ACJC CHEMISTRY DEPARTMENT

JC2 H2 Revision 2024

Summary of definitions for H2

This list of definitions **may not** be exhaustive. Students are to refer to the lecture notes given. Thank you, Chemistry Department

	Term	**Definition**
1	Relative **isotopic** mass	The relative isotopic mass of an element is the **mass of an atom of the isotope** relative to $\frac{1}{12}$ of the mass of an atom of ^{12}C.
2	Relative **atomic** mass	The relative atomic mass of an element is the **average mass of one atom** of the element relative to $\frac{1}{12}$ of the mass of one atom of ^{12}C.
3	Relative **molecular** mass (molecules)	The relative molecular mass of a molecule is the **average mass of one molecule** of a substance relative to $\frac{1}{12}$ of the mass of an atom of ^{12}C.
4	Relative **formula** mass (ions/ ionic compounds)	The relative formula mass of an ionic compound is the **mass of one formula unit** of a substance relative to $\frac{1}{12}$ of the mass of one atom of ^{12}C.
5	**Mole**	One mole is the amount of substance that contains 6.02×10^{23} particles (atoms, molecules, ions, electrons).
6	**Empirical formula**	**Empirical formula** of a compound is the formula that shows the **simplest whole number ratio** of the atoms of the different elements in the compound.
7	**Molecular formula**	**Molecular formula** of a compound is one which shows the **actual number** of atoms of each element present in one molecule of a compound.

8	**Disproportionation**	A disproportionation reaction is a redox reaction in which one species is **simultaneously** oxidised and reduced.
9	Mass number/ **nucleon** number	Defined as the ***total number of protons and neutrons*** in the nucleus of an atom
10	Atomic number / Proton number	Defined as the ***number of protons*** in the nucleus of an atom
11	Isotopes	Isotopes are atoms of the same element with the **same number of protons** but ***different number of neutrons***.
12	First ionisation energy	The **first ionisation energy** of an element is defined as the **energy absorbed** to remove **one mole of electrons from one mole of gaseous atoms** to form **one mole of singly charged positive gaseous ions.** Example: $Ca\ (g) \rightarrow Ca^{+}\ (g) + e^{-}$ $\Delta H_1 = +590\ kJ\ mol^{-1}$
13	Second ionisation energy	The **second ionisation energy** of an element is defined as the **energy absorbed** to remove **one mole of electrons from one mole of singly charged positive gaseous ions** to form **one mole of doubly charged positive gaseous ions.** Example: $Ca^{+}\ (g) \rightarrow Ca^{2+}\ (g) + e^{-}$ $\Delta H_2 = +1150\ kJ\ mol^{-1}$
14	Atomic radius	**Atomic radius** is half the shortest inter-nuclear distance found in the structure of the element.
15	Electronegativity	**Electronegativity** is a measure of the **ability of an atom** in a molecule to **attract the shared pair of electrons** in a covalent bond.
16	Ionic radius	Ionic radius of an ion is the radius of the spherical ion in an ionic compound.
17	Ionic bond	Ionic bond is the **strong electrostatic attraction** between **two oppositely charged ions**, formed by the transfer of electrons between the original atoms.
18	Covalent bond	Covalent bond is the **strong** electrostatic attraction between **the positive nuclei of the two atoms and the bonding electrons shared** between them.

19	Co-ordinate bond / dative covalent bond	A **co-ordinate *bond*** (also called a ***dative covalent bond***) is a ***covalent bond*** (a shared pair of electrons) in which both electrons come from the same atom.
20	Bond energy	**Bond energy** is the **energy absorbed** when **one mole of a given covalent bond** in a **gaseous compound** is broken into **gaseous atoms or molecules.**
21	Metallic bond	Metallic bond is the **strong electrostatic forces of attraction** between the **positive metal ions and mobile sea of delocalised electrons.**
22	Lattice Energy	**Lattice energy** is defined as the **energy released** when **1 mole of ionic solid** is formed from its constituent **gaseous ions** (at infinite distance apart).
23	Hess' Law	The **enthalpy change for a reaction** is the same whether the reaction takes place in one step or in a series of steps and is **determined by the initial and final states of the system**.
24	Standard Enthalpy change of Formation (ΔH_f^θ)	The **energy change** when **one mole** of substance is **formed from its constituent elements** in their standard states **at 298 K and 1 bar**. E.g. $\Delta H_f^{\o}$ (Na_2CO_3): $2Na(s) + C(s) + 3/2O_2(g) \rightarrow \mathbf{Na_2CO_3(s)}$ Note: $\Delta H_f^{\o} = 0$ for elements in standard states.
25	Standard Enthalpy change of Combustion (ΔH_C^θ)	The **energy evolved** when **one mole** of substance is **burnt completely in excess oxygen at 298 K and 1 bar**. E.g. $\Delta H^{\o}_c$ (CH_3COOH): $\mathbf{CH_3COOH(l)} + 2O_2(g) \rightarrow 2CO_2(g) + 2H_2O(l)$ Note: $\Delta H_c^{\o}$ is always **exothermic**. $\Delta H_c^\theta = \dfrac{-\text{ heat change}}{n_{\text{substance burnt}}}$

26	Standard Enthalpy change of neutralisation ($\Delta H_{neu}^{\varnothing}$)	The **energy evolved** when **one mole of water is formed** in the **reaction between an acid and a base at 298 K and 1 bar**. $NaOH(aq) + \frac{1}{2} H_2SO_4(aq) \rightarrow \frac{1}{2} Na_2SO_4(aq) + \mathbf{H_2O(l)}$ Note: $\Delta H_{neu}^{\varnothing}$ is always **exothermic**. $D H^{\ominus}_{n} = \dfrac{\text{- heat change}}{n_{water}\ \text{formed}}$
	Standard Enthalpy Change of Atomisation ($\Delta H_{atom}^{\ominus}$)	**Standard enthalpy change of atomisation of an element** is the **energy absorbed** when **one mole of gaseous atoms is formed** from the element in its standard state **at 298 K and 1 bar**. $Na(s) \rightarrow Na(g)$ $\Delta H_{atom}^{\ominus}(Na) = +107\ kJ\ mol^{-1}$
	Standard Enthalpy Change of Hydration (ΔH_{hyd})	**Standard Enthalpy change of hydration** is the **energy evolved** when **one mole of gaseous ions** is converted to **one mole of aqueous ions** at infinite dilution at **298 K and 1 bar**. $Mg^{2+}(g) + aq \rightarrow Mg^{2+}(aq)$ $\Delta H_{hyd}(Mg^{2+}) = -1926\ kJ\ mol^{-1}$ $F^{-}(g) + aq \rightarrow F^{-}(aq)$ $\Delta H_{hyd}(F^{-}) = -524\ kJ\ mol^{-1}$
	Standard Enthalpy Change of Solution (ΔH_{sol})	**Standard Enthalpy change of solution** is the **energy change** when **one mole** of substance (in its standard state) is **completely dissolved** in infinite dilution in an **aqueous solution at 298K and 1 bar**. $NaCl(s) + aq \rightarrow Na^{+}(aq) + Cl^{-}(aq)$ $\Delta H_{sol}(NaCl) = +5\ kJ\ mol^{-1}$
	Electron Affinity (EA)	**First electron affinity** is the **energy change** when **one mole of gaseous atoms** of an element **gain one mole of electrons** to form **one mole of singly charged negative gaseous ions**.

		Second electron affinity is the **energy change** when **one mole of singly charged negative gaseous ions** of an element **gain one mole of electrons** to form **one mole of doubly charged negative gaseous ions**. $O (g) + e^- \rightarrow O^- (g)$ 1st EA (O) = -141 kJ mol^{-1} $O^- (g) + e^- \rightarrow O^{2-} (g)$ 2nd EA (O) = $+780$ kJ mol^{-1}
27	Bond Energy (BE)	The **energy absorbed** when **one mole of a given covalent bond** in a **gaseous compound** is broken into **gaseous atoms or molecules**. E.g. Bond energy of O – H: **O –H (g)** → O (g) + H (g) Note: BE is always **endothermic (bond breaking).**
	Entropy, S	**Entropy** is a measure of the disorder/ disorderliness of a system.
	Gibbs Free Energy	Gibbs Free Energy is a **function that relates a reaction's enthalpy and entropy terms to its spontaneity**. $\boldsymbol{\Delta G^\ominus = \Delta H^\ominus - T\Delta S^\ominus}$ where $\Delta G^\ominus$: change in standard Gibbs free energy (kJ mol^{-1}) $\Delta H^\ominus$: standard enthalpy change of the system (kJ mol^{-1}) $\Delta S^\ominus$: standard entropy change of the system (J mol^{-1} K^{-1}) T: temperature of the system (in K)
28	Rate of reaction	The rate of a reaction is defined as the **change of amount or concentration** of a reactant or product **per unit time**. $\textbf{Rate of reaction} = \dfrac{\textbf{Final Concentration – Initial Concentration}}{\textbf{Time taken for reaction to complete}}$ Units: **mol dm^{-3} s^{-1}** or **mol dm^{-3} min^{-1}** or **mol dm^{-3} h^{-1} [mol dm^{-3} time^{-1}]**
29	rate equation, order of reactions	**Rate equation** (or **rate law**) is the mathematical relationship between the reaction rate and the concentration of the individual reactants.

<table>
<tr><td></td><td>(m and n),
rate constant (*k*)</td><td>

$a\text{A} + b\text{B} \longrightarrow c\text{C}$

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

<table>
<tr><th></th><th>Meaning</th><th>Units</th></tr>
<tr><td>rate</td><td>rate of reaction</td><td>mol dm^{-3} time^{-1}</td></tr>
<tr><td>[A] or [B]</td><td>concentration of A or B</td><td>mol dm^{-3}</td></tr>
<tr><td>*m*</td><td>order of reaction with respect to A</td><td>no units</td></tr>
<tr><td>*n*</td><td>order of reaction with respect to B</td><td>no units</td></tr>
<tr><td>*m* + *n*</td><td>overall order of the reaction</td><td>no units</td></tr>
<tr><td>*k*</td><td>rate constant</td><td>Depends on the values of m & n</td></tr>
</table>

Rate constant (***k***) is the **proportionality constant** for a given reaction at a **certain temperature**.
Order of reaction (with respect to a reactant) is the **power** to which that reactant's concentration is raised to in an experimentally determined rate equation.
Overall order of reaction is the **sum** of the individual orders of reaction with respect to all the reactants.</td></tr>
<tr><td>**30**</td><td>reaction mechanism</td><td>A **reaction mechanism** shows the sequence or series of steps in a reaction.These steps are called **elementary steps** individually.</td></tr>
<tr><td>**31**</td><td>Homogeneous/ heterogeneous catalyst</td><td>**Homogeneous** catalysts are catalysts which are in **same phases** as the **reactants**.
Heterogeneous catalysts are catalysts which are in **different phases** as the **reactants**.</td></tr>
<tr><td></td><td>Autocatalysis</td><td>**Autocatalysis** is a type of catalysis whereby one of the **products (or intermediate)** in the reaction catalyses the reaction.</td></tr>
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<table>
<tr><td></td><td>Enzymes</td><td>**Enzymes** are **proteins** which catalyse reactions in living organisms (biochemical reactions). Enzymes are **highly specific** as each has an **active site** (specific region of the enzyme) into which **only one type of substrate** (reactant) **can fit**.</td></tr>
<tr><td>**32**</td><td>Dynamic equilibrium</td><td>Dynamic equilibrium refers to a reversible process at equilibrium in which the **rate of the forward reaction equals to the rate of backward reaction**, with the **concentrations of reactants and products** remaining **constant** over time.</td></tr>
<tr><td>**33**</td><td>Le Chatelier's Principle</td><td>**Le Chatelier's Principle** states that if a **system in dynamic equilibrium** is subjected to a **change** which **disturbs the equilibrium**, the system responds in such a way to **counteract** the **effect** of the change.</td></tr>
<tr><td>**34**</td><td>**Arrhenius Theory** of Acids and Bases</td><td>
<table>
<tr><th colspan="2">Arrhenius Theory</th><th>Equation</th></tr>
<tr><td>Acid</td><td>substance that gives out protons (H^+) in water</td><td>$HCl(aq) + H_2O(l) \longrightarrow H_3O^+(aq) + Cl^-(aq)$</td></tr>
<tr><td>Base</td><td>substance that gives out hydroxide (OH^-) in water</td><td>$NaOH\ (aq) \longrightarrow Na^+\ (aq) + OH^-\ (aq)$</td></tr>
</table>
</td></tr>
<tr><td>**35**</td><td>**Brønsted-Lowry Theory** of Acids and Bases</td><td>
<table>
<tr><th colspan="2">Brønsted-Lowry Theory</th><th>Equation</th></tr>
<tr><td>Acid</td><td>substance which donates protons (H^+) to a base (i.e. proton donor)</td><td>$HCl(aq) + H_2O(l) \longrightarrow Cl^-\ (aq) + H_3O^+(aq)$
Water is the base (**H^+ acceptor**).</td></tr>
<tr><td>Base</td><td>substance that accepts protons (H^+) from an acid (i.e. proton acceptor)</td><td>$NH_3(aq) + H_2O(l) \rightleftharpoons NH_4^+(aq) + OH^-(aq)$
Water is the acid (**H^+ donor**).</td></tr>
</table>
</td></tr>
<tr><td>**36**</td><td>**Lewis** Acids & Bases</td><td>
<table>
<tr><th colspan="2">Lewis Theory</th><th>Examples</th></tr>
<tr><td>Acid</td><td>substance that accepts an electron pair from a base to form a dative covalent bond (i.e. an electron-pair acceptor)</td><td>$\underset{\text{Lewis acid}}{BF_3} + \underset{\text{Lewis base}}{:NH_3} \qquad H_3N\rightarrow BF_3$</td></tr>
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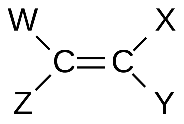
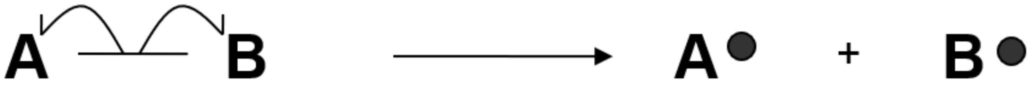
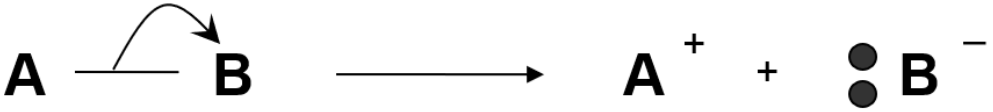
<table>
<tr><td></td><td></td><td><table><tr><td>Base</td><td>substance that donates an electron pair to an acid to form a dative covalent bond (i.e. an electron-pair donor)</td><td>$\underset{\text{Lewis acid}}{H^+} + \underset{\text{Lewis base}}{H_2O:} \longrightarrow H_3O^+$
$\underset{\text{Lewis acid}}{AlCl_3} + \underset{\text{Lewis base}}{:Cl^-} \longrightarrow AlCl_4^-$</td></tr></table></td></tr>
<tr><td>37</td><td>pH</td><td>The pH of a solution is defined as the negative logarithm to base ten of the hydrogen ion concentration in mol dm^{-3}.
$$\boxed{\mathbf{pH = -\lg [H^+]}}$$
Hence the concentration of H^+ can also be found from pH using $[H^+] = 10^{-pH}$</td></tr>
<tr><td>38</td><td>pOH</td><td>The pOH of a solution is defined as the negative logarithm to base ten of the hydroxide ion concentration in mol dm^{-3}.
$$\boxed{\mathbf{pOH = -\lg [OH^-]}}$$
Hence the concentration of OH^- can also be found from pH using $[OH^-] = 10^{-pOH}$
Relationship between pH and pOH at 298K (25°C)
$$\boxed{\mathbf{pH + pOH = 14}}$$
Hence pOH = 14 – pH at 25°C.</td></tr>
<tr><td>39</td><td>K_w</td><td>K_w, ionic product of water, is the equilibrium constant for the self-ionisation of water.
$$H_2O\,(l) \rightleftharpoons H^+\,(aq) + OH^-\,(aq)$$
$$K_w = [H^+][OH^-]$$
At 25 °C,
$$K_w = [H^+][OH^-] = 1.00 \times 10^{-14}\ \text{mol}^2\ \text{dm}^{-6}$$</td></tr>
</table>

		and consequently $\mathbf{p}K_w = pH + pOH = 14.0$
40	K_a	K_a is termed as the **acid dissociation constant** of acid HA. $HA\ (aq) \rightleftharpoons H^+\ (aq) + A^-\ (aq)$ $\boxed{K_a = \frac{[H_3O^+ (aq)]\ [A^- (aq)]}{[HA\ (aq)]} \quad \text{(units: mol dm}^{-3}\text{)}}$ $\boxed{pK_a = -\lg K_a}$
41	K_b	K_b is termed as the **base dissociation constant** of weak monoacidic base B. Similarly for a weak monoacidic base B (e.g. NH_3), that is partially ionised in aqueous solution, $B\ (aq) + H_2O\ (l) \rightleftharpoons BH^+ (aq) + OH^- (aq)$ $\boxed{K_b = \frac{[BH^+ (aq)]\ [OH^- (aq)]}{[B\ (aq)]} \quad \text{(units: mol dm}^{-3}\text{)}}$ The base dissociation constant, K_b, is a measure of the strength of a base. $\boxed{pK_b = -\lg K_b}$
42	Buffer solution	A buffer solution is a solution containing a weak acid and its conjugate base or a weak base and its conjugate acid which resists and **does not show significant changes in pH** when **small amount of strong acid or strong base are added to it.**

43	Constitutional isomers	Constitutional isomers are compounds that have the **same molecular formulae** but **different structural formulae.**
44	Stereoisomerism	Stereoisomerism occurs when compounds have the **same molecular formula** and the **same structural formula**, but the atoms in the molecules have **different spatial arrangements**.
45	Cis-trans isomerism	Cis-trans isomerism occurs in compounds where **free rotation about a bond is restricted** by the presence of multiple bonds, a ring structure or other steric factors (eg. large atoms). **Two criterion** for cis-trans isomerism: 1. **Restricted rotation about C=C double bond** (due to presence of of π bond). 2. **The 2 groups bonded to each carbon in the C=C bond must not be identical.**  where W ≠ Z, X ≠ Y **Maximum no. of cis-trans isomers = 2^n** where **n** is the number of C=C double bonds in molecule that shows cis-trans isomerism
46	Hybridisation	**Hybridisation** refers to the **mixing of atomic orbitals** in an atom to generate a set of hybrid orbitals.
47	Electrophile/ Nucleophile	**Electrophile** is an electron pair acceptor (Lewis acid). It is electron-deficient (or electron-poor) and is attracted to electron-rich region. E.g. AlCl3, HBr **Nucleophile** is an electron pair donor (Lewis base). It is electron-rich and is attracted to an electron-deficient region. E.g. NH_3, OH^-
48	Free radical	**Free radical** has an unpaired electron and is electrically neutral. E.g. $Cl\bullet$, NO_2

49	Homolytic fission	**Homolytic fission** is the breaking of a covalent bond in such a way that each fragment gets one of the shared electrons. Homolytic fission produces two free radicals. 
50	Heterolytic fission	**Heterolytic fission** is the breaking of a covalent bond such that the originally-shared pair of electrons goes to one of the fragments, forming a cation and an anion. 
51	Chiral centre	A **chiral centre** is an atom (usually carbon) that has four different substituent groups bonded to it.
52	Optically active	**Optically active** is the term used to describe **the ability of a substance to rotate the plane of polarization of a beam of light that is passed through it.**
55	Enantiomers	**Enantiomers are isomers that are non-superimposable mirror images of each other.** Enantiomers have identical chemical properties except in their interactions with another chiral molecule (eg. enzymes in biological systems). Enantiomers have identical physical properties except in the direction in which they rotate plane-polarised light.
54	Racemic mixture	A **racemic mixture** (or racemate) is one that has **equal amounts** of left-handed and right-handed enantiomers of a chiral molecule. A racemic mixture is **optically inactive**. o Although the two enantiomers rotate plane-polarised light in opposite directions, the rotations cancel out each other because the enantiomers are present in equal amounts. o There is **no net rotation of plane-polarised light**.
55	Meso compounds	**Meso compounds** are molecules that have:

		1. two or more chiral centres; 2. an internal plane of symmetry. Hence, meso compounds are not chiral compounds (i.e. they are **achiral**). A meso compounds is superimposable on its mirror image. Meso compounds are **optically inactive**.
56	Transition element	A transition element is a d-block element which forms **one or more** stable ions with a **<u>partially filled</u> d-subshell**.
57	Ligands	Ligands are **<u>neutral molecules</u>** or **<u>anions</u>** which contains at least one atom bearing **<u>a lone pair of electrons</u>** which can be **<u>donated</u>** into the **<u>low-lying vacant orbital</u>** of the central atom or ion **<u>forming a dative bond</u>**.
58	Complex ion	A complex ion the species formed when the number of ligands bonded to the central atom / cation exceeds the normal oxidation number of this atom.

<u>Useful formulas:</u>

For Chemical Energetics

In general, given $\Delta H_f^\ominus$ of the reactants and products,

$\Delta H_r = \sum m\Delta H_f^\ominus$ (Products) – $\sum n\Delta H_f^\ominus$ (Reactants)

where m = stoichiometric coefficients of the respective products
n = stoichiometric coefficients of the respective reactants

In general, given $\Delta H_c^\ominus$ of the reactants and products,

$\Delta H_r = \sum n\Delta H_c^\ominus$ (Reactants) – $\sum m\Delta H_c^\ominus$ (Products)

where m = stoichiometric coefficients of the respective products
n = stoichiometric coefficients of the respective reactants

In general, given BE of the reactants and products,

$\Delta H_r = \sum m$ Bond Energy (Reactants) – $\sum n$ Bond Energy (Products)

where m = total numbers of the respective bonds in the reactants
n = total numbers of the respective bonds in the products

- **For a solution,**

$$q = mc\Delta T$$

m = mass of **solution** (g)

c = specific heat capacity of solution (J g^{-1} K^{-1})

ΔT = change in temperature of the solution (K)

Assumptions: 1. density of water/aqueous solution is 1.0 g cm^{-3}

2. specific heat capacity of water/ aqueous solution is 4.18 J g^{-1} K^{-1} or 4.18 J cm^{-3} K^{-1}

(This means that 4.18 J of heat energy is required to increase the temperature of 1 g or 1 cm^3 *water/aqueous solution by 1* ^oC *or 1 K.)*

Chemical Eqm

Equilibrium constant K_c

Consider the following reversible reaction in equilibrium.

$$a\mathbf{A} + b\mathbf{B} \rightleftharpoons c\mathbf{C} + d\mathbf{D}$$

$$\boxed{K_c = \frac{[C]^c_{eqm}[D]^d_{eqm}}{[A]^a_{eqm}[B]^b_{eqm}}}$$

where **[ ]**$_{eqm}$ is the equilibrium concentration in mol dm^{-3}

Unit of K_c is **(mol dm^{-3})$^{c+d-a-b}$**

K_c is known as the ***equilibrium constant*** (expressed in terms of concentration).

Reaction Kinetics

Summary of Factors Affecting Rate of Reaction

Change	Effect on rate of reaction	Effect on rate constant, k	Remarks
Increase in concentration	Increases	No change	-
Increase in pressure	Increases	No change	Only applies to gaseous reactants
Increase in temperature	Increases	Increases	$k = Ae^{-\frac{E_a}{RT}}$
Use of catalyst	Increases	Increases	