). Helium is only known to exist as a gas and liquid.

10.2 Vapour Pressure

Evaporation takes place at the surface because molecules are bound to fewer neighbors. A dynamic equilibrium is when the rate of escaping matches the rate of returning.

Solids and gases aside, liquids with weak intermolecular forces have the highest vapour pressure; Liquids with strong intermolecular forces, ones capable of forming hydrogen bonds, have the lowest vapour pressure.

10.3 Clausius - Clapeyron Equation

$$\ln \frac{p_2}{p_1} = \frac{\Delta H_{\text{vap}}^\circ}{R} \left(\frac{1}{T_1} - \frac{1}{T_2} \right).$$

10.4 Water is Unusual

Most substances are more dense in the solid phase than liquid, water being an exception. Water is highly unusual, at 0.0°C, $\text{density}_{\text{liquid}} > \text{density}_{\text{solid}}$.

Solid water hydrogen bonds hold the molecules apart at low temperatures. As ice melts, the hydrogen bonds collapse, allowing water molecules to pack more closely.

10.5 Phase Diagram

A phase diagram is a map showing phases at different pressures and temperatures.

For water, a triple point is the temperature and pressure at which water exists as a solid, liquid, and vapour. The slope of the solid–liquid boundary depends on the relative density and for liquid water it is more dense than its solid.

Critical Properties

There is an end in the liquid–vapour phase boundary called the critical point.

The density of the vapour is so great that it is equal to the density of the liquid.

The surface boundary disappears into a single, uniform phase.

Here, the critical pressure and critical temperature mark the end of either liquid or vapour, and is now a supercritical fluid, a very dense fluid.

10.6 Raoult's Law

$p = \chi_{\text{solvent}} \cdot p_{\text{pure}}$, where p is the vapour pressure of the solvent, χ_{solvent} is the mole fraction of the solvent, and p_{pure} is the vapour pressure of the pure solvent.

$$\chi_{\text{solvent}} = \frac{n_{\text{solvent}}}{n_{\text{solute}} + n_{\text{solvent}}}$$

$$\chi_{A,\text{vapour}} = \frac{\chi_{A,\text{liquid}} \cdot p_{A,\text{pure}}}{\chi_{A,\text{liquid}} \cdot p_{A,\text{pure}} + \chi_{B,\text{liquid}} \cdot p_{B,\text{pure}}}$$

10.7 Distillation

The vapour pressure as well as the boiling point of the mixture will be intermediate between the two pure liquids.

Fractional distillation is a continuous re-distillation.

Higher BP (Boiling Point) vapour condenses and vapourizes over and over as it rises; the lower BP liquid drips back into the boiling mixture. Vapour becomes richer in the component with the lower boiling point. The final distillate is nearly pure (lower BP liquid), and the liquid in the pot is also nearly pure (higher BP liquid).

10.8 Pressure and Gas Solubility: Henry's Law

The solubility of a gas is directly proportional to its partial pressure: $s = k_H p$, where k_H is called Henry's constant.

10.9 Enthalpy of Solution

Measuring heat released or absorbed when a substance dissolves is called molar enthalpy of solution, ΔH_{sol} .

The first step sublimes solid ions to gas ions. Highly endothermic, this is the lattice enthalpy, ΔH_L ;

In the second step, gaseous ions plunge into water forming the final solution. This is the enthalpy of hydration, ΔH_{hyd} ;

Combining these steps: $\Delta H_{\text{sol}} = \Delta H_L + \Delta H_{\text{hyd}}$.

High charge and small ionic radius contribute to high lattice enthalpy (ΔH_L). However, often these can be the same properties that relate to low enthalpy of hydration (ΔH_{hyd}). Therefore it is very difficult to make reliable predictions and instead rationalise what is observed.

10.10 Solubility

Limits: unsaturated and saturated solution.

Like Dissolves Like: polar-polar, nonpolar-nonpolar.

10.11 Colligative Properties

Properties that depend on the numbers of solute and solvent molecules and not on chemical identity are called colligative properties.

Four colligative properties of major importance are the:

1. lowering of the vapour pressure,
2. raising boiling points,
3. lowering of freezing points, and
4. osmosis.

Colligative properties are measured using either mole fraction ($\chi_A = \frac{n_A}{n_A + n_B + \dots}$)

or molality ($\frac{n_{solute}(mol)}{m_{solvent}(kg)}$).

10.12 Boiling-Point Elevation and Freezing-Point Depression

A nonvolatile solute lowers the vapour pressure of the solvent, therefore increasing the boiling point and therefore it is called boiling-point elevation. The increased boiling temperature is usually quite small and is of little practical importance in science.

Freezing-point depression is more significant. An added solute lowers of the freezing point of a solvent. Melting point are also a method for determining the purity of a solid.

Boiling-point elevation = $k_b \times \text{molality}$, temperature increases.

Freezing-point depression = $k_f \times \text{molality}$, temperature decreases.

10.13 Van't Hoff i Factor

The van't Hoff i factor, is determined experimentally.

In very dilute solution, where all ions are independent, i = number of ions.

10.14 Osmosis

Osmotic pressure, $\Pi = iRTc$, where c is molarity.

Osmometry is the technique used to determine the molar mass of a solute from osmotic pressure measurements.

11 Chemical Equilibrium

11.1 Reactions at Equilibrium

The criteria for dynamic chemical equilibrium are:

1. The forward and reverse reactions are both taking place.
2. The forward rate equals the reverse rate.

Note, it's impossible to make more product when at equilibrium. The reaction just appears to have stopped moving.

11.2 Equilibrium Constant K

K is the same regardless of initial compositions.

The law of mass action

$$K = \left(\frac{\text{partial pressure of Products}}{\text{partial pressure of Reactants}} \right)_{\text{equilibrium}}$$

In general for $a A(g) + b B(g) \rightleftharpoons c C(g) + d D(g)$

$$K_p = \frac{(p_C)^c (p_D)^d}{(p_A)^a (p_B)^b}$$

here p means partial pressure, since the reactants and products are gases.

It can be shown empirically or thermodynamically that pure liquids or solids do not appear in K .

Aqueous solutions

In general for $a A(aq) + b B(aq) \rightleftharpoons c C(aq) + d D(aq)$

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

Note the change to brackets, $[]$, signifying, c , or molarity, $\text{mol} \cdot \text{L}^{-1}$.

Activity

It is common when deriving equations to simplify expressions without units.

Summary	Simplified form	Interpretation
Ideal gas	$a_J = \frac{p_J}{p^0}$	$[p^0 = 1 \text{ bar or } 1 \text{ atm}]$
Solute in a dilute solution	$a_J = \frac{[J]}{c^0}$	$[c^0 = 1 \text{ mol} \cdot \text{L}^{-1}]$
Pure solid or pure liquid	$a = 1$	unchanging throughout the reaction

11.3 Origins of K and ΔG

$\Delta G_r = \Delta G_r^\circ + RT \ln \frac{(a_C)^c (a_D)^d}{(a_A)^a (a_B)^b}$, here a = activity, either a gas or solute in a solution (molarity); r = the overall reaction.

Reaction quotient $Q = \frac{(a_C)^c (a_D)^d}{(a_A)^a (a_B)^b}$, then $\Delta G_r = \Delta G_r^\circ + RT \ln Q$.

Here K is a constant, it is known. Q is unknown and must be found. Q can be larger or smaller than K ; again it must be calculated.

Once $Q = K$ the reaction is at equilibrium, therefore $\Delta G = 0$. This leads to $\Delta G^\circ = -RT \ln K$. This of course links the thermodynamic tables ΔG° to K .

11.4 What Does K Mean?

1. $K \ll 1$ favours reactants at equilibrium (ΔG° is very positive);
2. When $10^{-3} < K < 10^3$ neither products nor reactants are favored; when $(-10 < \Delta G^\circ < 10)$ then temperature will be a factor;
3. $K \gg 1$ favours products at equilibrium (ΔG° is very negative).

11.5 The 5% rule

The 5% rule says we can ignore " x " when there is less than 5% decomposition or ionisation change (the definition of a weak acid).

A note of good practice is, once solving for x , always plug x back into the original equilibrium table to make sure you're below 5% and to get your final values.

11.6 Le Chatelier's Principle

Concentration

Le Chatelier's principle suggests a good way to ensure that a reaction goes on generating a substance: simply remove products as they are formed.

Inert Gas

Adding an inert gas does not interfere with the reacting gases, so the reacting gases continue to occupy the same volume, and so their individual molar concentrations and partial pressures remain unchanged despite the presence of an inert gas.

Temperature

Since ΔH and ΔS are independent of range of temperatures,

$$\ln \frac{K_2}{K_1} = \frac{\Delta H^\circ}{R} \left(\frac{1}{T_1} - \frac{1}{T_2} \right), \text{ here we see } K_2 \text{ over } K_1 \text{ and no "-" on the right.}$$

12 Acid-Base

12.1 Definitions

Svante Arrhenius

An acid is a compound that contains hydrogen and reacts with water to form hydrogen (H^+) ions.

A base is a compound that produces hydroxide ions (OH^-) in water.

Brønsted–Lowry

An acid is a proton donor, and a base is a proton acceptor.

A strong acid is fully deprotonated in solution (\rightarrow).

A weak acid is only partly deprotonated in solution (\rightleftharpoons).

A strong base is completely protonated in solution (\rightarrow).

A weak base is only partially protonated in solution (\rightleftharpoons).

Acid $\xrightarrow{\text{donates } H^+}$ conjugate base. Base $\xrightarrow{\text{accepts } H^+}$ conjugate acid.

Brønsted vs Arrhenius

Brønsted acids and bases are more general than the Arrhenius definitions.

Lewis

An acid is an electron pair acceptor.

A base is an electron pair donor.

Lewis vs Brønsted vs Arrhenius

Lewis theory is more general than Brønsted or Arrhenius acid–base theory.

12.2 Hydronium Ion

The H_3O^+ ion is called the hydronium ion. We say that the H_2O becomes strongly hydrated in a solution to form H_3O^+ .

12.3 Amphiprotic vs Amphoteric

Amphiprotic: a substance that can be a proton donor or receiver. (H_2O)

Amphoteric: a substance that can behave as either an acid or a base. (Al)

12.4 Autoprotolysis Constant

For $H_2O(l) + H_2O(l) \rightleftharpoons H_3O^+(aq) + OH^-(aq)$, $K_w = [H_3O^+][OH^-]$.

In pure water at 25°C , $K_w = [H_3O^+][OH^-] = 1.0 \times 10^{-14}$. The concentrations of H_3O^+ and OH^- are very low, which explains why pure water is such a poor conductor of electricity.

Autoprotolysis reaction is endothermic, so K_w increases with temperature.

12.5 The pH Scale

$$pH = -\log[H_3O^+], pOH = -\log[OH^-], pK_w = -\log K_w. pH + pOH = pK_w.$$

Most solutions used in chemistry have a pH ranging from 0 to 14, but values outside this range are possible.

12.6 Acid Strength

There is no general theory that describes acid strength. We only have general trends to work with and compare. 1. Equilibrium constants, K_a and K_b ; 2. Solvent.

Two different trends are observed, one for periods, the other for groups.

Period: Acid strength is based on bond polarity;

Group: Acid strength is based on bond strength. ($HF < HCl < HBr < HI$)

There are two ways to compare oxoacids.

In groups with the same number of oxygen atoms:

The greater the electronegativity the greater the acidity; ($HClO > HBrO > HIO$)

A family of the same element with different numbers of oxygen atoms:

The greater the number of oxygen atoms, the stronger the acid.

12.7 The pH of Salt Solutions

Common Cations in Water		
Character	Interpretation	Examples
Acidic	Conjugate acids of weak bases	NH_4^+
	Small, highly charged metal cations	$Fe^{3+}, Al^{3+}, Cu^{2+}$
Neutral	Group 1 and 2 cations	$Li^+, Na^+, K^+, Mg^{2+}, Ca^{2+}$
	metal cations with charge +1	Ag^+
Basic	None	

Common Anions in Water		
Character	Interpretation	Examples
Acidic	Very few	$HSO_4^-, H_2PO_4^-$
Neutral	Conjugate bases of strong acids	$Cl^-, Br^-, I^-, NO_3^-, ClO_4^-$
Basic	Conjugate bases of weak acids	$F^-, O^{2-}, OH^-, S^{2-}, HS^-, CN^-, CO_3^{2-}, PO_4^{3-}, NO_2^-, CH_3CO_2^-$, other carboxylate ions

12.8 The pH of Polyprotic Acids

For all polyprotic acids or bases, K_{a_1} is always the greatest contributor to the overall pH of the solution — except for H_2SO_4 . Only H_2SO_4 has a large K_{a_1} and K_{a_2} .

$pH = \frac{1}{2}(pK_{a_1} + pK_{a_2})$ as long as initial concentration $[X]_{initial} \gg K_{a_1}$.

12.9 Very Dilute Solutions

of Strong Acids and Bases

$$K_w = [H_3O^+]([H_3O^+] - [HCl]_{initial}).$$

$$K_w = [H_3O^+]([H_3O^+] - [NaOH]_{initial}).$$

of Weak Acids

Four unknowns exist for a weak acid HA , A^- , H_3O^+ , and OH^- .

Both $K_w = [H_3O^+][OH^-]$ and $K_a = \frac{[H_3O^+][A^-]}{[HA]}$ are used to arrive at the following cubic equation. $[H_3O^+]^3 + K_a[H_3O^+]^2 - (K_w + K_a[HA]_{initial})[H_3O^+] - K_aK_w = 0$.

13 Aqueous Solution Equilibrium

Dzsyang

14 Electrochemistry

14.1 Redox (Reduction Oxidation)

Redox reaction is combination of two separate reactions called half-reactions.

Electrons removed: called an oxidation reaction, reducing agent.

Electron taken: called a reduction reaction, oxidising agent.

Redox couple:

Couples are written as reductions, so Zn^{2+}/Zn implies $Zn^{2+}(aq) + 2e^- \rightarrow Zn(s)$.

14.2 Galvanic Cells

Galvanic cells are spontaneous reactions generating electric current. Batteries are a collection of galvanic cells joined in series, where the total voltage is the sum of each cell. A battery prevents direct contact of the anode and cathode reactions.

The **anode** is where oxidation takes place. The "-" charge is because electrons are being generated in the oxidation reaction.

The **cathode** is where reduction occurs. The "+" potential means electrons are being attracted (pulled) here.

Daniell cells contain a porous pot, which is a barrier through which ions travel. The porous pot allows ion exchange, thereby electrically neutralising the anode and cathode.

14.3 Cell Potential

A cell's electrical force is measured by its cell potential, E .

E is the ability of a cell to force electrons through a circuit. If both the anode and cathode have a lot of pushing-and-pulling power, this will generate a high cell potential. Exhausted batteries at equilibrium have no pushing or pulling power, and the cell potential is zero.

Electrical work is neither a change in pressure nor volume. Electrical work is nonexpansion work. Gibbs free energy is the measure of the maximum nonexpansion work that a reaction can do at constant pressure and temperature: $\Delta G = W_e$.

The total electrical work done is $W_e = -neN_A E$.

Faraday's constant, $F = eN_A = 96485 C \cdot mol^{-1}$.

With Faraday's constant: $W_e = -nFE$ or $\Delta G = -nFE$.

E is also taken to represent the electromotive force, **emf**, of a cell.

Overpotential

In practice, the applied potential must be substantially greater than the cell emf. The additional potential is called the overpotential.

14.4 The Notation for Cells: IUPAC, for a Daniell Cell

Anode electrode | anode electrolyte || cathode electrolyte | cathode electrode.

"|" means phase change like an electrode and solution.

"||" means a salt bridge.

The anode is also called the anodic compartment, and cathodic compartment is used for the cathode.

The salt bridge allows ions to move back and forth.

Voltmeters measure voltage

$(-)(+)$, $E > 0$, so $G < 0$ and this is a spontaneous cell.

$(+)(-)$, $E < 0$, so $G > 0$ and this is a nonspontaneous cell.

14.5 Standard Potentials

$$E^\circ = E^\circ(\text{cathode}) - E^\circ(\text{anode}).$$

The standard potential, E° , is only measuring a difference.

Standard Hydrogen Electrode (SHE) $2H^+(aq) + 2e^- \rightarrow H_2(g)$, $E^\circ = 0.0V$.

14.6 Standard Potentials and Equilibrium Constants

The standard reaction Gibbs free energy is related to the standard emf

$$\Delta G^\circ = -nFE^\circ = -RT\ln K. \text{ Therefore, } \ln K = \frac{nFE^\circ}{RT} = \frac{nE^\circ}{0.025693V}.$$

Since $\Delta G_r = \Delta G_r^\circ + RT\ln Q$,

$$\text{The Nernst Equation } E = E^\circ - \frac{0.025693V}{n} \ln Q.$$

14.7 Electrolytic Cells

Electrolytic cells are nonspontaneous cells, where a current is forced in the opposite direction to drive the reaction.

This process is called electrolysis, and is the opposite of a Galvanic cell.

Where as electrons flow spontaneously in a Galvanic cell from the anode to cathode, electrolytic cells apply an emf at the cathode, making the cathode "-" and the anode "+".

Reduction still takes place at the cathode, but electrons are being pumped in from an external power source.

Oxidation still takes place at the anode, but electrons are "flowing" upstream, like salmon during a spawning-run.

Notice the lack of a salt bridge.

15 Kinetics

15.1 Thermodynamics vs Kinetics

Thermodynamics tells us which way a chemical reaction will go.

Kinetics tells us how fast a chemical reaction will go and how to control the rate.

Dzsyang