



# Grade 12 Chemistry Review Workbook

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# 1. Organic Chemistry

General IUPAC rules for naming organic compounds:

- 1) The lowest numbers possible must be used to denote positions.
- 2) To separate numbers, use commas.
- 3) To separate numbers and letters, use hyphens.
- 4) Names of branches and parent chains are not to be separated.
- 5) Unnecessary numbers should be omitted.

The **parent chain** is the longest carbon chain of an organic compound. The length of the parent chain is associated with a certain prefix, as shown below.

Prefix	Number of Carbons in Parent Chain	Prefix	Number of Carbons in Parent Chain
meth-	1	hex-	6
eth-	2	hept-	7
prop-	3	oct-	8
but-	4	non-	9
pent-	5	dec-	10

## 1.1 Alkanes, Alkenes, and Alkynes

A **hydrocarbon** is an organic compound that contains only carbon and hydrogen atoms.

Hydrocarbon Type	Carbon-Carbon Bonds	Saturation
Alkane	Only single bonds	Saturated
Alkene	At least one double bond	Unsaturated
Alkyne	At least one triple bond	Unsaturated

A **saturated hydrocarbon** is one with only single bonds, such that it is "saturated" with the maximum number of hydrogen atoms. In contrast, **unsaturated hydrocarbons** contain double or triple bonds, allowing for less hydrogen atoms to be bonded to each carbon atom.

### 1.1.1 Naming Branched Alkanes, Alkenes, and Alkynes

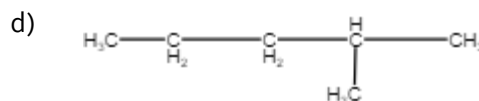
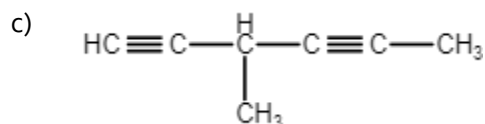
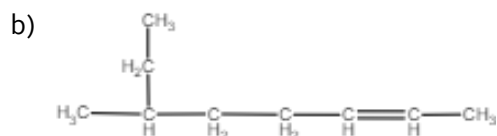
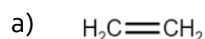
Straight chains become branched chains with the addition of **alkyl groups**, which are hydrocarbon groups derived from alkanes by the removal of an H atom, e.g.  $-\text{CH}_3$  = "methyl,"  $-\text{C}_2\text{H}_5$  = "ethyl." Use the steps below as a guide to name branched alkanes.

- 1) Identify the parent chain.
- 2) Number the carbon atoms, starting with the end closest to the multiple bond, if present. If no multiple bonds are present, start naming with the end closest to the branch(es).
- 3) Indicate the location of the multiple bond, if present, by the number of the carbon atom that begins the multiple bond. Indicate multiple double or triple bonds by the suffixes "di-" (2), "tri-" (3), "tetra-" (4), etc.
- 4) Name each branch and indicate its location on the parent chain by the number of the carbon atom at the point of attachment. If more than 1 branch is present, list the branches in alphabetical order. If there are multiple identical branches, insert a comma between their location numbers and use the prefixes "di-" (2), "tri-" (3), "tetra-" (4), etc. before the branch name.
- 5) Write the complete IUPAC name, including both multiple bonds and branches. Use the suffix "-ane" for alkanes, "-ene" for alkenes, and "-yne" for alkynes.



### 1.1 Example

Name the following organic compounds.



### 1.2 Halogenoalkanes

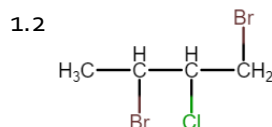
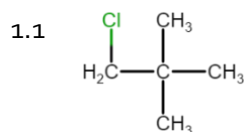
**Halogenoalkanes** (also called haloalkanes or alkyl halides) are formed when halogen atoms (fluorine, chlorine, bromine, or iodine) are added to a parent chain to replace a hydrogen atom. Halogenoalkanes have the general formula  $\text{C}_n\text{H}_{2n+1}\text{X}$ , where X denotes the halogen atom.

#### 1.2.1 Naming Halogenoalkanes

Halogenoalkanes are named by treating the halogens as branches. The halogen branch names are derived by removing the "-ine" from the end of the halogen name and replacing it with "-o," e.g. "fluoro-," "chloro-," "bromo-," "iodo-." Multiple identical halogen branches are indicated by the suffixes "di-," "tri-," "tetra-," etc.

### 1.2 Example

Name the following organic compounds.



### 1.3 Alcohols

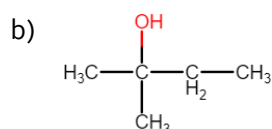
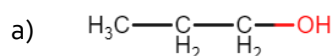
**Alcohols** are organic compounds which include a hydroxyl functional group (-OH) and have a general formula  $\text{C}_n\text{H}_{2n+1}\text{OH}$ . Alcohols with the same formula may differ by the location of the hydroxyl group.

#### 1.3.1 Naming Alcohols

To indicate the hydroxyl functional group, add the suffix "-ol" to the parent chain (remove the "-e" first). Numbering must start at the end nearest the hydroxyl group. The hydroxyl group position is indicated by the number associated with the carbon atom to which it is attached. This number can either be placed between the "-an" and the "-ol" (e.g. butan-2-ol) or before the "-anol" (e.g. 2-butanol).

### 1.3 Example

Name the following organic compounds.



## 1.4 Ethers

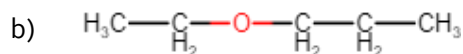
**Ethers** are organic compounds containing 2 alkyl groups attached to an oxygen atom.

### 1.4.1 Naming Ethers

“-oxy” is placed after the prefix of the smaller hydrocarbon group, and the alkane name of the larger hydrocarbon group is added at the end.

### 1.4 Example

Name the following organic compounds.



## 1.5 Aldehydes and Ketones

Both aldehydes and ketones contain a **carbonyl functional group** (C=O). Aldehydes and ketones differ by the position of their carbonyl groups, as shown below.

Organic Compound	Carbonyl Group Position	General Structure
Aldehyde	“Terminal”: bonded to a hydrogen atom	$\begin{array}{c} \text{O} \\    \\ \text{R}-\text{C}-\text{H} \end{array}$
Ketone	Bonded to 2 carbon atoms	$\begin{array}{c} \text{O} \\    \\ \text{R}-\text{C}-\text{R}' \end{array}$

### 1.5.1 Naming Aldehydes

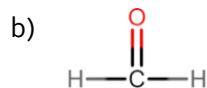
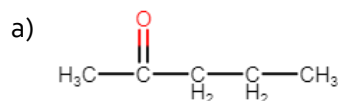
Drop the final “-e” from the parent alkane name and add the suffix “-al.” No numbering is necessary to indicate the position of the carbonyl group since it is always located at the end of the molecule.

### 1.5.2 Naming Ketones

Drop the final “-e” from the parent alkane name and add the suffix “-one.” If the carbon chain containing the carbonyl group has more than 4 carbon atoms, it is necessary to use a numerical prefix to specify the location of the carbonyl group (if less than 4 carbons, the number would be redundant).

### 1.5 Example

Name the following organic compounds.





## 1.6 Carboxylic Acids and Esters

Both **carboxylic acids** and **esters** contain a carbonyl functional group with an additional oxygen atom single-bonded to the carbon atom. Carboxylic acids and esters differ by the position of their carbonyl groups, as shown below.

Organic Compound	Carbonyl Group Position	General Structure
Carboxylic Acid	"Terminal": The added oxygen is bonded to a hydrogen atom	$\begin{array}{c} \text{O} \\ \parallel \\ \text{R}-\text{C}-\text{OH} \end{array}$
Ester	The added oxygen is bonded to a carbon atom	$\begin{array}{c} \text{O} \\ \parallel \\ \text{R}-\text{C}-\text{OR}' \end{array}$

### 1.6.1 Naming Carboxylic Acids

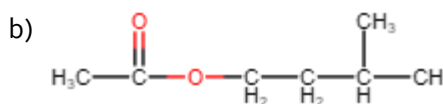
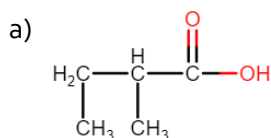
Replace the "-e" at the end of the parent chain containing the functional group with the suffix "-oic" followed by the word "acid."

### 1.6.2 Naming Esters

Write the name of the alkyl group attached to the singly bonded oxygen on the carbonyl group. Follow with the name of the chain bonded to the other side of the carbonyl group, changing the suffix to "-oate."

### 1.6 Example

Name the following organic compounds.



## 1.7 Amines

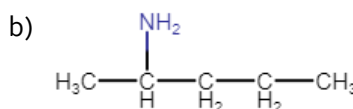
An **amine** is an ammonia molecule in which one or more of the hydrogen atoms are substituted by alkyl or aromatic groups. Labelling amines as **primary**, **secondary**, and **tertiary** differentiate those that have had 1, 2, or 3 of their hydrogen atoms substituted, respectively.

### 1.7.1 Naming Amines

Add the suffix "-amine" to the end of the name of the carbon chain to which the ammonia derivative is attached. Indicate the nitrogen atom's position with a numerical prefix if required (avoid redundancy).

### 1.7 Example

Name the following organic compounds.





## 1.8 Amides

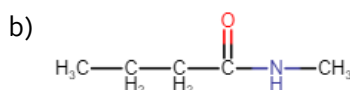
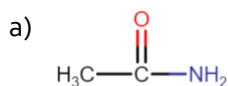
An **amide** is characterized by the presence of a carbonyl functional group (C=O) bonded to a nitrogen atom.

### 1.8.1 Naming Amides

Amide Type	Naming Procedure
Primary (N attached to one C)	Add the suffix "-amide."
Secondary (N attached to two C's)	Name alkyl groups attached to the nitrogen atom as substituents using "N."

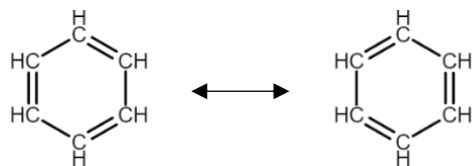
### 1.8 Example

Name the following organic compounds.

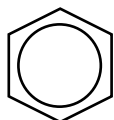


## 1.9 Aromatic Hydrocarbons

**Benzene**, C<sub>6</sub>H<sub>6</sub>, is a hydrocarbon in the form of a hexagonal ring of carbon atoms. Each of the 6 carbon atoms is bonded to 2 other carbon atoms and a hydrogen atom. Benzene has two resonance structures which differ in the location of the 3 carbon-carbon double bonds, shown below.



In reality, each of the carbon-carbon bonds are equivalent, somewhere between being single bonds and being double bonds. To represent this, benzene is sometimes represented using the structure below.

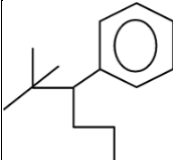


All **aromatic hydrocarbons** are derivatives of benzene (note that benzene itself is also considered an aromatic hydrocarbon).

### 1.9.1 Naming Aromatic Hydrocarbons

Complexity	Naming Convention	Example
Any number of substituents on benzene with 6 or less carbons	Treat benzene as a parent molecule with alkyl branches. For multiple substituents, number the carbons in the benzene ring as small as possible to specify alkyl positions	<p>ethylbenzene      1-methyl-4-propylbenzene</p>



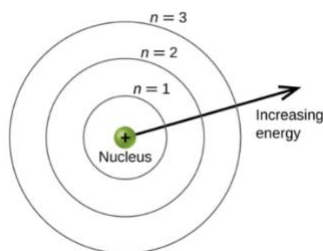
1 substituent on benzene with more than 6 carbons	Treat benzene as a branch on the larger parent chain, called a "phenyl" group	 2,2-dimethyl-3-phenylhexane
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### 1.10 Types of Chemical Reactions

Reaction Type	Explanation
Substitution	One or more atoms in a molecule are replaced with another atom or group of atoms
Addition	An atom or molecule is added to an unsaturated molecule (has at least one multiple bond) to make a single product
Elimination	Adjacent atoms are removed from a molecule to form a multiple bond and to release a small molecule
Condensation	Two molecules combine to form a single molecule and to produce another small molecule, often water.
Esterification	A condensation reaction where an alcohol and a carboxylic acid combine to form an ester and to produce a water molecule.
Hydrolysis	Water is added to a molecule to break the molecule apart (the reverse of a condensation reaction). The breakage typically occurs at a C-O-C linkage

## 2. Atomic Structure, The Periodic Table, Quantum Mechanics & Electron Configuration

### 2.1 Bohr Atomic Model



The Bohr atomic model represents the atom as a central nucleus containing protons and neutrons, surrounded by circular electron shells at set distances. The shells closest to the nucleus have less energy than further shells. Each shell is designated using a number and the letter "n," with increasing numbers at increasing distances from the nucleus (1, 2, then 3).

The **valence shell** of an atom is the outermost electron shell, and the electrons in the valence shell are called **valence electrons**. The number of valence electrons determines the atom's reactivity. A full valence shell is most stable/least reactive. The **octet rule** states that most elements require 8 electrons in their valence shell to be stable.

### 2.2 Electron Configurations and the Periodic Table

An element's **atomic number** is the number of protons each of its atoms has. The periodic table organizes elements into rows and columns with similar properties.



Generally, the number of valence electrons is the same for all elements in the same column, and going from left to right within a row, elements have an increasing number of valence electrons. As a result, an element's group number is indicative of its reactivity, as shown below.

Group	Valence Electrons	Reactivity
1 "Alkali Metals"	1	Elements are very unstable in single atomic form so are very reactive. They attain stability by losing their valence electron to form a cation (a positively charged ion).
2 "Alkali Earth Metals"	2	Elements are slightly more stable and less reactive than Group 1. They lose their valence electrons to form cations and attain stability.
14	4	Elements typically share electrons and form bonds with multiple other atoms to attain a full valence shell.
17	7	Elements become anions (negatively charged ions) by accepting 1 electron from another atom to achieve a full valence shell.
18 "Inert/Noble Gases"	8 (except He with 2)	Elements are very stable in single atomic form due to their full valence shell, making them non-reactive.

### 2.3 Blocks and Orbitals of the Periodic Table

Block	Block Description	Orbital Description
s	The s sublevel has 1 orbital which can hold a maximum of 2 electrons.	Spherical shape.
p	The p sublevel has 3 orbitals so can hold a maximum of 6 electrons.	Dumbbell shape.
d	The d sublevel has 5 orbitals so can hold a maximum of 10 electrons.	Various shapes, depending on orientation.
f	The f sublevel has 7 orbitals so can hold a maximum of 14 electrons.	Complex shape.

### 2.4 Electron Configurations

The **Heisenberg Uncertainty Principle** states that both the position and momentum of an electron cannot be determined at the same time. There are 3 rules and principles for completing electron configuration diagrams, as summarized below.

Rule or Principle	Description
Pauli Exclusion Principle	No 2 electrons in an atom can have the same 4 quantum numbers.
Aufbau Principle	Each electron is added to the lowest energy orbital available.
Hund's Rule	When there are several orbitals at the same energy level (p, d, f), electrons are placed singly in each orbital with the same spin before being paired up.

There are 2 methods which can be used to determine an atom's electron configuration, as shown below.

Method	Description
Aufbau Principle	Fill orbitals in the order dictated by the chart below (start at the top with 1s):





Periodic Table	<p>Fill orbitals in the order dictated by the blocks of the periodic table:</p>

### 2.4.1 Anomalous Electron Configurations

**Anomalous electron configurations** allow atoms to have lower overall energy and to be more stable. Half-filled and filled subshells are more stable than unfilled subshells (unfilled means neither half-filled nor filled). This is especially true for d orbitals, more so than s orbitals. 2 important anomalous configurations to remember are those of chromium and of copper, as shown below

Element	Expected Configuration	Actual Configuration
Chromium	$1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d_4$	$1s^2 2s^2 2p^6 3s^2 3p^6 4s^1 3d_5$
Copper	$1s^2 2s^2 2p^6 3s^2 3p^6 4s^2 3d_9$	$1s^2 2s^2 2p^6 3s^2 3p^6 4s^1 3d_{10}$

### 2.4.2 Electron Configurations of Ions

Ion Type	Procedure
Cation (positive)	Identify the orbital with the highest n which is furthest from the nucleus in the electron configuration of the neutral atom, and remove the necessary number of electrons from that orbital first (e.g. remove from 3p before 3s, and remove from 4s before 3d).
Anion (negative)	Identify the electron configuration of the neutral atom, and add the necessary number of electrons in the appropriate order according to the Aufbau principle.

### 2.4.3 Shorthand/Noble Gas Electron Configurations

Indicate an atom's core electrons (electrons occupying inner orbitals) using the preceding noble gas, and then write out the rest of the electron configuration for its valence electrons only. For example, the shorthand electron configuration of barium is  $[\text{Xe}]6s_2$ .



## 2.4 Example

- Write both the full and shorthand electron configurations of germanium (Ge).
- Write the full electron configuration for  $\text{Ge}_{4+}$ .

# 3. Bonding Theory

## 3.1 Lewis Structures

The following steps can be used as a guide to draw **Lewis Structures**:

- Determine the total number of valence electrons (make sure to account for any charges).
- Arrange all the atoms around a central atom, generally the least electronegative one. Connect each atom to the central atom with a single bond, represented by one electron pair or by one line.
- Distribute remaining electrons on the outer atoms as lone pairs to complete their octets.
- Arrange all remaining electrons on the central atom.

## 3.2 VSEPR Theory

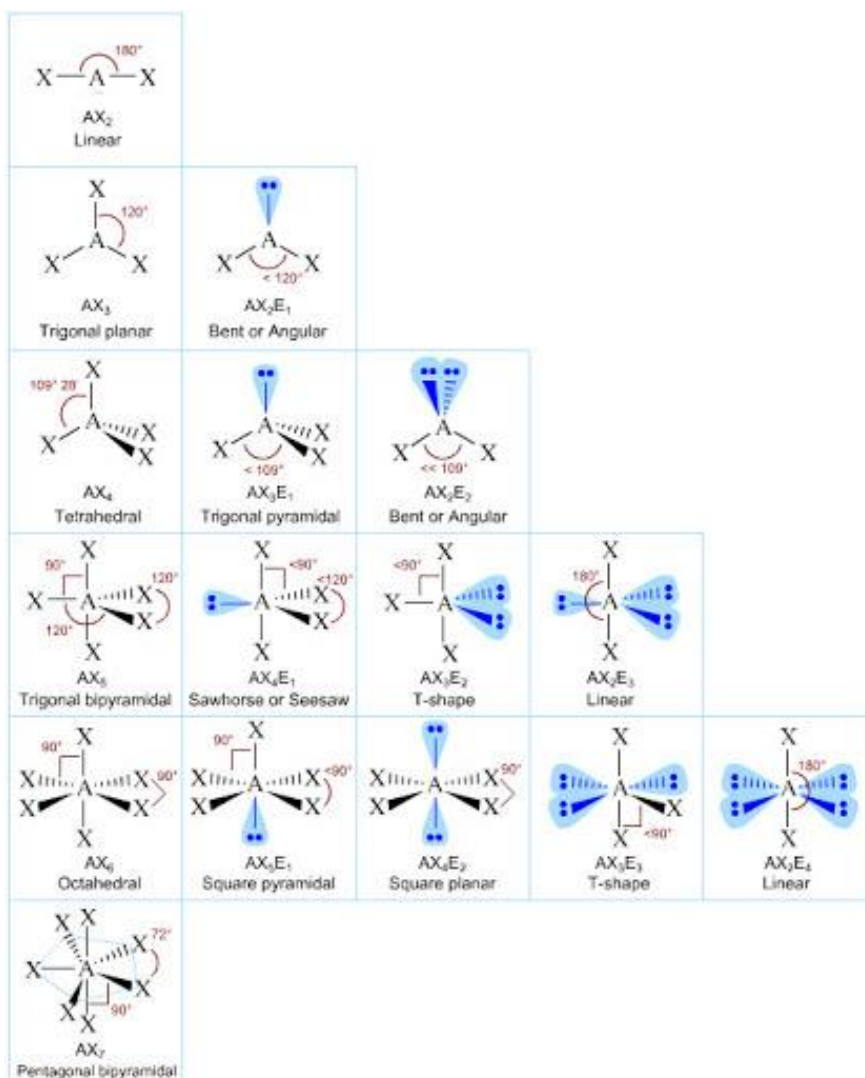
**Valence Shell Electron Pair Repulsion (VSEPR)** is a model used to predict the shapes and bond angles of simple molecules based on the theory that molecules will adopt a shape/arrangement which minimizes the repulsion between valence electrons by maximizing the distance between them. Lone pairs of electrons exert a greater repelling effect than bonding pairs do, such that lone pair-lone pair repulsion > lone pair-bonding pair repulsion > bonding pair-bonding pair repulsion.

In the figure below, the leftmost column lists all electron-pair geometries, each one representing molecules with a specific number of regions of electron density. All other columns to the right are the molecular structures associated to each electron-pair geometry, and they differ depending on the number of lone pairs in the molecule. The following steps can be used to determine the electron pair geometry and molecular structure of a molecule:

- Draw the Lewis Structure of the molecule.
- Count the number of regions of electron density (including both bonding and non-bonding pairs), considering all types of bonds (single, double, or triple) as a single region of electron density each.
- Determine the electron-pair geometry based on the number of regions of electron density.



4) Determine the molecular structure by considering the number of lone pairs.



### 3.3 Polarity

**Molecular polarity** is determined by considering a molecule's shape and bond polarities. The following steps can be used to determine molecular polarity:

- 1) Draw the Lewis Structure and determine the molecular shape. Draw the shape diagram.
- 2) Consider the electronegativity differences between atoms to determine bond polarity. If a bond is polar, illustrate by drawing an arrow pointing towards the more electronegative atom with a plus sign at the tail next to the bonds in the structural formula. Below is a reminder of how to determine bond polarity/type.

Electronegativity Difference, $\Delta EN$	Bond Type
$0.0 \leq 0.4$	Pure covalent
$0.4 < 1.7$	Polar covalent
$1.7 < 3.3$	Ionic



- 3) Decide whether the molecule as a whole is polar by considering the dipoles (the arrows indicating bond polarity) as vectors and taking the vector sum. If the sum is zero, the molecule is non-polar. If the sum is not zero, the molecule is polar, so draw a large arrow near the structural diagram to indicate the direction of net polarity.

### 3.3 Example

What is the electron pair geometry and molecular structure of  $\text{NI}_3$ ? Is  $\text{NI}_3$  polar or non-polar, and why?

## 4. Intermolecular Forces, Solids, and Physical Properties

### 4.1 Intermolecular Forces

**Intermolecular forces** are attractive forces between molecules and are considerably weaker than covalent bonds. **Dipole-dipole forces** and **London dispersion forces** are both classified as van der Waals forces. The table below explains van der Waals forces and Hydrogen bonds.

Intermolecular Force	Description	Strength
London dispersion forces	A temporary dipole in one atom can induce an opposite dipole in a neighbouring atom. The induced dipole results in a weak and short-lasting attractive force. London dispersion forces are always present but are most significant in larger atoms, in molecules with many atoms, and at low temperatures.	Weak
Dipole-dipole forces	Polar molecules attract each other electrostatically by lining up so that the positive and negative ends are close to each other. This causes dipole-dipole forces. These forces weaken as the distance between dipoles increases, making them relatively unimportant in gases.	Strong
Hydrogen bonds	Hydrogen bonds are unusually strong dipole-dipole forces in molecules where a hydrogen atom is covalently bonded to a highly electronegative atom, specifically nitrogen, oxygen, or fluorine.	Strongest

### 4.2 Solid Structures

Type of Solid	Structural Unit	Forces of Attraction	Properties
Ionic	Oppositely charged ions (cations and anions)	Electrostatic attractions	Hard, brittle, very high melting points, poor conductors of electricity (molten and aqueous conduct, but solids do not)
Covalent Network	Enormous network of atoms	Covalent bonds	Very hard, strong, very high melting points, often poor conductors of electricity



Molecular	Polar or non-polar molecules	Intermolecular forces	Can vary based on which intermolecular forces are present, fairly soft, poor conductors of electricity, low to moderate melting points
Metallic	Fixed positive metallic ions surrounded by a "sea" of delocalized electrons	Metallic bonds (non-directional bonds)	Soft to very hard, malleable, ductile, excellent conductors of electricity, low to very high melting points, sheen

### 4.3 Physical Properties

Physical Property	General Trend	Reason for Trend
Atomic radius	Decreasing left to right, increasing top to bottom	Increasing effective nuclear charge left to right, so electrons are more strongly attracted to and therefore more closely held to the nucleus. Increasing number of energy levels top to bottom, taking up more space.
1 <sup>st</sup> ionization energy	Increasing left to right, decreasing top to bottom	Increasing effective nuclear charge left to right, so valence electrons more are strongly attracted to nucleus and are therefore harder to remove. Top to bottom, valence electrons are further from the nucleus so experience weaker attraction and are easier to remove.
Electronegativity	Increasing left to right, decreasing top to bottom	Increasing effective nuclear charge left to right, so electrons are more strongly attracted to the nucleus. Top to bottom, valence electrons are further from the nucleus so experience weaker attraction.

## 5. Thermochemistry

### 5.1 Heat Transfer

The **heat transfer**,  $q$ , of a substance of mass  $m$  with specific heat capacity  $c$  over a change in temperature  $\Delta T = T_{\text{final}} - T_{\text{initial}}$  can be calculated using the equation  $q = mc\Delta T$ . A negative  $q$  signifies the release of energy in a reaction (exothermic) while a positive  $q$  signifies the absorption of energy (endothermic). Due to the 1<sup>st</sup> law of thermodynamics, all energy released from one entity is absorbed by the other, as represented by the equation  $q_A = -q_B$ .

#### 5.1 Example

Determine the final temperature (in Kelvin) when a 70.5 g piece of copper at 330 K is placed into 130 g of water at 305 K. The specific heat capacity of copper is 0.385 J/g°C and the specific heat capacity of water is 4.184 J/g°C.



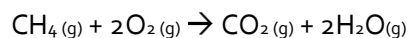
## 5.2 Enthalpy

The **enthalpy change**,  $\Delta H$ , of a system is calculated using the equation  $\Delta H = \Delta H_{\text{products}} - \Delta H_{\text{reactants}}$ . When the enthalpy change is positive, the reaction is **endothermic**. When the enthalpy change is negative, the reaction is **exothermic**.

The **standard enthalpy of formation**,  $\Delta H^{\circ}_f$ , is the enthalpy change that occurs when one mole of a compound is formed from its elements in their **standard states** (conditions of 1 atm of pressure and 25°C). Based on standard enthalpies of formation, the standard enthalpy of reaction,  $\Delta H^{\circ}_{\text{rxn}}$ , can be calculated using the equation  $\Delta H^{\circ}_{\text{rxn}} = \sum \Delta H^{\circ}_f (\text{products}) - \sum \Delta H^{\circ}_f (\text{reactants})$ .

### 5.2 Example

Calculate the standard enthalpy change for the following reaction given that the standard enthalpies of formation of methane, oxygen gas, carbon dioxide, and water are -75 kJ/mol, 0 kJ/mol, -394 kJ/mol, and -284 kJ/mol, respectively.



## 5.3 Hess's Law

**Hess's Law** states that, regardless of the multiple stages or steps of a reaction, the total enthalpy change for the reaction is the sum of all the enthalpy changes, such that  $\Delta H_{\text{overall}} = \sum \Delta H_{\text{steps}}$ . There are 2 rules to remember when applying Hess's Law:

- 1) If a chemical equation is reversed, the sign of  $\Delta H$  changes but the magnitude remains constant.
- 2) If the coefficients of a chemical equation are multiplied by a constant,  $\Delta H$  is multiplied by the same constant.

### 5.3 Example

What is the enthalpy change for the formation of 1 mole of butane ( $\text{C}_4\text{H}_{10}$ ) gas from its elements?

Target/Overall reaction:	$4\text{C}(\text{s}) + 5\text{H}_2(\text{g}) \rightarrow \text{C}_4\text{H}_{10}(\text{g})$	$\Delta H^{\circ} = ?$
Steps:	[1] $\text{C}_4\text{H}_{10}(\text{g}) + 13/2 \text{O}_2(\text{g}) \rightarrow 4\text{CO}_2(\text{g}) + 5\text{H}_2\text{O}(\text{g})$	$\Delta H^{\circ}_1 = -2657.4 \text{ kJ}$
	[2] $\text{C}(\text{s}) + \text{O}_2(\text{g}) \rightarrow \text{CO}_2(\text{g})$	$\Delta H^{\circ}_2 = -393.5 \text{ kJ}$
	[3] $2\text{H}_2(\text{g}) + \text{O}_2(\text{g}) \rightarrow 2\text{H}_2\text{O}(\text{g})$	$\Delta H^{\circ}_3 = -483.6 \text{ kJ}$



## 6. Chemical Kinetics

### 6.1 Collision Theory

**Collision theory** states that:

- 1) Reactions occur because of collisions of reactants with one another or the walls of their container.
- 2) Collisions are only effective if they have sufficient energy and correct alignment/orientation.
- 3) Reaction rate is dependent on collision frequency and fraction of effective collisions.

To increase reaction rate, several factors can be manipulated, as summarized below.

Factor	Effect on Reaction Rate
Concentration	A higher concentration of reactants means a greater number of particles per unit volume, and thus there is a greater chance of collision of the particles in the same fixed space. Thus, concentration of reactants affects the collision frequency.
Surface area (for heterogeneous systems only)	Collisions can only occur at the interface where 2 reactants in different phases are in contact. Increasing the surface area of a reactant allows for more "exposure" between the 2 phases. Thus, surface area affects the collision frequency.
Temperature	At higher temperatures, the average kinetic energy of reacting molecules is higher. Therefore, the molecules are moving around more quickly and collide more often. Also, at higher temperatures, since molecules are moving more quickly, collisions are more forceful, so a greater fraction of reacting molecules possess the minimum threshold (activation energy) when they collide. Thus, temperature affects the collision frequency and the fraction of effective collisions.
Catalysts	A catalyst works by providing an alternative pathway from reactants to products. The activation energy of the catalyzed reaction is lower than for the uncatalyzed reaction. So, at a particular temperature, more collisions meet the minimum threshold energy requirement for a successful collision, simply because the energy barrier has been lowered. Thus, catalysts affect the fraction of effective collisions.

### 6.2 Rate Laws and Order of Reaction

Reaction rate is proportional to the product of the initial concentration of reactants raised to some exponential values. This is represented by the rate law equation  $r = k[X]^m[Y]^n$ , where  $r$  is the rate of reaction in units of concentration over time,  $k$  is the rate constant (units vary),  $[X]$  and  $[Y]$  are the concentrations of reactants  $X$  and  $Y$ , respectively, and  $m$  and  $n$  represent the orders of reaction for reactants  $X$  and  $Y$ , respectively. The sum of the individual orders of reaction ( $m + n$ ) represents the **overall order of reaction**. The rate law for a specific reaction can be experimentally determined by conducting a series of experiments at the same temperature where only the initial concentration of one reactant varies at a time.

#### 6.2 Example

Chlorine dioxide reacts with hydroxide ions to produce a mixture of chlorate and chlorite ions:



The data in the table below were determined at a constant temperature. Find the rate law equation, the overall order of reaction, and the rate constant.



Experiment	Initial [ClO <sub>2</sub> ] (mol/L)	Initial [OH <sup>-</sup> ] (mol/L)	Initial Rate of Formation of Products (mol/L·s)
1	0.0150	0.0250	$1.30 \times 10^{-3}$
2	0.0150	0.0500	$2.60 \times 10^{-3}$
3	0.0450	0.0250	$1.16 \times 10^{-2}$

## 7. Equilibrium

### 7.1 Equilibrium Conditions

**Chemical equilibrium** is achieved when the forward and reverse reactions of a reversible reaction occur at equal rates. For true chemical equilibrium to exist, the system must be closed and at constant temperature and pressure.

### 7.2 Le Châtelier's Principle

**Le Châtelier's Principle** states that when a system at equilibrium is exposed to stress, the equilibrium will shift to relieve the stress. The 3 stresses that can be applied to a system at equilibrium and the shifts they cause are summarized below.

Stress	Effect
Changes in temperature	Endothermic reactions are favoured by an increase in temperature. Exothermic reactions are favoured by a decrease in temperature.
Changes in pressure or volume (for gas phase reactions only)	Reactions producing fewer moles of gaseous particles are favoured by an increase in pressure or decrease in volume. Reactions producing more moles of gaseous particles are favoured by a decrease in pressure or increase in volume.
Changes in concentration	Reactions producing substances whose concentrations have been decreased are favoured. Reactions consuming substances whose concentrations have been increased are favoured.

### 7.3 Equilibrium Constants

A general equation for a reversible reaction can be written as  $aA + bB \rightleftharpoons cC + dD$ . The **reaction quotient**,  $Q$ , for this equation is  $Q = \frac{[C]^c[D]^d}{[A]^a[B]^b}$ . At equilibrium, the reaction quotient always has the same value and is called the **equilibrium constant**,  $K$ , of the reaction. The table below summarizes the relationship between the values of  $Q$  and  $K$ .

Q vs. K	Significance
$Q < K$	The system is not at equilibrium and must shift right/toward the products to reach it.
$Q = K$	The system is at equilibrium.
$Q > K$	The system is not at equilibrium and must shift left/toward the reactants to reach it.





### 7.3 Example

Consider the reaction  $2\text{HI}(\text{g}) \rightleftharpoons \text{H}_2(\text{g}) + \text{I}_2(\text{g})$  with  $K = 0.020$ . If  $[\text{HI}(\text{g})] = 0.20 \text{ mol/L}$ ,  $[\text{H}_2(\text{g})] = 0.15 \text{ mol/L}$ , and  $[\text{I}_2(\text{g})] = 0.09 \text{ mol/L}$ , is the system at equilibrium? If not, in which direction must it shift to reach equilibrium?

### 7.4 Calculating Equilibrium Concentrations

An **ICE Table** is a convenient way to solve equilibrium problems. To create an ICE table, write the reaction equation on the top row and create columns beneath each species in the reaction. Label the rows of the first column I, C, and E (symbolizing initial, change, and equilibrium). To use the ICE table, plug in all known concentrations and work through the problem to find all necessary information.

#### 7.4 Example

$K$  for the reaction  $\text{I}_2(\text{g}) + \text{Br}_2(\text{g}) \rightleftharpoons 2\text{IBr}(\text{g})$  is found to be  $1.20 \times 10^2$ . Starting with 4.00 mol each of iodine and bromine in a 2.00 L flask, calculate the equilibrium concentrations of all reaction components.

## 8. Acids and Bases

### 8.1 Brønsted-Lowry Acids and Bases

A **Brønsted-Lowry acid** is a proton donor and a **Brønsted-Lowry base** is a proton acceptor. The **conjugate base** of an acid is the base that remains after the acid donates a proton. The **conjugate acid** of a base is the acid that is formed after the base accepts a proton. Every acid-base reaction involves the transfer of a proton between the acid and the base.

### 8.2 Autoionization of Water

Water can act as a proton donor or proton acceptor depending on the circumstances, meaning it is **amphoteric** (also called **amphiprotic**). Water is also capable of acting as a proton donor and proton acceptor toward itself. The spontaneous reaction between 2 water molecules producing a hydronium ion and a hydroxide ion is called the **autoionization of water** and is represented by the equation  $2\text{H}_2\text{O}(\text{l}) \rightleftharpoons \text{H}_3\text{O}^+(\text{aq}) + \text{OH}^-(\text{aq})$ . Since this is an equilibrium system, it has an equilibrium constant called the **ion product constant for water**,  $K_w = [\text{H}_3\text{O}^+][\text{OH}^-] = 1.0 \times 10^{-14}$  (at  $25^\circ\text{C}$ ).



### 8.3 pH and pOH

The **pH** of a solution is calculated using the equation  $\text{pH} = -\log[\text{H}_3\text{O}^+]$ . If the pH is known, the hydronium ion concentration of a solution can be calculated using the equation  $[\text{H}_3\text{O}^+] = 10^{-\text{pH}}$ . Similarly, the **pOH** of a solution is calculated using the equation  $\text{pOH} = -\log[\text{OH}^-]$  and the hydroxide ion concentration can be calculated using the equation  $[\text{OH}^-] = 10^{-\text{pOH}}$ . At 25°C, it is always the case that the sum of a solution's pH and pOH is 14.00.

At 25°C, a solution is **acidic** if it has a pH of less than 7.00, which corresponds to a hydronium ion concentration of more than  $1.0 \times 10^{-7}$ . In contrast, a solution is **basic** at 25°C if it has a pH of more than 7.00, which corresponds to a hydronium ion concentration of less than  $1.0 \times 10^{-7}$ .

#### 8.3 Example

What is the hydronium ion concentration in a 0.075 M solution of NaOH? What are the pH and pOH and is the solution acidic, basic, or neutral? Assume a temperature of 25°C. Note that NaOH is a strong base and completely dissociates.

### 8.4 Acid and Base Strength

**Strong acids** and **strong bases** ionize completely in aqueous solution. The most common strong acids include solutions of HI, HBr, HCl,  $\text{HClO}_3$ ,  $\text{HClO}_4$ ,  $\text{H}_2\text{SO}_4$ , and  $\text{HNO}_3$ . The most common strong bases include solutions of Group I and Group II hydroxides, including NaOH, KOH, RbOH, CsOH,  $\text{Ca}(\text{OH})_2$ ,  $\text{Sr}(\text{OH})_2$ , and  $\text{Ba}(\text{OH})_2$ .

### 8.5 Neutralization Reactions

In a **neutralization reaction**, an acid and a base react to form water and a salt. The neutralization reactions of a strong acid with a strong base, a strong acid with a weak base, and a weak acid with a strong base, result in pH values equal to 7, less than 7, and more than 7, respectively.

### 8.6 Titrations

When a titration reaches its **equivalence point**, the solutions have been mixed in exactly the right proportions according to the equation. In contrast, when a titration reaches its **endpoint**, the indicator changes colour. The equivalence point and endpoint are not necessarily exactly the same. From a **titration curve** which illustrates the change in pH which results from adding a certain volume of titrant, the equivalence point can be determined.

## 9. Electrochemistry

### 9.1 Oxidation-Reduction (REDOX) Reactions

A **redox reaction** is a reaction which involves the movement of electrons between the reactions, with one reactant being the electron donor and the other the electron acceptor. A redox reaction involves 2



sub-steps: **reduction**, where electrons are gained, and **oxidation**, where electrons are lost. Helpful mnemonics include LEO the lion says GER (LEO = lost electrons → oxidation; GER = gained electrons → reduction), and OIL RIG (OIL = oxidation is loss; RIG = reduction is gain). The **reducing agent** is the reactant which undergoes oxidation itself and facilitates reduction of the other reactant. The **oxidizing agent** is the reactant which undergoes reduction itself and facilitates oxidation of the other reactant.

## 9.2 Oxidation Numbers

The partial movement of electrons within covalent bonds can be shown more easily by using **oxidation numbers**. Oxidation numbers are assigned to each atom within a molecule and their magnitude indicates the number of electrons an atom may lose or gain control of in polar covalent bonding. Use the steps below as a guide to assign oxidation numbers.

- 1) Assign all atoms in elemental form an oxidation number of zero.
- 2) Assign monatomic cations and anions oxidation numbers equal to the charge on the ions.
- 3) Assign a hydrogen atom an oxidation number of +1 (except in a hydride, where it is combined with a metal, it should be -1).
- 4) Assign an oxygen atom an oxidation number of -2 (except in hydrogen peroxide,  $\text{H}_2\text{O}_2$ , where it is -1, and in  $\text{OF}_2$ , where it is +2).
- 5) Assign group 1 elements an oxidation number of +1, group 2 elements an oxidation number of +2, and aluminum an oxidation number of +3.
- 6) Assign a halogen atom an oxidation number of -1.
- 7) Deduce the oxidation numbers of all other elements by considering that the sum of all oxidation numbers within a neutral molecule must be zero and that the sum of all oxidation numbers within a polyatomic ion must equal the charge on the ion.

A positive oxidation number signifies a loss of control of electrons, a negative oxidation number signifies a gain of control of electrons, and an oxidation number of zero signifies equal electron sharing. Whenever the oxidation number of at least one atom changes, a redox reaction is occurring. When the oxidation number of an atom reduces or increases, it is undergoing reduction or oxidation, respectively.

### 9.2 Example

For each of the following chemical reactions, assign oxidation numbers to all atoms/ions, indicate if the reaction represents a redox reaction and why, and if it does, identify the oxidizing and reducing agents.



## 9.3 Balancing Redox Reactions Using the Half Reaction Method

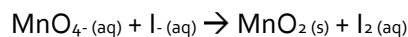
Redox reactions can be represented using **half reactions** which show oxidation and reduction independently. The **half reaction method** of balancing redox reactions considers each half reaction separately, balancing each to ensure that the number of electrons lost in the oxidation half reaction equals the number of electrons gained in the reduction half reaction. For a complete redox reaction, the half reactions must be combined after balancing each individually. No electrons can appear in the overall redox reaction. Use the steps below as a guide to balance using the half reaction method.



- 1) Write the 2 unbalanced half reactions showing reactant/products.
- 2) Balance atoms (excluding O & H).
- 3) Balance oxygen by adding  $\text{H}_2\text{O}(\text{l})$ .
- 4) Balance hydrogen by adding  $\text{H}^+(\text{aq})$ .
- 5) For reactions occurring in **basic** solutions, complete these additional steps:
  - i. Add as many  $\text{OH}^-(\text{aq})$  to both sides as the number of  $\text{H}^+(\text{aq})$ .
  - ii. Simplify by combining  $\text{H}^+(\text{aq})$  and  $\text{OH}^-(\text{aq})$  to form  $\text{H}_2\text{O}(\text{l})$ .
  - iii. Simplify  $\text{H}_2\text{O}(\text{l})$  molecules.
- 6) Balance charge by adding electrons.
- 7) If necessary, multiply 1 or both of the half reactions by the lowest possible number so that they both contain equal numbers of electrons.
- 8) Add the balance half reactions together.
- 9) Simplify by removing any species which appear on both sides of the overall reaction.
- 10) Verify correctness by checking that the numbers of each atom on each side are equal and that the total charges are balanced.

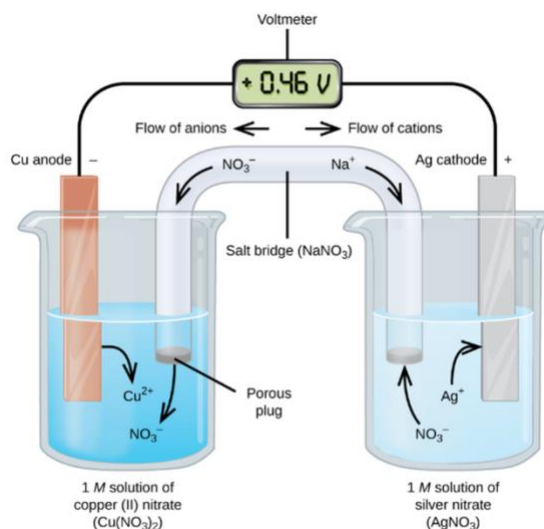
### 9.3 Example

Balance the skeleton redox equation below using the half reaction method. Assume it occurs in a basic medium.



### 9.4 Galvanic Cells

A **galvanic cell**, also known as an **electrochemical cell**, is a cell where a spontaneous redox reaction is used to produce an electric current. Each beaker containing the solution in contact with its metal is called a half cell, since a half reaction of the redox reaction occurs therein. The **anode** is the electrode where oxidation occurs and carries a negative charge. The **cathode** is the electrode where reduction occurs and carries a positive charge. Electrons flow from the anode to the cathode. The **salt bridge** is necessary in the cell to complete the electrical circuit without introducing any more metal into the system and to maintain the electrical neutrality of each half cell. Galvanic cells are usually drawn in the style shown below.

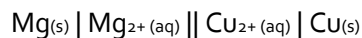


**Cell notation** allows for a short-hand way to represent a galvanic cell. A single vertical line, |, represents a phase boundary between the solid electrode and the aqueous solution. A double vertical line, ||, represents the salt bridge. Cell notation is written in the order shown below:

anode | cation of anode || cation of cathode | cathode

#### 9.4 Example

For the galvanic cell indicated by the cell notation below, identify the anode, cathode, oxidizing agent, reducing agent, reduction and oxidation half reactions, overall cell reaction.



#### References:

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