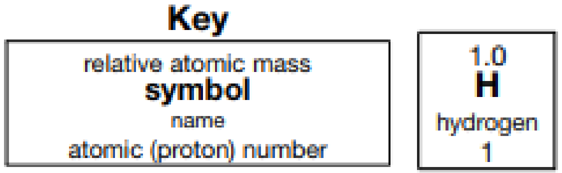
Unit 1: Atomic Structure

1. Keywords

Mass Number:	Total number of protons and neutrons (bigger no.)
Atomic number:	number of protons (smaller no.) also the number of electrons in an atom.
Isotope:	An atom that has the same number of protons but a different number of neutrons
Ionisation:	Removal of one or more electrons (endothermic)
First Ionisation Energy:	The energy needed to remove 1 electron from each atom in 1 mole of **gaseous** atoms. $M(g) \rightarrow M^{+}(g) + e^{-}$
Second ionisation energy:	The energy needed to remove 1 electron from each atom in 1 mole of **gaseous** +1 ions. $M^{+}(g) \rightarrow M^{2+}(g) + e^{-}$
Successive ionisation energies:	Removing each electron in turn from a mole of gaseous atoms. Provides evidence of energy levels and orbitals
Mass spectrometry	Technique used to calculate the mass of atoms and molecules



2. Time of Flight Mass Spec.

1	Ionisation	Sample dissolved and pushed through nozzle at high pressure and 4000v. As solvent evaporates particles gain a H+ ion
2	Acceleration	+ ions accelerated by -5000v electric field. Have a fixed kinetic energy
3	Ion drift	Region of no electric field, so drift (lighter move faster, heavier ions slower.)
4	Detection	+ ions discharge creating a flow of electrons in the detector which registers the current and plots the mass spectrum.

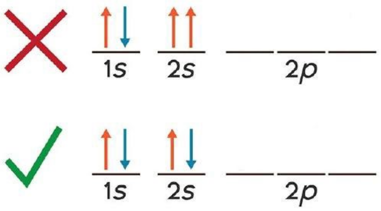
3. Rules for electron configuration

1	Aufbau "building up" principle	Electrons always fill the lowest energy level first
2	Hund's rule of maximum multiplicity (bus rule)	Electrons will fill the empty orbital of an energy level before pairing
3	Pauli's exclusion principle	When electrons pair in an orbital they have opposite spin

4. Energy levels

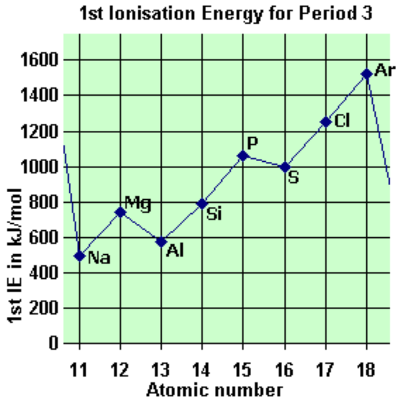
Energy level	Sub levels	Orbitals	Electrons
1	s	1	2
2	s,p	1,3	8
3	s,p,d	1,3,5	18
4	s,p,d	1,3,5	18

REMEMBER the 4s fills before the 3d (use you periodic table to remind you)



5. Drops in 1st ionisation energies across period 3

Ar	symbol	Electron config	Reason for drop?
13	Al	$3s^2\ 3p^1$	Electron is in p orbital further from nucleus to easier to remove
16	S	$3s^2\ 3p^4$	Electron is paired in 3p orbital so easier to remove



Unit 2: Amount of substance

1. Keywords

Relative Atomic Mass: (Ar)	Average mass of an atom $^1/_{12}$ th Carbon 12
Relative Molecular Mass: (Mr)	Average mass of a molecule $^1/_{12}$ th Carbon 12
Relative Isotopic Mass:	Average mass of an isotope $^1/_{12}$ th Carbon 12
Avogadro's constant	The number of particles that make up 1 mole of a substance.
Mole	The unit the amount of a substance is measured in. The number of particles need to make 12.00g of Carbon-12
Concentration	The amount of particles in a fixed volume. Measured in moles per litre (Mol dm^{-3})
Ideal gas	Ideal gases are any gas which behaves in accordance with the ideal gas equation. It does not matter what substances are in the gas.
Empirical formula	Simplest whole number ratio of the elements in a compound
Molecular formula	The actual ratio of elements in a specific compound. Should add up to the Mr.
Balanced full equation	A balanced chemical equation showing all atoms and their relative amounts and states
Ionic equation	An equation which only shows the species which change during a chemical reaction
Spectator ions	The ions omitted from an ionic equations because they are not involved
Atom economy	Mr desired product x100 Mr of all reactants

2. Calculating moles

$$Mass = Mr\ x\ moles$$

$Mass$	g
Mr	g mole^{-1}
$moles$	moles

3. Calculating concentration

$$Concentration = \frac{moles}{Volume}$$

$Concentration$	Mol dm^{-3}
$moles$	moles
$Volume$	dm^3

4. Ideal gas equation

$$pV = nRT$$

p	Pressure	Pa (pascals)	1 atm = 1x10^5 pa
V	Volume	m^3	1m^3 = 1x10^6 cm^3
n	No. of moles	Moles	
R	Boyles gas const.	J/mol K.	8.314
T	Temperature	K (kelvin)	T ^{o}C + 273

5. Ratio method for titration calculations

$$m_2C_1V_1 = m_1C_2V_2$$

m	Mole co-efficient (ratio)	
C	Concentration	Mol dm^{-3}
V	Volume	cm^3

Unit 3: Bonding

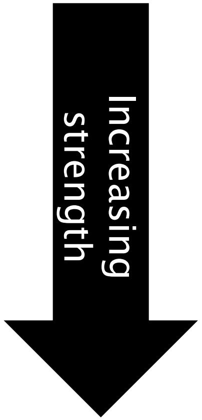
1. Keywords

Ionic bond	Bond formed by the transfer of electrons from a metal to a non-metal. Strong attraction between oppositely charged ions.
Covalent bond	Bond between non-metals. A shared pair of electrons
Metallic bond	Positive metal ions in a ‘sea’ of delocalised electrons
Macromolecular (Giant covalent)	Large covalently bonded molecule. Eg diamond, graphite, silicon dioxide
Molecular covalent molecule	Small covalently bonded molecules that are held together by intermolecular forces. Eg Iodine, water, carbon dioxide.
Co-ordinate bond	A type of covalent bond where both electrons are donated by one atom.
Bonding pair	A pair of electrons in a covalent bond
Lone pair	A pair of un-bonded electrons. Repel more than bonding pairs
Electronegativity	The power of an atom to attract the electrons in a covalent bond
Polar covalent bond	A bond with a unequal distribution of electrons due to a difference in electronegativity of the bonding atoms
Intermolecular forces	The forces between molecules. They are responsible for the trends in melting and boiling points of substances

2. Common anions

Sufate	SO_4^{2-}	Hydroxide	OH^-
Carbonate	CO_3^{2-}	Ammonium	NH_4^+
Nitrate	NO_3^-		

3. Intermolecular forces


Increasing strength

Van der Waals	Temporary dipoles induce complimentary dipoles in neighbours	Happens in all molecules	Eg. Alkanes
Permanent dipole- dipole	Attraction between slightly positive and negative ends of bond	Happens in any asymmetric bond with different electronegativity	Eg. Hydrogen chloride
Hydrogen bonding	Attraction between slightly positive and negative ends of bond	Happens when H bonded to O, N, F only	Eg. Water, Ammonia, Alcohol

4. VSEPR molecular shapes

Electron pairs	Geometry	Bonding pairs	Lone pairs	Shape	Angle	Example
2	Linear	2	0	Linear	180	$BeCl_2$
3	Trigonal planar	3	0	Trigonal planar	120	SO_3
4	Tetrahedral	4	0	Tetrahedral	109.5	CH_4
		3	1	Trigonal pyramidal	107	NH_3
		2	2	V-shape	104.5	H_2O
5	Trigonal bipyramidal	5	0	Trigonal bipyramidal	120, 90	PCl_5
		4	1	See-saw	120, 90	$TeCl_4$
		3	2	T-shape	87.5	ClF_3
6	Octahedral	6	0	Octahedral	90	SF_6
		4	2	Square planar	90	ICl_4^-

Unit 4: Energetics

1. Keywords

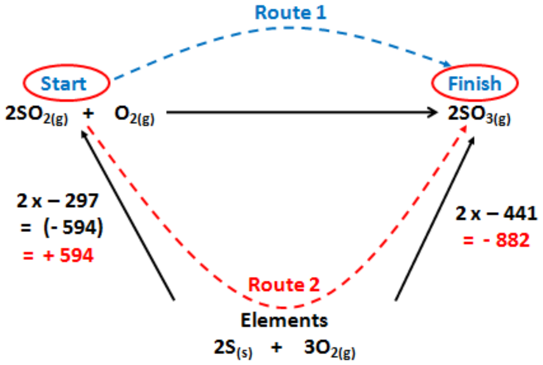
Enthalpy change ΔH	Heat energy change at constant pressure
Standard enthalpy change ΔH_{298}°	Enthalpy change under standard conditions eg 100Kpa and 298K
Standard enthalpy change of formation $\Delta_f H^{\circ}$	The enthalpy change when one mole of substance is formed from it's elements under standard conditions
Standard enthalpy change of combustion $\Delta_c H^{\circ}$	The enthalpy change when one mole of substance is completely burnt in oxygen. Reactants and products in their standard states under standard conditions.
Calorimetry	The process of measuring the heat from a chemical reaction
Hess' law	The enthalpy change of a reaction is independent of the route taken
Mean bond enthalpy	The average enthalpy change when one mole of a specific bond is broken in a range of different **gaseous** compounds

2. Calorimetry

$$q = mc\Delta T$$

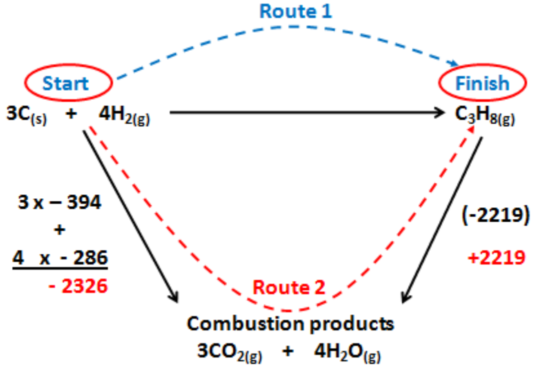
q	Energy / J	c	Specific heat capacity / J/Kg°C
m	mass / g	ΔT	Temperature change / °C

3. Enthalpy of formation cycle



Formation: Arrows go up

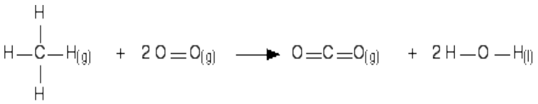
4. Enthalpy of combustion cycle



Combustion: Arrows go down

5. Bond enthalpies

Shopping list approach

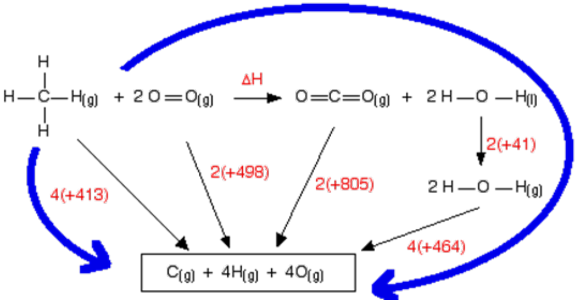


$$CH_{4(g)} + 2\,O{=}O_{(g)} \longrightarrow O{=}C{=}O_{(g)} + 2\,H{-}O{-}H_{(l)}$$

C-H	4x (+413)	C=O	2x (+805)
O=O	2x (+498)	O-H	4x (+464)
	= +2348		= +3466

= +2348 - +3466 = -1118kJmol^{-1}

Cycle approach



= +2348 - +3466 = -1118kJmol^{-1}

Unit 5: Kinetics

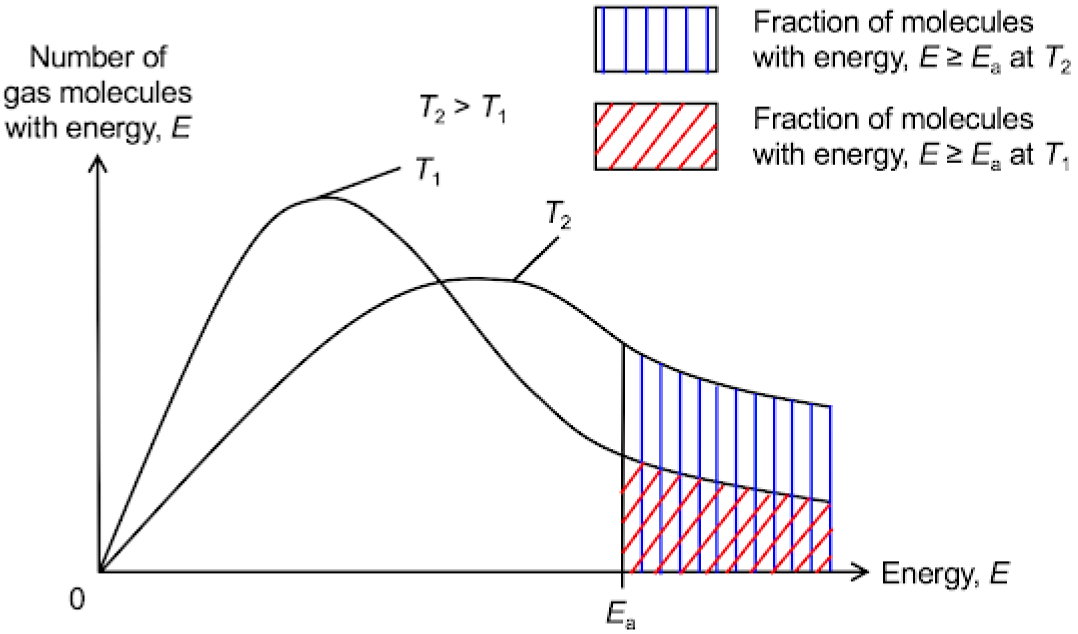
1. Keywords

Collision theory	For a reaction to occur the reactants have to collide with sufficient energy and in the correct orientation
Activation energy	The minimum energy needed for a collision to create a successful reaction
Rate of reaction	The amount of product made or reactant used up in a given time

2. Factors the increase the rate of reaction

concentration	More particles in a given space so more collisions and increased rate
Pressure	More particles in a given space so more collisions and increased rate
Surface area	High surface area increases the number of particles available to collide so increases rate
Temperature	Increases the kinetic store of the particles so they collide more frequently and with a greater proportion above the activation energy
Catalyst	Speed up reaction by providing an alternative route with a lower activation energy

3. Maxwell-Boltzmann distribution



Line starts at origin	No particles have no energy
Asymptotic	some particles have infinite energy
Area under the curve	Number of particles
Height of curve	Most probable energy

Unit 6: Equilibria Kc

1. Keywords

Equilibrium	When a reversible reaction has the same rate in both directions and the concentrations do not change. Happens in a closed system
Le Charteliers principle	If a change is made to a system at equilibrium then the equilibrium will shift to oppose that change

3. Kc

$$aA + bB \rightleftharpoons cC + dD$$

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

[ ]	Concentration / mol dm^{-3}
Kc >1	In favour of products
Kc<1	In favour of reactants

2. Le Charteliers principle

Factor	Effect on equilibrium	Reason	Example
Pressure Increase	Shifts to side with least moles	To decrease the pressure to reach the equilibrium pressure again	It will shift to the right as it has the least moles
Pressure Decrease	Shifts to side with most moles	To increase the pressure	It will shift to the left as it has the most moles
Concentration Increase	Shifts to other side	To decrease the concentration of what was increased	If the concentration of A is increased it will shift to the right
Concentration Decrease	Shifts to the same side	To increase the concentration of what was decreased	If the concentration of A is decreased it will shift to the left
Temperature Increase	Shifts in the endothermic direction	To cool down the reaction	It will shift to the left as this is the endothermic direction
Temperature Decrease	Shifts in the exothermic direction	To heat up the reaction	It will shift to the right as it is the exothermic direction

4. Units of Kc

$$2A + B \rightleftharpoons C$$

$$\downarrow$$

$$\frac{moldm^{-3}}{(moldm^{-3})^2(moldm^{-3})} \nearrow \frac{\cancel{moldm^{-3}}}{(moldm^{-3})^2\cancel{(moldm^{-3})}}$$

$$= \frac{1}{(moldm^{-3})^2}$$

$$= mol^{-2}dm^{6}$$

Unit 7: Redox

1. Keywords

Oxidation	The loss of electrons, gain of oxygen, increase in oxidation number
Reduction	The gain of electrons, loss of oxygen, decrease in oxidation number
Disproportionation	When one species is both oxidised and reduced simultaneously
Half equation	A balanced equation which only shows either the oxidation or reduction during a redox reaction
Oxidising agent	A substance which oxidises another substance and in doing so becomes reduced
Reducing agent	A substance which reduces another substance and in doing so becomes oxidised

2. Rules to assign oxidation numbers

Element	0
Mono atomic ion	Charge on the ion
Hydrogen	+1 (unless hydride = -1)
Oxygen	-2 (unless peroxide = -1)
Fluorine	-1
Group 1	+1
Group 2	+2
Group 7	-1 mostly

3. Steps to balance half equations

1 Balance the existing atoms

2 Determine the oxidation state before and after the reaction

3 Add electrons to balance the change in oxidation state

4 If it is a complex ion add H^+ ions to produce 1 water for every oxygen

5 Check the atoms and charge balance both sides

4. Steps to combine half equations

1 Determine which is being oxidised and which is being reduced

2 Write out the balanced half equations

3 Factor up the half equations until the number of electrons balance

4 Combine the half equations

5 Cancel any redundant species

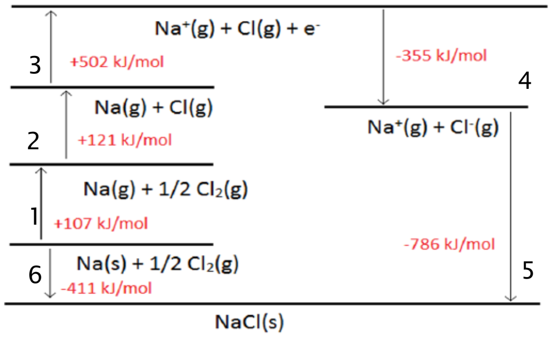
Unit 8: Thermodynamics

1. Keywords

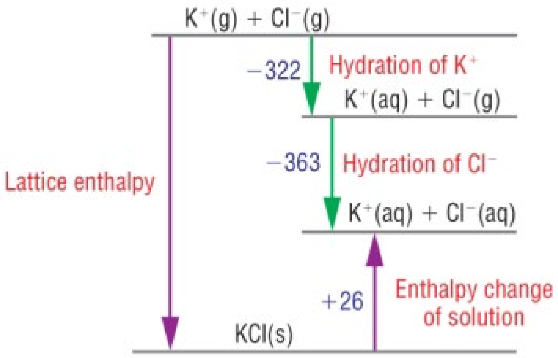
Standard enthalpy change of atomisation $\Delta_{at}H^\circ$	The enthalpy change when one mole of **gaseous** atoms are formed from a substance at its standard states
Ionisation enthalpy $\Delta_{ie}H^{1st}$	The enthalpy change when one mole of ions are formed from one moles of **gaseous** atoms under standard conditions
Electron affinity $\Delta H_{ea}{}^{1st}$	The enthalpy change when one mole of **gaseous** atoms gain one mole of electrons under standard conditions
Lattice association enthalpy $\Delta_{Lf}H$	The enthalpy change when one mole of ionic lattice is *formed* from its **gaseous** ions under standard conditions
Lattice dissociation enthalpy $\Delta_{Ld}H^\circ$	The enthalpy change when one mole of an ionic lattice dissociates into isolated gaseous ions under standard conditions
Hydration enthalpy $\Delta_{Hy}H^\circ$	The enthalpy change when one moles of gaseous ions are completely surrounded by water
Enthalpy of solution $\Delta_{sol}H^\circ$	The enthalpy change when I mole of solute is completely dissolved in solvent so that the ions are infinitely diluted, under standard conditions.
Entropy	The measure of disorder within a system measured in J mol^{-1}
Entropy change ΔS	$\Sigma\, entropy\ products - \Sigma\, entropy\ reactants$
Gibbs free energy ΔG	A measure of the feasibility of a reaction. For a reaction to be feasible ΔG must ≤ 0

2. Born-Haber cycle: basic layout

1	Atomise
2	Atomise
3	Ionise
4	Electron affinity
5	Lattice enthalpy
6	Enthalpy of formation



3. Born-Haber cycle: enthalpy of solution



4. Gibbs free energy

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

ΔG	Gibbs free energy / KJ
ΔH	Enthalpy change / KJ mol^{-1}
T	Temperature / K
ΔS	Entropy change / **J mol^{-1}**

Unit 9: Rate equations

1. Rate equation

$$Rate = k[A]^m[B]^n$$

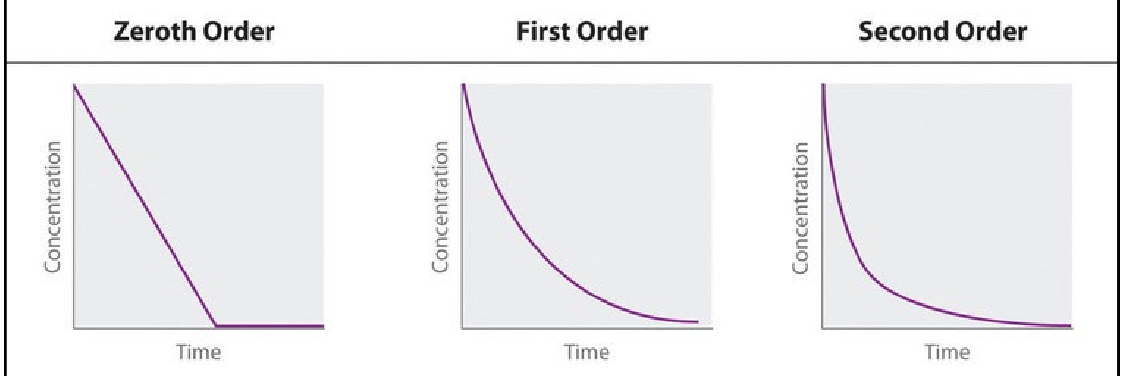
$Rate$	The rate of reaction	Mol dm^{-3} s^{-1}
k	The rate constant (temperature dependent)	variable
$[A]$	Concentration of A	Mol dm^{-3}
m	Order of reaction with respect to A	
$[B]$	Concentration of B	Mol dm^{-3}
n	Order of reaction with respect to B	

2. Arrhenius equation

$$k = Ae^{-Ea/RT}$$

k	The rate constant (temperature dependent)	variable
A	Arrhenius constant	S^{-1}
e	Euler's number (magic number e)	2.71
Ea	Activation energy	KJ mol^{-1}
R	Boyles gas constant	8.31 J/mol K
T	Temperature	K

3. Determining order of reaction graphically

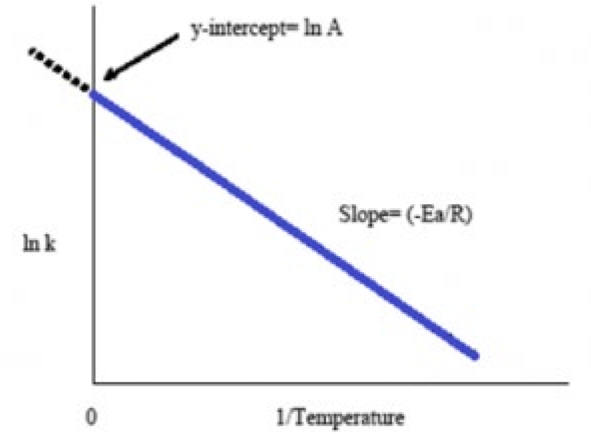


4. Determining the activation energy graphically

$$k = -\frac{Ea}{RT} + lnA$$

$$y = mx + c$$

Y axis	k	Rate constant
X axis	$\frac{1}{T}$	Reciprocal of temperature
m	$-\frac{Ea}{R}$	Multiply by -R to determine Ea



Unit 10: Equilibria Kp

1. Keywords

Mole fraction	$\frac{\text{The number of moles of a species}}{\text{The total number of moles}}$
Partial pressure	The mole fraction of a species x total pressure

2. Kp Expression

$$K_p = \frac{(P_C)^c\,(P_D)^d}{(P_A)^a\,(P_B)^b}$$

Kp	Equilibrium constant	Variable units
(P_c)	Partial pressure of C	Pascals
c	Order with respect to C	
(P_D)	Partial pressure of D	Pascals
d	Order with respect to D	
(P_A)	Partial pressure of A	Pascals
a	Order with respect to A	
(P_B)	Partial pressure of B	Pascals
b	Order with respect to B	

Unit11: Electrode potentials and electrochemical cells

1. Keywords

Potential difference	The difference in voltage between two points.
Electrode potential E°	The is the potential difference of a cell built of two electrodes: on the left-hand side of the cell diagram is the standard hydrogen electrode (SHE), and. on the right-hand side is the electrode in question under standard conditions (1mol dm^{-3}, 100kPa, 298K)
Standard hydrogen electrode (SHE)	The electrode given the electrode potential of 0.00v to establish all other electrode potentials
Platinum electrode	Unreactive electrode used in half cells when neither species is a solid metal
Electrochemical cell	A cell produced when 2 half cells of different electrode potentials are linked by a salt bridge
Salt bridge	A medium connecting two half cells. Contains inert ions to allow charge to flow without interfering with the electrochemistry
EMF	Electro motive force. The potential difference of a cell when no current flows
Feasible reaction	A spontaneous redox reaction which generates a positive E° for the cell
Anode	The electrode where oxidation happens. In an electrochemical cell this is the negative terminal
Cathode	The electrode where reduction happens. In an electrochemical cell this is the positive terminal. Often on the right
Non-rechargeable cell	A cell with a spontaneous reaction which cannot be reverse
Rechargeable cell	A cell with a spontaneous reaction which can be reversed by applying an electric current
Fuel cell	A cell which generates an EMF providing a continuous flow of chemicals are provided
Hydrogen fuel cell	A fuel cell which uses hydrogen and oxygen to generate an EMF. Water is the only waste product

2. Standard hydrogen electrode

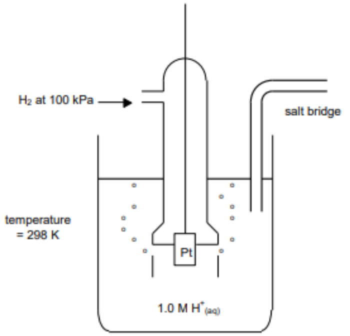
Half cell equation:

$$2\,H^+_{(aq)} + 2\,e^- \rightleftharpoons H_{2(g)}$$

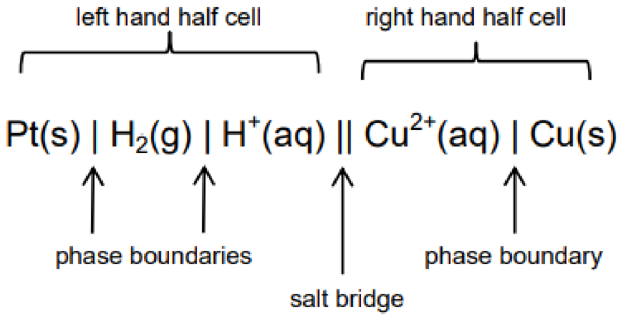
$$E^\circ = +0.00\ V$$

Cell notation:

$$Pt(s)\ |\ H_2(g)\ |\ H^+(aq)$$



3. How to write conventional cell notation



$$\underbrace{Pt(s)\ |\ H_2(g)\ |\ H^+(aq)}_{\text{left hand half cell}}\ ||\ \underbrace{Cu^{2+}(aq)\ |\ Cu(s)}_{\text{right hand half cell}}$$

Highest oxidation state is nearest the salt bridge

If platinum electrode is present it goes on the far edges

4. Calculating the EMF of a cell

$$EMF = E_{cell} = E_{right} - E_{left}$$

Note: SHE always goes on the left

Unit 12: Acids and Bases

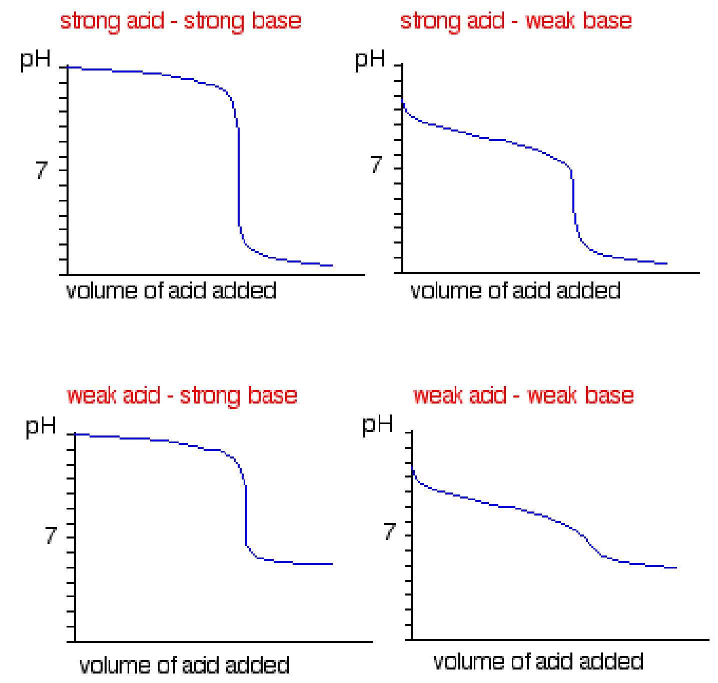
1. Keyword

Bronsted-Lowry acid	Proton donor
Bronsted-Lowry base	Proton acceptor
Conjugate base	An **acid** that has donated its hydrogen
Conjugate acid	A **base** with a hydrogen ion added to it.
Strong	Fully ionises/dissociates in solution
Weak	Partially ionises/dissociates in solution
pH	$-\mathrm{Log}_{10}\,[H^+]$
Kw	Ionic product of water = $[H^+]\,[OH^-] = 1\mathrm{x}10^{-14}$ At 298k
Ka	Acid dissociation constant. A measure of dissociation of a weak acid. Units = Mol dm^{-3}
Half equivalence point	Half the volume required to neutralise and acid or base = pKa
Indicator	A compound or compounds that change colour when pH changes. Should fall in the vertical rise on a pH curve
10^x	Inverse log button on your calculator duh!
Buffer	Solution which resists small changes in pH. Made from a weak acids and its salt.

2. Useful acids

Name	Formula	Strength	Protic
Hydrochloric	HCl	Strong	Mono
Sulphuric	H_2SO_4	Strong	Di
Nitric	HNO_3	Strong	Mono
Phosphoric	H_3PO_4	Strong	Tri
Ethanoic	CH_3COOH	Weak	Mono
Ethanedioic	$HOOCCOOH$	Weak	Di

3. pH curves



4. Ka Expression

$$HA + H_2O \rightleftharpoons H_3O^+ + A^-$$

$$Ka = \frac{[H_3O^+][A^-]}{[HA]}$$