

Chemistry Honors Study Guide

Semester 2 Final Exam

Test date: the calendar is glitching help

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1 Limiting and Excess Reagents

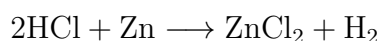
1.1 Definitions

- **limiting (lim) reagent:** The reactant that limits the amount of product yielded to a certain number.
- **excess (xs) reagent:** The reactant left over after all of the limiting reagent is used.
- **theoretical yield:** The theoretical amount of product produced found through calculations.
- **actual yield:** The amount of product actually produced, likely affected by sources of error, found by doing the experiment. (Also *experimental yield*.)
- **percent yield:** The percentage of the theoretical yield that is actually produced through experimentation.

1.2 Finding Limiting Reagent and Theoretical Yield

A limiting reagent limits the amount of product that can be created in a chemical reaction (no matter how much of the other reactant you have, the limiting reagent controls the maximum reactant produced, known as theoretical yield).

Example:



If you react 17 g HCl with 17 g Zn, what is the limiting reagent and theoretical yield of ZnCl₂?

Step 1: Find how many grams of ZnCl₂ each reactant provides.

$$(17 \text{ g HCl}) \times \left(\frac{1 \text{ mol HCl}}{36.461 \text{ g HCl}} \right) \times \left(\frac{1 \text{ mol ZnCl}_2}{2 \text{ mol HCl}} \right) \times \left(\frac{136.286 \text{ g ZnCl}_2}{1 \text{ mol ZnCl}_2} \right) \\ \approx 31.77 \text{ g ZnCl}_2$$

$$(17 \text{ g Zn}) \times \left(\frac{1 \text{ mol Zn}}{65.38 \text{ g Zn}} \right) \times \left(\frac{1 \text{ mol ZnCl}_2}{1 \text{ mol Zn}} \right) \times \left(\frac{136.286 \text{ g ZnCl}_2}{1 \text{ mol ZnCl}_2} \right) \\ \approx 35.44 \text{ g ZnCl}_2$$

Step 2: Find the limiting reagent by taking the element that produces the smallest amount of the product (the theoretical yield). The excess reagent is the other reactant.

limiting reagent: HCl
excess reagent: Zn
theoretical yield: 31.77 g ZnCl₂

1.3 Finding Percent Yield

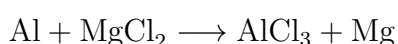
To find the percent yield, use this formula:

$$\frac{\text{actual yield}}{\text{theoretical yield}} \times 100\%$$

1.4 Finding Amount of Excess Reagent

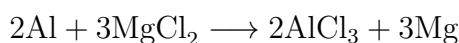
To find the amount of excess reagent left over after a reaction has occurred, work backwards from the theoretical yield.

Example:



If you react 10 g Al with 10 g MgCl₂, how many grams of excess reagent are there?

Step 1: Always check if the equation is balanced, and if not, balance it.



Step 2: Find the theoretical yield by applying the steps shown in the prior example.

$$(10 \text{ g Al}) \times \left(\frac{1 \text{ mol Al}}{26.982 \text{ g Al}} \right) \times \left(\frac{2 \text{ mol AlCl}_3}{2 \text{ mol Al}} \right) \times \left(\frac{133.341 \text{ g AlCl}_3}{1 \text{ mol AlCl}_3} \right) \\ \approx 49.42 \text{ g AlCl}_3$$

$$(10 \text{ g MgCl}_2) \times \left(\frac{1 \text{ mol MgCl}_2}{95.211 \text{ g MgCl}_2} \right) \times \left(\frac{2 \text{ mol AlCl}_3}{3 \text{ mol MgCl}_2} \right) \times \left(\frac{133.341 \text{ g AlCl}_3}{1 \text{ mol AlCl}_3} \right) \\ \approx 9.34 \text{ g AlCl}_3 \leftarrow \text{theoretical yield, lim. reagent is MgCl}_2$$

Step 3: Work backwards from the theoretical yield and find how many grams of excess reagent (Al) is needed to produce approximately 9.3 grams of AlCl₃.

$$(9.3 \text{ g AlCl}_3) \times \left(\frac{1 \text{ mol AlCl}_3}{133.341 \text{ g AlCl}_3} \right) \times \left(\frac{2 \text{ mol Al}}{2 \text{ mol AlCl}_3} \right) \times \left(\frac{26.982 \text{ g Al}}{1 \text{ mol Al}} \right) \\ \approx 1.88 \text{ g Al}$$

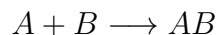
Step 4: Subtract this number from the amount of aluminum used in the reaction to get $10 - 1.88 = \boxed{8.12 \text{ g}}$ of excess reagent.

Step 5: Check if the answer is reasonable. There was a large difference between the product yielded by the magnesium chloride and the aluminum, so it makes sense that there is a significant amount of aluminum left over.

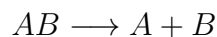
2 Ionic Reactions

2.1 Types of Reactions

Combination reaction Two or more reactants combine to create one product



Decomposition reaction Reactant decomposes into two or more products



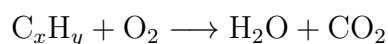
Double displacement reaction Two ionic compounds switch anions



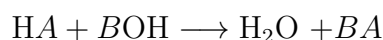
Single displacement reaction The anion switches from one reactant to the other (note: when predicting products, diatomic elements that stand alone should have 2 as a subscript)



Combustion reaction Substance reacts with oxygen, creating energy (heat)

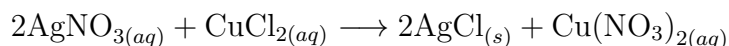


Neutralization reaction An acid reacts with a base to form water and an ionic compound

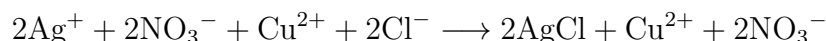


2.2 Ionic Equations

Balanced chemical eqn. (B.C.E.) or molecular equation



(Overall) ionic eqn. (I.E.): break apart aq. compounds (bring subscript to the front)



(do not need subscripts)

Compounds that can be cancelled out (**spectator ions**) are not involved in the reaction. ($2NO_3^-$ and Cu^{2+})

Net ionic eqn. (N.I.E.): get rid of compounds that cancel out (should be charge-balanced)



This is the actual reaction that occurs.

No net ionic equation, all spectator ions, all aqueous = NO REACTION

Precipitation rxn w/ ppt (precipitate, (s))



2.3 Solubility Rules

These rules are provided on tests and quizzes and do not need to be memorized. Soluble = aqueous, insoluble = solid.

1. Alkali metals (group 1) are always soluble.
2. Ammonium, (NH_4^+) is always soluble.
3. Nitrates (NO_3^-), chlorates (ClO_3^-), and perchlorates (ClO_4^-)
4. Most hydroxides (OH^-) are insoluble except those paired with an alkali metal or barium.
5. Most chlorides (Cl^-), bromides (Br^-), and iodides (I^-) are soluble except when they are paired with silver (Ag^+), lead (Pb^{2+}), and mercury (Hg^{2+}).
6. Carbonates (CO_3^{2-}), phosphates (PO_4^{3-}), and sulfides (S^{2-}) are insoluble except when paired with ammonium (NH_4^+) and alkali metals.
7. Most sulfates (SO_4^{2-}) are soluble except barium sulfate (BaSO_4), lead sulfate (PbSO_4), and mercury sulfate (HgSO_4).

3 Solutions

3.1 Definitions

- **solution:** A homogeneous mixture in the liquid phase.
- **solvent:** The larger part of the solution in the liquid phase.
- **solute:** The smaller part of the solution that is dissolved in the solvent. Can be solid, liquid, or gas.
- **percent by mass:** The percentage of the solution that is the solvent.
- **molarity:** How many moles of dissolved solute is present per liter of solution.
- **dilution:** To make a solution less concentrated.

3.2 Calculating Concentration

$$\text{Percent by mass} = \frac{\text{mass of solute}}{\text{mass of solution}} \times 100\%$$

$$\text{Molarity or } \frac{\text{mol}}{\text{L}} \text{ or } M = \frac{\text{moles of solute}}{\text{liters of solution}} \times 100\%$$

(units of molarity are $\frac{\text{mol}}{\text{L}}$ or M)

Example: If you dissolve **6.3 g NaCl** in **92 g H₂O**, the resulting solution has a volume of **90 ml**. What is the percent by mass of the solute and the molarity?

Step 1: Find the percent by mass using the formula above.

$$\frac{6.3 \text{ g NaCl}}{98.3 \text{ g H}_2\text{O}} \times 100\% = 6.41\%$$

Step 2: Find the number of moles of solute with the molar mass of NaCl.

$$(6.3 \text{ g}) \times \left(\frac{1 \text{ mol}}{58.44 \text{ g}} \right) = 0.11 \text{ mol NaCl}$$

Step 3: Calculate molarity using the formula above.

$$\frac{0.11 \text{ mol NaCl}}{0.09 \text{ L}} = 1.22 \text{ M}$$

percent by mass: 6.41%
molarity: 1.22 M

3.3 Dilution

$$C_1 V_1 = C_2 V_2$$

where:

C_1 is the concentration of the stock solution (more concentrated);

V_1 is its volume;

C_2 is the concentration of the diluted solution;

V_2 is its volume.

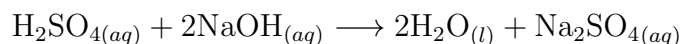
Example: How many ml of **12 M HCl** do you need to make **250 ml** of **0.5 M HCl**?

$$\begin{aligned} C_1 V_1 &= C_2 V_2 \\ (12 \text{ M}) V_1 &= (0.5 \text{ M})(250 \text{ ml}) \\ &= \boxed{10.4 \text{ ml}} \end{aligned}$$

3.4 Solution Stoichiometry

How many liters of **0.5 M H₂SO₄** are needed to fully react with **0.25 L** of **0.75 M NaOH**?

Step 1: Set up the balanced chemical equation with coefficients and subscripts, determining subscripts using solubility rules (page 6) and predicting products using types of reactions (page 5).



Step 2: Using molarity as a conversion factor, convert to moles, then perform the usual calculation.

$$(0.25 \text{ L NaOH}) \times \left(\frac{0.75 \text{ mol NaOH}}{1 \text{ L}} \right) \times \left(\frac{1 \text{ mol H}_2\text{SO}_4}{2 \text{ mol NaOH}} \right) \times \left(\frac{1 \text{ L H}_2\text{SO}_4}{0.5 \text{ mol H}_2\text{SO}_4} \right) \\ = \boxed{0.1875 \text{ L H}_2\text{SO}_4}$$

4 Acids and Bases

4.1 Definitions

- **Arrhenius acid:** Releases H^+ into the solution (starts with H).¹
- **Arrhenius base:** Releases OH^- into the solution (ends with OH).
- **Bronsted-Lowry acid:** A proton (H^+) donor.
- **Bronsted-Lowry base:** A proton acceptor. Does not necessarily have to end with OH.²
- **titration:** A laboratory technique, adding a substance with known concentration to a second substance with unknown concentration to determine the concentration or identity of the second.
- **titrant:** The substance in a titration with known concentration.
- **analyte:** The substance in a titration with unknown concentration.
- **monoprotic acid:** An acid that begins with 1 hydrogen atom.
- **diprotic acid:** An acid that begins with 2 hydrogen atoms.
- **monobase:** A base that ends with 1 hydroxide molecule.
- **dibase:** A base that ends with 2 hydroxide molecules.

4.2 pH and pOH

pH and pOH are negative log scales, so as they increase by 1 integer, the concentration of H^+ or OH^- decreases by tenfold. For example, a substance with pH 3 has 10 times the H^+ concentration as a substance with pH 4.

These are the three formulas:

Formula for pH:

$$\text{pH} = -\log[\text{H}^+] \quad (1)$$

Formula for pOH:

$$\text{pOH} = -\log[\text{OH}^-] \quad (2)$$

Relationship between pH and pOH:

$$\text{pH} + \text{pOH} = 14 \quad (3)$$

¹The exception to this is water; it is an **ampholyte** (**amphoteric**, able to act both as an acid and a base)

²A square is always a rectangle, but a rectangle is not always a square. Similarly, an Arrhenius base is always a Bronsted-Lowry base, but a Bronsted-Lowry base is not always an Arrhenius base.

And these are the formulas that are derived from the above:

Derived from 1:

$$[\text{H}^+] = 10^{-\text{pH}}$$

Derived from 2:

$$[\text{OH}^-] = 10^{-\text{pOH}}$$

Derived from 3:

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

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

The pH scale ranges from 0 to 14, with 0 being the most acidic, 14 being the most basic, and 7 being neutral (water).

The pOH scale also ranges from 0 to 14 but works in the opposite direction, with 0 being the most basic, 14 being the most acidic, and 7 being neutral.

4.3 Strong/Weak Acids and Bases

Strong acids/bases dissociate fully (use \longrightarrow for dissociation reaction) and weak acids/bases do not dissociate fully (use \rightleftharpoons for dissociation reaction).

The strong acids will be given on the test. Strong bases are **groups 1+2 OH^-** .

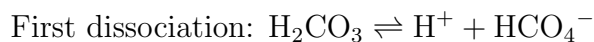
When an acid or base is strong, the molarity of H^+ or OH^- is equal to the molarity of the original compound. For example:



This is a strong acid, so $[\text{H}^+] = [\text{HCl}]$. For the dissociation of a weak acid/base, the concentrations are not equal.

4.4 Dissociation reactions

Diprotic dissociation:

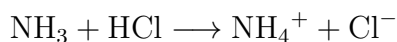


\rightleftharpoons denotes a reversible reaction and should be used when a weak acid/base is present.

Overall dissociation:



4.5 Conjugate Acids and Bases



acid, HCl, *donates a proton* (H^+) to become the **conjugate base**, Cl^-

base, NH_3 , *accepts a proton* to become the **conjugate acid** (NH_4^+)

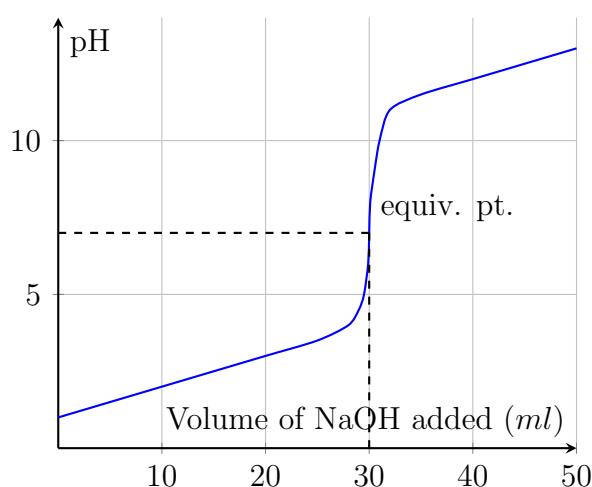
4.6 Naming Acids

Acids are named by their anions³ because they all begin with H.

Anion suffix	Acid suffix	Mnemonic device
-ide	hydro- -ic	My ride has <u>hy</u> draulics
-ate	-ic	I <u>a</u> te some <u>i</u> cky acid
-ite	-ous	Dynam <u>i</u> te is danger <u>o</u> us

4.7 Acid-Base Titrations

Acid-base titrations are the most common type of titration and are done by titrating an acid with a strong base or a base with a strong acid.



When performing a titration, try to aim for a light pink color. The **equivalence point** (where just enough of the titrant is added to completely neutralize the analyte) of a strong base/strong acid titration or a strong acid/strong base titration is at $\text{pH} = 7$. For dibases/diprotic acids, there are 2 equivalence points, and the pH at these points varies.

4.7.1 Titration Stoichiometry

You use 0.5 M NaOH to titrate a monoprotic acid. It takes 15 ml NaOH to reach the equivalence point, and you started with 25 ml of acid. What is the acid concentration?

Step 1: Write and balance the chemical formula.



Step 2: Calculate the number of moles of acid that you have.

$$\begin{aligned} (0.015\text{ L NaOH}) \times \left(\frac{0.5\text{ mol NaOH}}{1\text{ L NaOH}} \right) \times \left(\frac{1\text{ mol HCl}}{1\text{ mol NaOH}} \right) \\ = 0.0075\text{ mol HCl} \end{aligned}$$

³Keep prefixes (e. g. hypo-, per-)

Step 3: Put this number over the number of liters of acid you started with to find the concentration (molarity).

$$\frac{0.0075 \text{ mol HCl}}{0.025 \text{ L}} = \boxed{0.3 \text{ M HCl}}$$

5 Serial Dilution

A serial dilution is a series of multiple dilutions in a row.

5.1 Dilution Steps

1. Take 1 *ml* of the original solution and add it to a second container.
2. Add $x - 1$ *ml* of water.
3. Repeat as many times as necessary. The molarity should have decreased by x times (e.g. if you are performing a 1 : 4 dilution and the original molarity was 1.0 *M*, the molarity after diluting once is 0.25 *M*).

5.2 Beer's Law

$$A = \epsilon lc \quad (4)$$

where:

- A is the absorbance
- ϵ is the molar absorptivity (slope)
- l is the pathlength
- c is the concentration

6 States of Matter

6.1 Three States of Matter

State of Matter	Conforms?	Closely packed?	Compressible?
solid	no	yes	no
liquid	yes	yes	no
gas	yes	no	yes

Solids are not always more closely packed than liquids! There are exceptions, such as water.

6.2 Temperature

Kelvin K Fahrenheit °F Celsius °C

$$^{\circ}\text{F} = 1.8(^{\circ}\text{C}) + 32$$

$$^{\circ}\text{C} = \frac{^{\circ}\text{F} - 32}{1.8}$$

$$\text{K} = ^{\circ}\text{C} + 273$$

$$^{\circ}\text{C} = \text{K} - 273$$

$$\Delta T (\text{in}^{\circ}\text{F}) = 1.8 \cdot \Delta T (\text{in}^{\circ}\text{C or K})$$

6.3 Kinetic Energy

$$\text{KE} = \frac{1}{2}mv^2 \quad (5)$$

Example: Which gas would move 4x faster than O_2 ?

Solution: Using 5:

$$\frac{1}{2}(32 \text{ g/mol})(1 \text{ (arbitrary unit of velocity)}) = \frac{1}{2}(x \text{ g/mol})(4)$$

$$x = 2$$



7 Gasses

7.1 Definitions

- **gas pressure**: Collisions of molecules with the container.
- **diffusion**: The movement of particles from high concentration to low concentration until the concentration is consistent throughout.
- **effusion**: The movement of particles through a small hole in the container.
- **long-range order**: A regular, repetitive arrangement of particles. Solids with long-range order are called crystalline, and those without are called amorphous.
- **vapor pressure**: Measure of the tendency of a liquid to vaporise.

7.2 Pressure Units

760 mmHg (millimeters of mercury) = 1 atm (atmosphere)

7.3 The Gas Laws

where:

- P is the pressure
- V is the volume
- T is the temperature **in Kelvin**
- n is the number of moles
- R is the ideal gas constant equal to 0.0821

$$\text{Boyle's Law: } P_1V_1 = P_2V_2 \quad (6)$$

$$\text{Charles' Law: } \frac{V_1}{T_1} = \frac{V_2}{T_2} \quad (7)$$

$$\text{Gay-Lussac's Law: } \frac{P_1}{T_1} = \frac{P_2}{T_2} \quad (8)$$

$$\text{Combined Gas Law: } \frac{P_1V_1}{T_1} = \frac{P_2V_2}{T_2} \quad (9)$$

$$\text{Avogadro's Law: } \frac{V_1}{n_1} = \frac{V_2}{n_2} \quad (10)$$

$$\text{Ideal Gas Law: } PV = nRT \quad (11)$$

Important:

- $P \propto T$ and $V \propto T$. $P \propto \frac{1}{V}$. (P and V are directly proportional to T, and P is proportional to the reciprocal of V or inversely proportional to V.)
- Always use K for temperature.
- Laws 6, 7, and 8 are only true **if the temperature, pressure, and volume are held constant**.
- The ideal gas law uses specific units. The constant R , 0.0821, uses the units $\frac{L \cdot atm}{mol \cdot K}$, so volume must be in L , pressure in atm , and, as always, temperature in K.
- STP (standard temperature and pressure) is at $0^\circ C$ or 273K when 1 mol of any gas occupies 22.4 l .

Example: The density of an unknown gas at $20^\circ C$ and 749 $mmHg$ is 1.31 g/L . Calculate the molar mass of the gas.

Solution: Using 11, assuming 1 L of the gas, and converting to proper units:

$$\begin{aligned} (0.986 \text{ atm})(1 \text{ L}) &= \left(0.0821 \frac{L \cdot atm}{mol \cdot K}\right) (293K)(n) \\ n &= 0.041 \text{ mol} \\ 1.31 \text{ g} \div 0.041 \text{ mol} \\ &= \boxed{32.0 \text{ g/mol}} \end{aligned}$$

7.4 Diffusion and Effusion

7.4.1 Law of Effusion

$$\frac{\text{rate}_A}{\text{rate}_B} = \sqrt{\frac{MM_B}{MM_A}} \quad (12)$$

where MM denotes molar mass.

Example: How many times faster does He effuse than N_2 ?

Solution: Using 12:

$$\begin{aligned}\frac{\text{rate}_{\text{He}}}{\text{rate}_{\text{N}_2}} &= \sqrt{\frac{28.014}{4.003}} \\ &= \boxed{2.65}\end{aligned}$$

Example: Which gas diffuses 2.23 times faster than Br₂?

$$\begin{aligned}2.23 &= \sqrt{\frac{159.808 \text{ g/mol}}{x}} \\ x &\approx 32 \text{ g/mol} \\ &\boxed{\text{O}_2}\end{aligned}$$

7.4.2 Other Formulas

$$\textbf{Total pressure: } P_{\text{total}} = P_A + P_B + P_C + \dots \quad (13)$$

The total pressure of a mixture of gasses is equal to the sum of the pressures of its constituent parts.

$$\textbf{Definition of the mole fraction: } \chi_a = \frac{n_A}{n_{\text{total}}} \quad (14)$$

The mole fraction is defined as the number of moles of a substance over the moles of total gas.

$$\textbf{Dalton's Law of Partial Pressure: } P_a = \chi_a \cdot P_{\text{total}} \quad (15)$$

The pressure exerted by an individual gas is equal to the mole fraction times the total pressure.

Example: If you have 2.1 *mol* N₂, 0.8 *mol* CO₂, and 0.3 *mol* O₂ at 297K and occupying 1.8 *L*, what is P_{total} and the partial pressures?

Step 1: Add the number of moles together.

$$2.1 + 0.8 + 0.3 = 3.2 \text{ mol}$$

Step 2: Use $PV = nRT$ to find the total pressure.

$$\begin{aligned}PV &= nRT \\ (1.8 \text{ L})(P) &= \left(0.0821 \frac{\text{L} \cdot \text{atm}}{\text{mol} \cdot \text{K}}\right) (3.2 \text{ mol})(297\text{K}) \\ P &= 43.35 \text{ atm}\end{aligned}$$

Step 3: Find the mole fractions for each element with 14.

$$\chi_{\text{N}_2} = \frac{2.1 \text{ mol}}{3.2 \text{ mol}} = 0.656$$

$$\chi_{\text{CO}_2} = \frac{0.8 \text{ mol}}{3.2 \text{ mol}} = 0.25$$

$$\chi_{\text{O}_2} = \frac{0.3 \text{ mol}}{3.2 \text{ mol}} = 0.09$$

Step 4: Solve for partial pressure with 15.

$$P_{\text{N}_2} = 0.656 \cdot 43.35 = 28.44 \text{ atm}$$

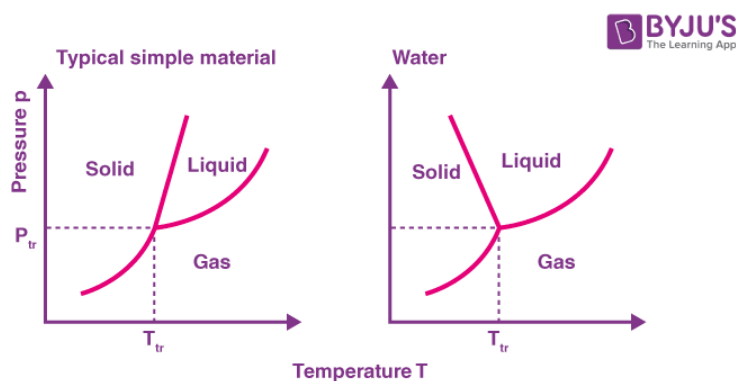
$$P_{\text{CO}_2} = 0.25 \cdot 43.35 = 10.84 \text{ atm}$$

$$P_{\text{O}_2} = 0.09 \cdot 43.35 = 3.90 \text{ atm}$$

$$P_{\text{N}_2} = 28.44 \text{ atm}; P_{\text{CO}_2} = 10.84 \text{ atm}; P_{\text{O}_2} = 3.90 \text{ atm}.$$

7.4.3 Phase Diagrams

Below are the phase diagrams for typical material and water:⁴



Note that in water, increasing the pressure leads to a phase change, since ice is less dense than water. The center point where all three lines intersect is called the **triple point**.

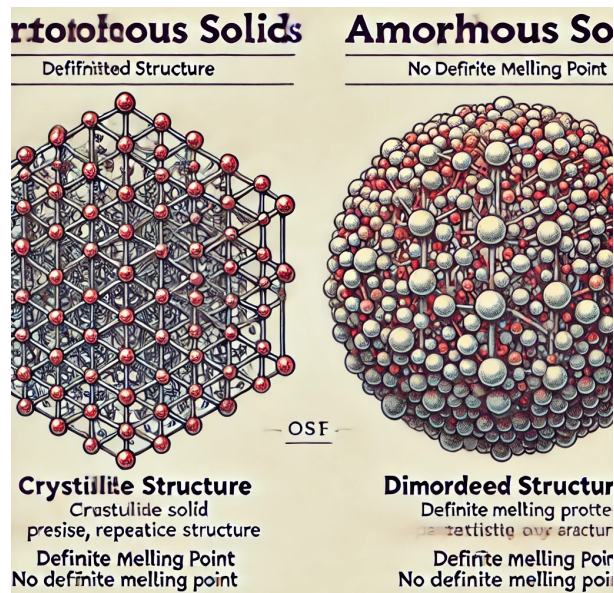
7.4.4 Melting Points of Solids

Distinct melting point (H_2O , metals...) v. **MP range** (fats, glass, plastic, rubber...)

Solids with distinct melting points are **crystalline** solids with long-range order. Solids with a melting point range are **amorphous** solids without long range order that melt at multiple temperatures.⁵

⁴<https://byjus.com/chemistry/phase-diagram-of-water/>

⁵<https://elibraryportal.com/amorphous-and-crystalline-solids/>



This is somewhat irrelevant, but here is an amusing AI-generated picture that attempts to explain this concept. (They tried.)

7.4.5 Vapor Pressure

As temperature increases, vapor pressure will also increase until vapor pressure is equal to the atmospheric pressure (i.e. the boiling point). There is an exchange of molecules happening at the surface of a liquid, and if you seal the container, it will eventually reach equilibrium.

7.4.6 Solutions

Freezing point depression happens when a non-volatile solute is added to a solvent that disrupts the lattice, interfering with the solvent's ability to form its regular crystalline structure.

Boiling point elevation happens when a non-volatile solute is added to a solvent and blocks the solvent from evaporating.

8 Thermodynamics

8.1 Definitions

- **heat**: The kinetic energy of particles.
- **temperature**: The measurement of heat.
- **endothermic process**: A process that absorbs heat from surroundings, going from a lower energy state to a higher energy state.
- **exothermic process**: A process that releases heat into surroundings, going from a higher energy state to a lower energy state.
- **calorie**: The amount of energy it takes to raise 1 gram of water by 1 degree Celsius (specific heat capacity of water).

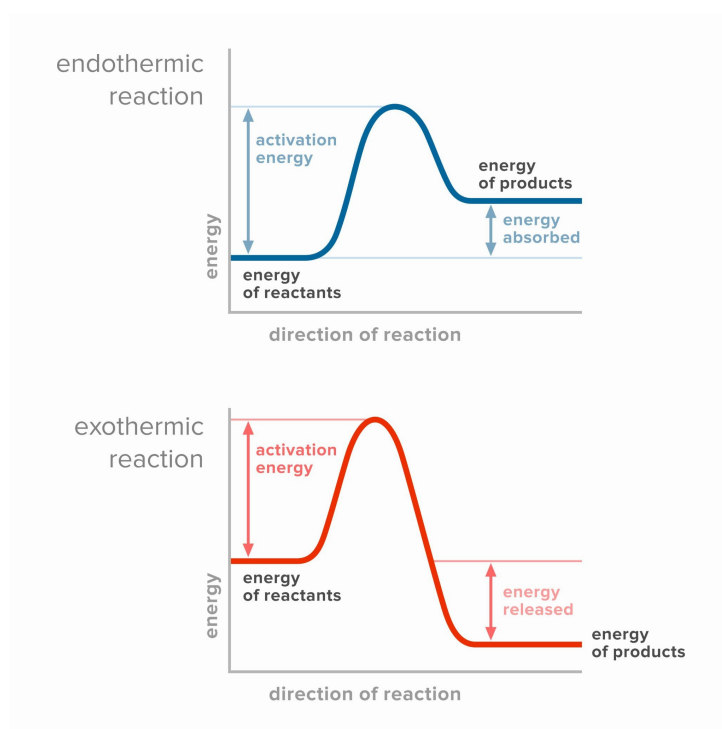
- **specific heat capacity:** The amount of energy it takes to raise the temperature of 1 gram of a substance by 1 degree Celsius. (For water, the specific heat capacity is 1 calorie.)
- **bond energy:** The heat required to break a bond or released when the bond is made.

8.2 Laws of Thermodynamics

1. Energy cannot be created nor destroyed.
2. Heat disperses until equal.

8.3 Reaction Diagrams for Chemical Changes

Below are the reaction diagrams for endothermic and exothermic reactions:⁶



Activation energy, or E_a , is the energy required to undergo a specific reaction. Energy released/absorbed is denoted by ΔH .

8.4 Heat Transfer

$$q = mc\Delta T \quad (16)$$

where:

- q is heat in joules
- m is mass in grams
- c is the specific heat capacity in joules over degrees Celsius
- ΔT is the change in temperature in degrees Celsius or Kelvin

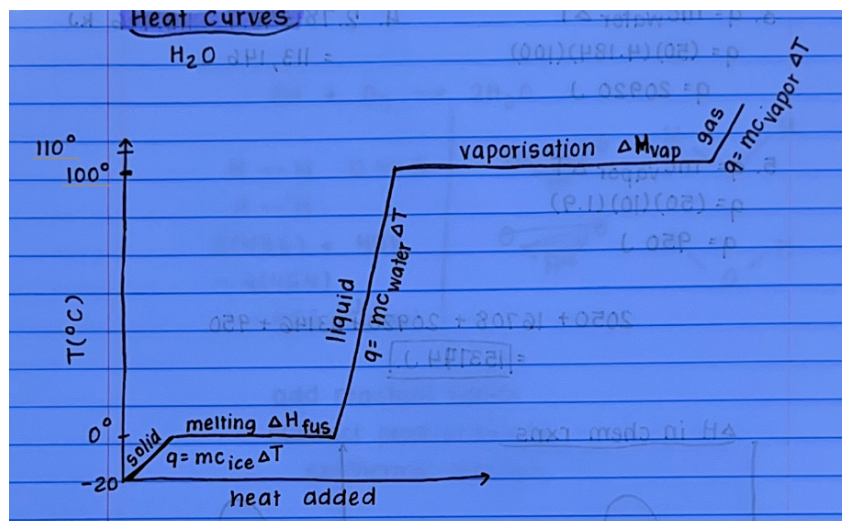
⁶<https://online-learning-college.com/knowledge-hub/gcses/gcse-chemistry-help/energy-level-diagrams/>

8.4.1 Heat Units

1 kilocalorie/food calorie (kcal or Cal) = 1000 calories (cal) = 4184 joules (J)

8.5 Heat Curves

A heat curve for water:



ΔH_{fusion} : amt of heat/mol that is absorbed when melting or released when freezing

ΔH_{vapor} : amt of heat/mol absorbed when vaporising/released when condensing

Example: How much heat does it take to change 50 g of water from a solid state at -20°C to a vapor state at 110°C ?

Given:

- melting point of water: 0°C
- boiling point of water: 100°C
- $c_{\text{ice}} = 2.05 \frac{\text{J}}{\text{g}^\circ\text{C}}$
- $c_{\text{water}} = 4.814 \frac{\text{J}}{\text{g}^\circ\text{C}}$
- $c_{\text{vapor}} = 1.9 \frac{\text{J}}{\text{g}^\circ\text{C}}$
- $\Delta H_{\text{fus}} = 6.01 \frac{\text{kJ}}{\text{mol}}$
- $\Delta H_{\text{vap}} = 40.7 \frac{\text{kJ}}{\text{mol}}$

This solution takes five steps. At the end, add them all together.

Step 1: Solid state. The melting point of water is 0°C and the ΔT is 20. Use $q = mc_{\text{ice}}\Delta T$:

$$q = (50)(2.05)(20) = 2050 \text{ J}$$

Step 2: Fusion (melting). By a simple calculation, we can find that there are 2.78 mol of water. Convert to kilojoules, then to joules:

$$6.01 \frac{\text{kJ}}{\text{mol}} \cdot 2.78 \text{ mol} = 16.708 \text{ kJ} = 16708 \text{ J}$$

Step 3: Liquid state. The boiling point of water is 100 °C and the ΔT is 100. Use $q = mc_{\text{water}}\Delta T$:

$$q = (50)(4.184)(100) = 20920 \text{ J}$$

Step 2: Vaporisation (boiling). Using the same 2.78 mol of water, convert ΔH_{vap} to kilojoules, then to joules:

$$40.7 \frac{\text{kJ}}{\text{mol}} \cdot 2.78 \text{ mol} = 113.146 \text{ kJ} = 113146 \text{ J}$$

Step 5: Vapor state. The final temperature is 110 °C and the ΔT is 10. Use $q = mc_{\text{vapor}}\Delta T$:

$$q = (50)(1.9)(10) = 950 \text{ J}$$

Step 6: Add each change in heat together.

$$\begin{aligned} 2050 + 16708 + 20920 + 113146 + 950 \\ = \boxed{153774 \text{ J}} \end{aligned}$$

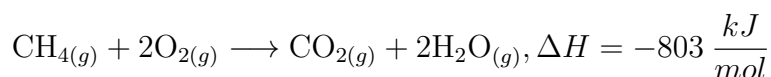
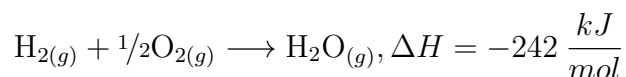
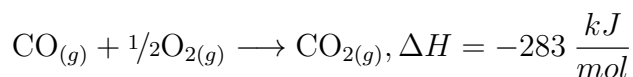
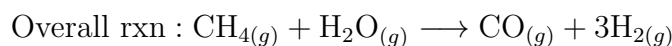
8.6 ΔH in Reactions

Breaking bonds **absorbs** heat, and making bonds **releases** heat. A positive ΔH suggests an endothermic reaction, and a negative ΔH suggests an exothermic reaction. To find the ΔH from a chemical equation, subtract the bond energies of the products from those of the reactants.

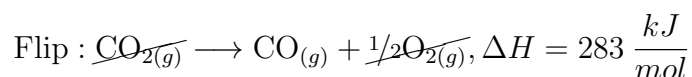
8.7 Hess' Law

When adding reactions, you can also add ΔH values.

Example: Find the ΔH of the overall reaction.



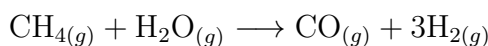
Manipulate the three equations so that all the unnecessary elements cancel out:



Flip and multiply by 3 : $\cancel{3\text{H}_2\text{O}_{(g)}} + \text{H}_2\text{O}_{(g)} \longrightarrow 3\text{H}_{2(g)} + \cancel{3/2\text{O}_{2(g)}}, \Delta H = 726 \frac{\text{kJ}}{\text{mol}}$

Leave as is : $\text{CH}_{4(g)} + \cancel{2\text{O}_{2(g)}} \longrightarrow \cancel{\text{CO}_{2(g)}} + \cancel{2\text{H}_2\text{O}_{(g)}}, \Delta H = -803 \frac{\text{kJ}}{\text{mol}}$

Leaving us with the overall reaction:

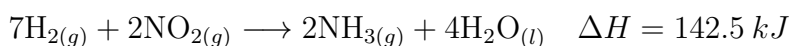


Adding the ΔH values:

$$283 + 726 - 803 = \boxed{206 \frac{\text{kJ}}{\text{mol}}}$$

9 Gasses and Heat in Stoichiometry

9.1 Heat of Reaction Stoichiometry



In the above equation, what would ΔH be if 4.8 mol of H_2 were reacted? Is this reaction endothermic or exothermic? Where does heat belong in this equation: with the reactants or products?

This reaction is **endothermic** because the ΔH is positive, meaning it absorbs heat. Heat belongs on the **reactants** side.



Step 1: Find the heat per mole of H_2 .

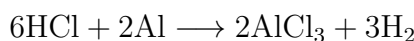
$$\frac{142.5 \text{ kJ}}{7 \text{ mol H}_2} = 20.36 \frac{\text{kJ}}{\text{mol}}$$

Step 2: Find the total heat by multiplying by 4.8 mol.

$$20.36 \frac{\text{kJ}}{\text{mol}} \times 4.8 \text{ mol} = \boxed{97.73 \text{ kJ}}$$

If 4.8 mol of H_2 were reacted, the ΔH would be 97.73 kJ.

9.2 Gasses



At STP, how many *ml* of H_2 gas are produced from 12 *g* of solid Al? (1 *mol* = 22.4 *L* at STP)

Using stoichiometry:

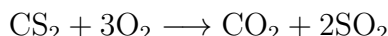
$$\begin{aligned} (12 \text{ g Al}) \times \left(\frac{1 \text{ mol Al}}{26.98 \text{ g Al}} \right) \times \left(\frac{3 \text{ mol H}_2}{2 \text{ mol Al}} \right) \times \left(\frac{22.4 \text{ L H}_2}{1 \text{ mol H}_2} \right) \times \left(\frac{1000 \text{ ml}}{1 \text{ L}} \right) \\ = \boxed{14944.4 \text{ ml}} \end{aligned}$$

9.3 Heats of Formation

ΔH_f is the heat absorbed/released when compounds are formed from elemental units. The ΔH_f of elements, including diatomic elements, is always 0.

Heats of formation equation:

$$\Delta H_{\text{rxn}} = \sum \Delta H_{\text{f(products)}} - \sum \Delta H_{\text{f(reactants)}} \quad (17)$$



Find the heat of formation given the following:

$$\begin{aligned} \Delta H_f(\text{CO}_2) &= -393.5 \frac{\text{kJ}}{\text{mol}} \\ \Delta H_f(\text{SO}_2) &= -296.8 \frac{\text{kJ}}{\text{mol}} \\ \Delta H_f(\text{CS}_2) &= 87.9 \frac{\text{kJ}}{\text{mol}} \end{aligned}$$

Solution: Using 17:

$$\begin{aligned} [-393.5 + 2(-296.8)] - [3(0) + 87.9] \\ = \boxed{1075 \frac{\text{kJ}}{\text{mol}}} \end{aligned}$$

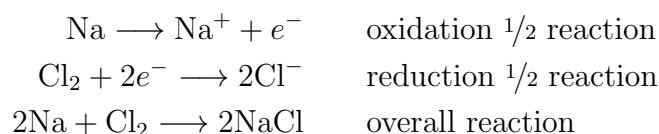
10 Redox

10.1 Definitions

- **redox reaction**: Short for *oxidation/reduction reaction*.
- **oxidation**: Losing electrons (becoming more positive).
- **reduction**: Gaining electrons (becoming more negative).
- **oxidizing agent**: Oxidizing another compound and is being reduced.
- **reducing agent**: Reducing another compound and is being oxidized.

10.2 Reactions

LEO the lion says GER (Losing Electrons = Oxidation, Gaining Electrons = Reduction)



10.3 Assigning Oxidation Numbers

10.3.1 Rules

1. An element in its elemental state is neutral.
2. H in a compound is always +1.
3. O in a compound is -2 except for H_2O_2 ; in that case, the oxidation number of O is -1.
4. Monatomic ions are whatever ionic charge they would normally form.
5. All oxidation numbers must add to the overall charge of the molecule or ion.

Example: What is the oxidation number of I in $\text{Mg}(\text{IO}_3)_2$?

Solution: By looking at the charge on Mg, the charge on $(\text{IO}_3)_2$ can be found to be 2-, so the charge on one ion is 1-, making the ion $(\text{IO}_3)^-$. The oxidation number of O in this compound is 2- by rule 3. There are 3 oxygen atoms, so:

$$3(-2) + \text{I} = -1$$

Solving for I:

$$-6 + \text{I} = -1$$

$$\text{I} = -1 + 6$$

$$\text{I} = 5$$

The charge on I is $\boxed{+5}$.

10.4 Writing Half-Reactions

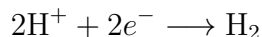


To write half-reactions, look at the oxidation numbers on each element. If the oxidation number has increased, it has been oxidised. If the number has decreased, it has been reduced.

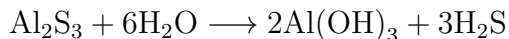
Oxidation $1/2$ reaction:



Reduction $1/2$ reaction:



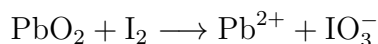
If *none* of the oxidation numbers change, the reaction is *not* a redox reaction. For example,



is not a redox reaction.

10.5 Balancing Redox Reactions

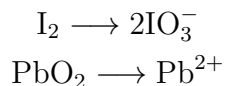
Assume an acidic, aqueous environment:



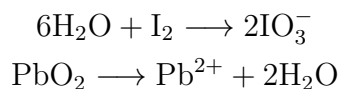
1. Separate into oxidation and reduction half-reactions by looking at oxidation numbers.



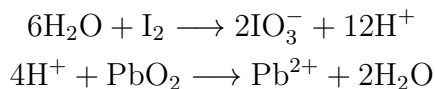
2. Balance elements that are *not* H or O.



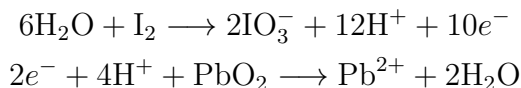
3. Balance O by adding H_2O .



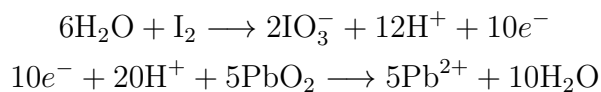
4. Balance H by adding H^+ .



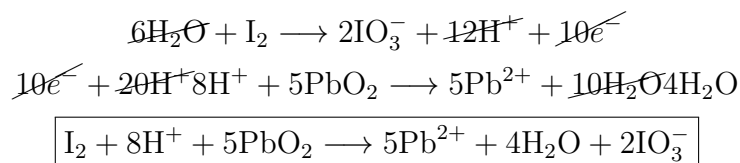
5. Add electrons to balance the charges. When balanced correctly, electrons should be on opposite sides (one on reactants, one on products).



6. Multiply to cancel out the electrons.

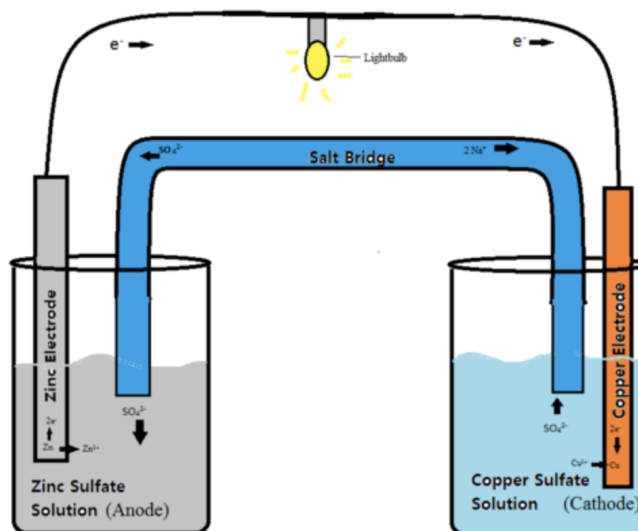


7. Combine and cancel.



11 Batteries

Below is a diagram of a Voltaic/Galvanic cell⁷:



containing:

electrodes to carry electricity. The **anode** is oxidised, losing mass as the material is released into solution, and the **cathode** is reduced, gaining mass as solid metal deposits onto it.

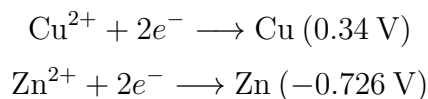
salt bridge to close the circuit.

solutions that the electrodes are suspended in.

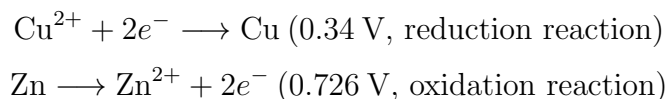
Electrons flow from the anode to the cathode.

To find the voltage of this battery:

⁷bit.ly/4jMzl6a



Flip the equation that is higher on the standard reduction potential table:



Combine:



N.B. When **multiplying** an equation to balance out the electrons, **the voltage does not change**. It only changes in sign if the equation is flipped.

12 Nuclear Chemistry

12.1 Types of Radioactive Decay

Radioactive decay, or **nuclear radiation**, is the release of particles/energy from unstable nuclei (radio isotopes).

There are three types:

α -radiation ${}^4_2\text{He}$ released, penetrates $\sim \frac{1}{20} \text{ mm}$ into body tissue, blocked by paper

β -radiation ${}^0_{-1}e$ released, penetrates $\sim 4 \text{ mm}$ into body tissue, blocked by aluminum foil

γ -radiation ${}^0_0\gamma$ released, penetrates several inches of lead and feet of concrete

12.2 Half-lives

The half-life of a radioisotope sample is the amount of time it takes for 50% of it to decay. Each sample has its own distinct half-life. The amount of sample left follows an exponential decay model that reaches 0 asymptotically.

The half-life formula is:

$$N = N_0 \left(\frac{1}{2} \right)^{\frac{t}{t_{1/2}}} \quad (18)$$

where:

- N = final amount of sample
- N_0 = initial amount of sample
- t = time elapsed
- $t_{1/2}$ = half-life

Example: What % of C-14 is left after 8,000 years if the half-life is 5,730 years?

$$N = N_0 \left(\frac{1}{2} \right)^{\frac{t}{t_{1/2}}}$$

$$N = 100 \left(\frac{1}{2} \right)^{\frac{8000}{5730}}$$

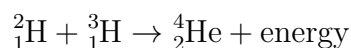
$$N = 38$$

38%

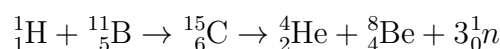
12.3 Nuclear Fusion and Fission

12.3.1 Nuclear Fusion

Combination of 2+ smaller nuclei into a larger nucleus. Happens at > 40 million $^{\circ}\text{C}$ (in the sun). Releases more energy than fission.



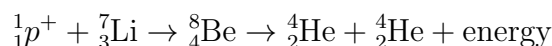
(can also be used to start a chain reaction)



12.3.2 Nuclear Fission

The nucleus is bombarded with neutrons, splitting into smaller nuclei.

Fission first happened in 1932 in the Cavendish Laboratory where Li-7 split into helium atoms:



proton and nucleus \rightarrow unstable isotope \rightarrow splits or decays

(triggers decay)

Modern fission (neutrons) in 1938



fissionable nucleus \rightarrow stable isotope \rightarrow fission

supercritical: more neutrons than in (such as in a bomb)

1 kilogram of U-235 is as destructive as 20,000 tons of dynamite!

fusion = :)
fission = :(