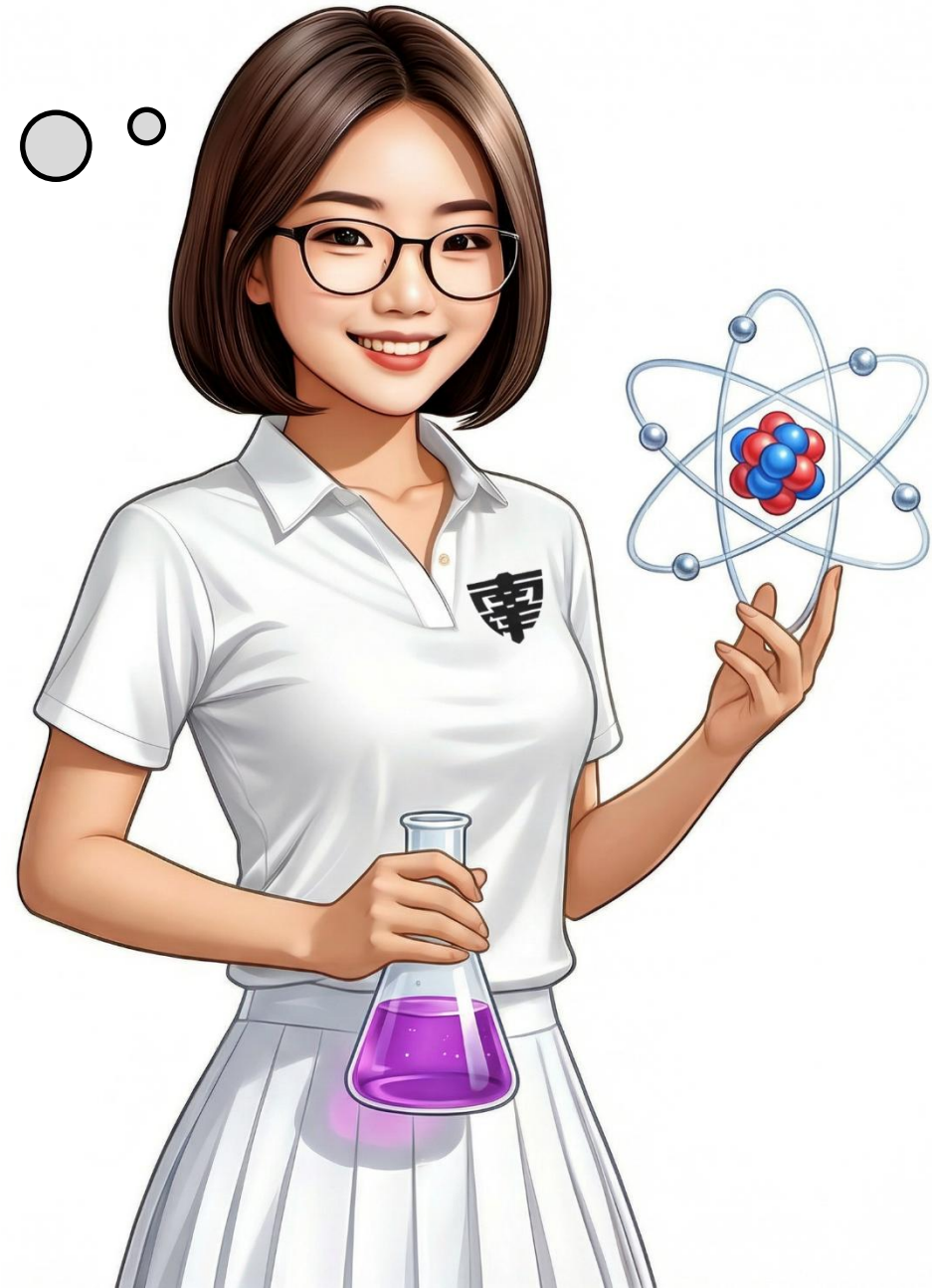


Introduction to Atomic Structure

What do I
need to know
about atomic
structure?



Atomic Structure – Knowledge and Skills

Atomic structure

- (a) State the relative charges and approximate relative masses of a proton, a neutron and an electron.
- (b) Describe, with the aid of diagrams, the structure of an atom as containing protons and neutrons (nucleons) in the nucleus and electrons arranged in shells (energy levels). A copy of the Periodic Table will be given in all examinations.
- (c) Define proton (atomic) number and nucleon (mass) number.
- (d) Interpret and use symbols such as $^{12}_6\text{C}$.
- (e) Define the term isotope.
- (f) Deduce the numbers of protons, neutrons and electrons in atoms and ions given proton and nucleon numbers.

The Periodic Table of Chemical Elements

Group																							
1	2											13	14	15	16	17	18						
		<div style="border: 1px solid black; padding: 5px; width: fit-content; margin: auto;"> Key atomic number atomic symbol name relative atomic mass </div>										1											2
																						H hydrogen 1.0	
3	4											5	6	7	8	9	10						
Li lithium 6.9	Be beryllium 9.0											B boron 10.8	C carbon 12.0	N nitrogen 14.0	O oxygen 16.0	F fluorine 19.0	Ne neon 20.2						
11	12	3	4	5	6	7	8	9	10	11	12	13	14	15	16	17	18						
Na sodium 23.0	Mg magnesium 24.3											Al aluminium 27.0	Si silicon 28.1	P phosphorus 31.0	S sulfur 32.1	Cl chlorine 35.5	Ar argon 39.9						
19	20	21	22	23	24	25	26	27	28	29	30	31	32	33	34	35	36						
K potassium 39.1	Ca calcium 40.1	Sc scandium 45.0	Ti titanium 47.9	V vanadium 50.9	Cr chromium 52.0	Mn manganese 54.9	Fe iron 55.8	Co cobalt 58.9	Ni nickel 58.7	Cu copper 63.5	Zn zinc 65.4	Ga gallium 69.7	Ge germanium 72.6	As arsenic 74.9	Se selenium 79.0	Br bromine 79.9	Kr krypton 83.8						
37	38	39	40	41	42	43	44	45	46	47	48	49	50	51	52	53	54						
Rb rubidium 85.5	Sr strontium 87.6	Y yttrium 88.9	Zr zirconium 91.2	Nb niobium 92.9	Mo molybdenum 95.9	Tc technetium –	Ru ruthenium 101.1	Rh rhodium 102.9	Pd palladium 106.4	Ag silver 107.9	Cd cadmium 112.4	In indium 114.8	Sn tin 118.7	Sb antimony 121.8	Te tellurium 127.6	I iodine 126.9	Xe xenon 131.3						
55	56	57–71	72	73	74	75	76	77	78	79	80	81	82	83	84	85	86						
Cs caesium 132.9	Ba barium 137.3	lanthanoids	Hf hafnium 178.5	Ta tantalum 180.9	W tungsten 183.8	Re rhenium 186.2	Os osmium 190.2	Ir iridium 192.2	Pt platinum 195.1	Au gold 197.0	Hg mercury 200.6	Tl thallium 204.4	Pb lead 207.2	Bi bismuth 209.0	Po polonium –	At astatine –	Rn radon –						
87	88	89–103	104	105	106	107	108	109	110	111	112	113	114	115	116	117	118						
Fr francium –	Ra radium –	actinoids	Rf rutherfordium –	Db dubnium –	Sg seaborgium –	Bh bohrium –	Hs hassium –	Mt meitnerium –	Ds darmstadtium –	Rg roentgenium –	Cn copernicium –	Nh nihonium –	Fl flerovium –	Mc moscovium –	Lv livermorium –	Ts tennessine –	Og oganesson –						

lanthanoids	57	58	59	60	61	62	63	64	65	66	67	68	69	70	71
	La lanthanum 138.9	Ce cerium 140.1	Pr praseodymium 140.9	Nd neodymium 144.2	Pm promethium –	Sm samarium 150.4	Eu europium 152.0	Gd gadolinium 157.3	Tb terbium 158.9	Dy dysprosium 162.5	Ho holmium 164.9	Er erbium 167.3	Tm thulium 168.9	Yb ytterbium 173.1	Lu lutetium 175.0
actinoids	89	90	91	92	93	94	95	96	97	98	99	100	101	102	103
	Ac actinium –	Th thorium 232.0	Pa protactinium 231.0	U uranium 238.0	Np neptunium –	Pu plutonium –	Am americium –	Cm curium –	Bk berkelium –	Cf californium –	Es einsteinium –	Fm fermium –	Md mendelevium –	No nobelium –	Lr lawrencium –

A' Level Periodic Table of the Chemical Elements

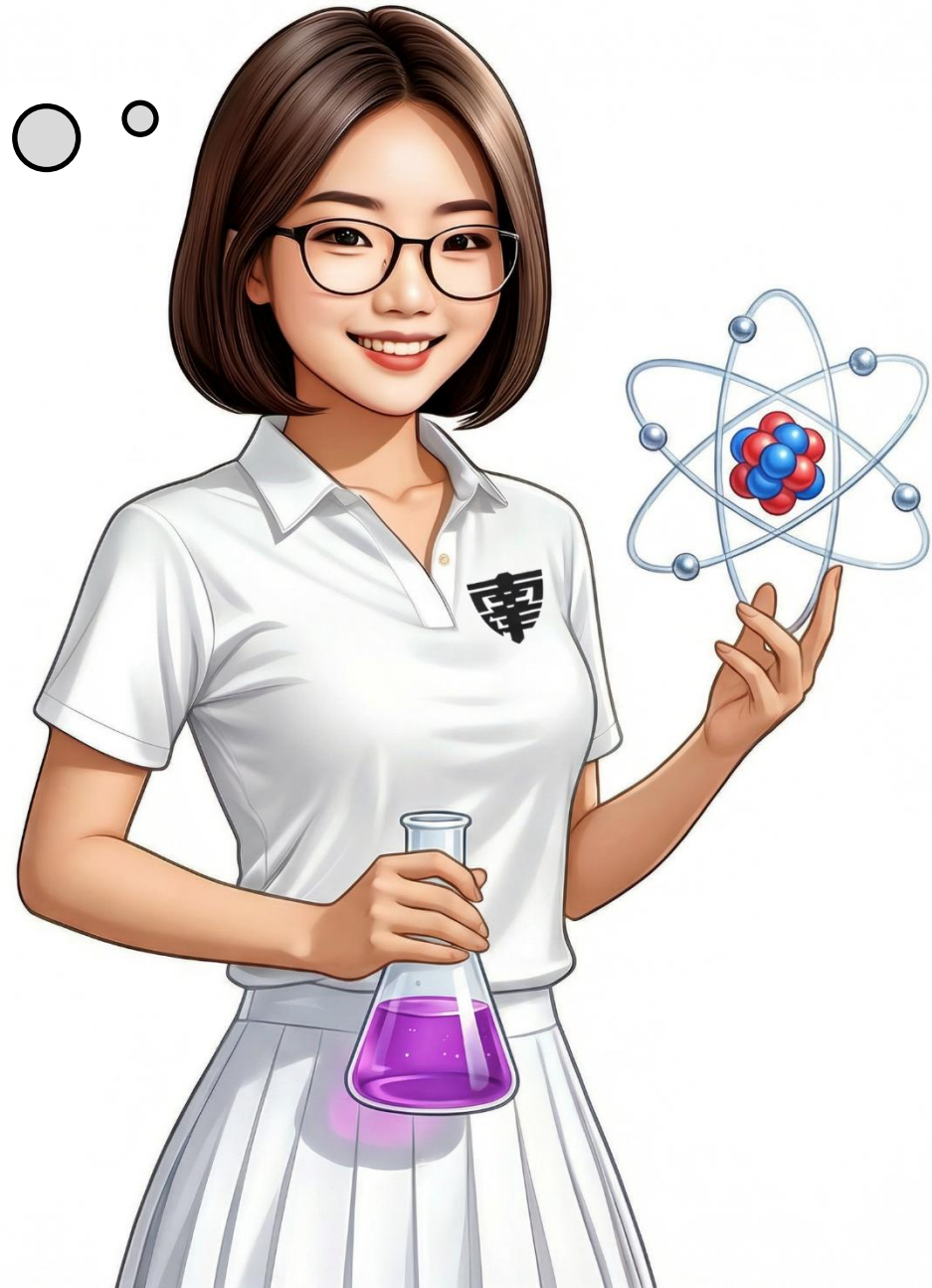
Group																				
1	2													16	17	18				
																2 He helium 4.0				
<div style="display: flex; justify-content: space-between;"> <div style="border: 1px solid black; padding: 5px;"> <p>Key</p> <p>atomic number</p> <p>atomic symbol</p> </div> <div style="border: 1px solid black; padding: 5px;"> <p>hydrogen</p> <p>1.0</p> </div> </div>																				
3 Li lithium 6.9	4 Be beryllium 9.0													5 B boron 10.8	6 C carbon 12.0	7 N nitrogen 14.0	8 O oxygen 16.0	9 F fluorine 19.0	10 Ne neon 20.2	
11 Na sodium 23.0	12 Mg magnesium 24.3	3 Al aluminium 27.0	4 Si silicon 28.1	5 P phosphorus 31.0	6 S sulfur 32.1	7 Cl chlorine 35.5	8 Ar argon 39.9											16 S sulfur 32.1	17 Cl chlorine 35.5	18 Ar argon 39.9
19 K potassium 39.1	20 Ca calcium 40.1	21 Sc scandium 45.0	22 Ti titanium 47.9	23 V vanadium 50.9	24 Cr chromium 52.0	25 Mn manganese 54.9	26 Fe iron 55.8	27 Co cobalt 58.9	28 Ni nickel 58.7	29 Cu copper 63.5	30 Zn zinc 65.4	31 Ga gallium 69.7	32 Ge germanium 72.6	33 As arsenic 74.9	34 Se selenium 79.0	35 Br bromine 79.9	36 Kr krypton 83.8			
37 Rb rubidium 85.5	38 Sr strontium 87.6	39 Y yttrium 88.9	40 Zr zirconium 91.2	41 Nb niobium 92.9	42 Mo molybdenum 95.9	43 Tc technetium -	44 Ru ruthenium 101.1	45 Rh rhodium 102.9	46 Pd palladium 106.4	47 Ag silver 107.9	48 Cd cadmium 112.4	49 In indium 114.8	50 Sn tin 118.7	51 Sb antimony 121.8	52 Te tellurium 127.6	53 I iodine 126.9	54 Xe xenon 131.3			
55 Cs caesium 132.9	56 Ba barium 137.3	57-71 lanthanoids	72 Hf hafnium 178.5	73 Ta tantalum 180.9	74 W tungsten 183.8	75 Re rhenium 186.2	76 Os osmium 190.2	77 Ir iridium 192.2	78 Pt platinum 195.1	79 Au gold 197.0	80 Hg mercury 200.6	81 Tl thallium 204.4	82 Pb lead 207.2	83 Bi bismuth 209.0	84 Po polonium -	85 At astatine -	86 Rn radon -			
87 Fr francium -	88 Ra radium -	89-103 actinoids	104 Rf rutherfordium -	105 Db dubnium -	106 Sg seaborgium -	107 Bh bohrium -	108 Hs hassium -	109 Mt meitnerium -	110 Ds darmstadtium -	111 Rg roentgenium -	112 Cn copernicium -	113 Nh nihonium -	114 Fl flerovium -	115 Mc moscovium -	116 Lv livermorium -	117 Ts tennessine -	118 Og oganesesson -			

- The *Groups* are numbered in Arabic numerals from 1 to 18.

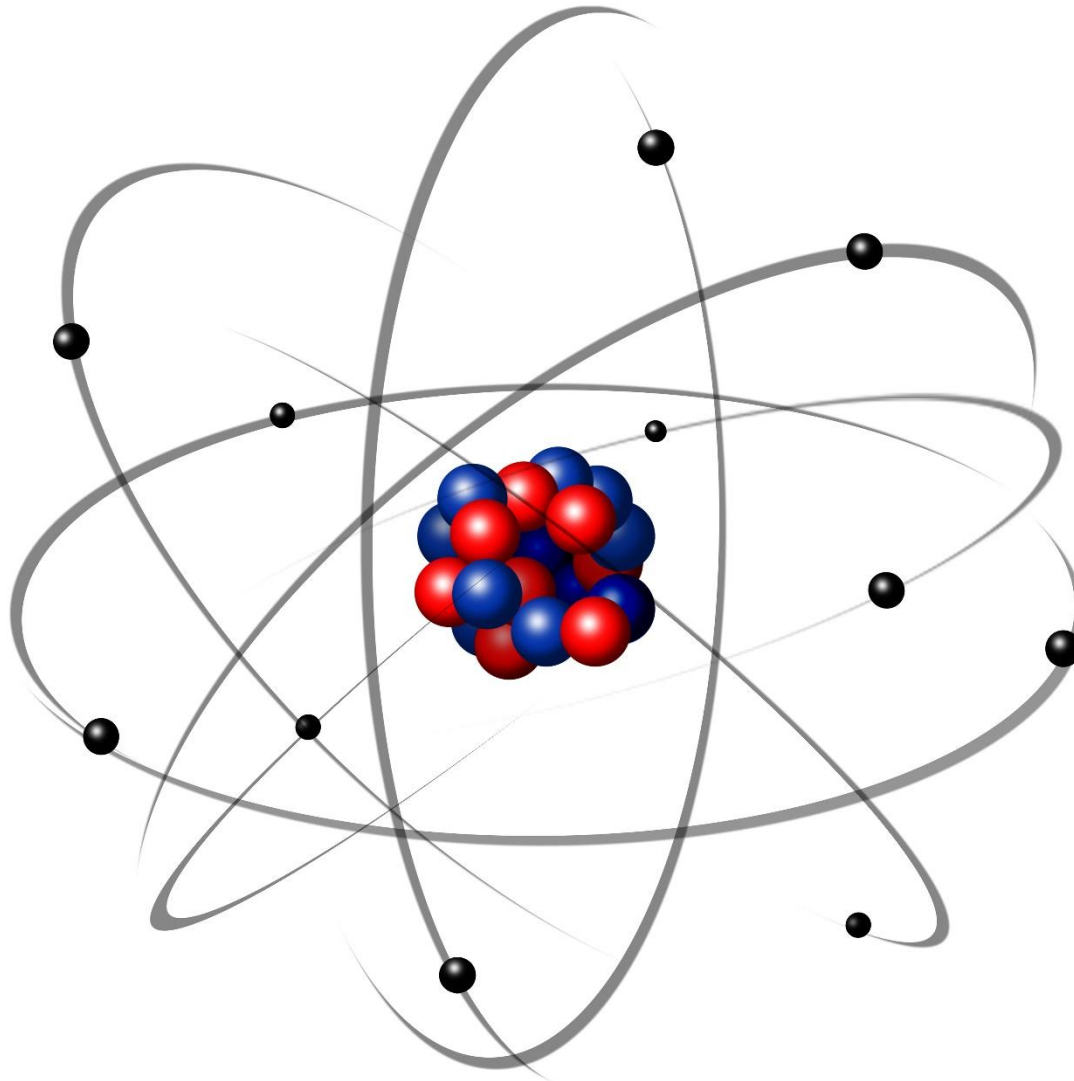
- The *relative atomic masses* are all written to one decimal place.

- The *atomic number* is written *above* the symbol of the chemical element, while the *relative atomic mass* is written *below* the symbol of the chemical element.

What is the
modern theory
of atomic
structure?

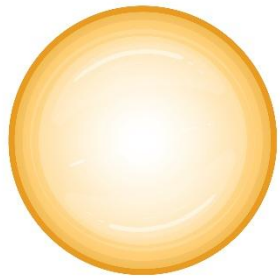


Introduction to Atomic Structure

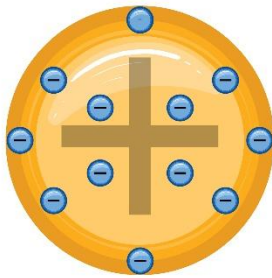


Introduction to Atomic Structure

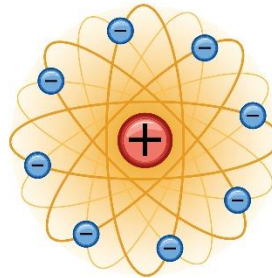
Atomic Models



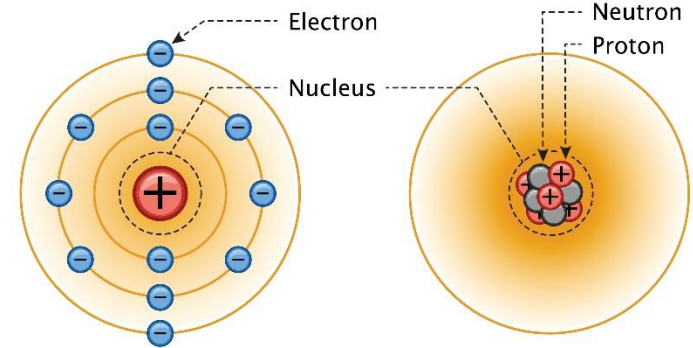
Solid sphere model
(Dalton, 1803)



Plum pudding model
(Thomson, 1897)



Nuclear model
(Rutherford, 1911)

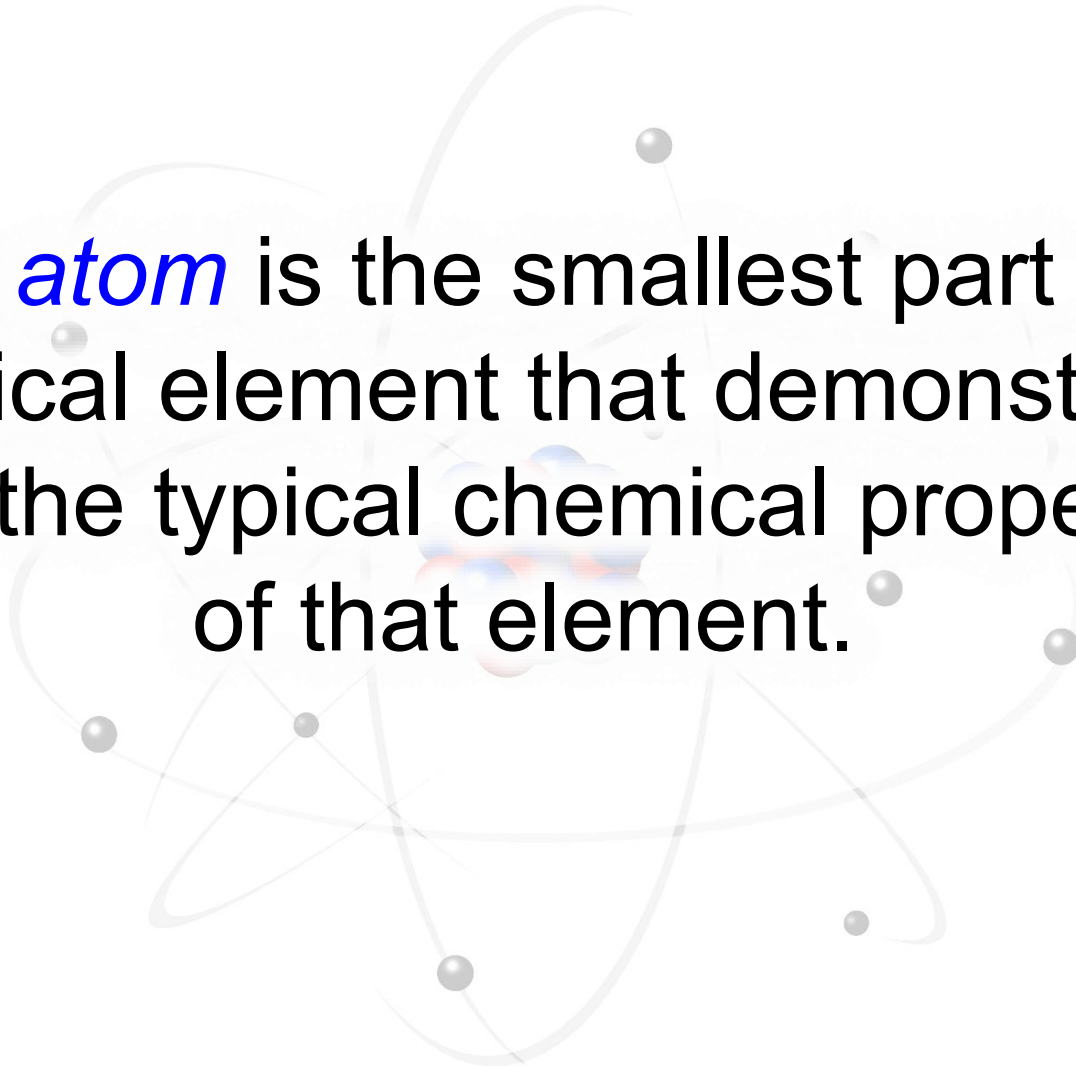


Planetary model
(Bohr 1913)

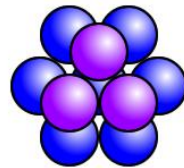
Quantum model
(Schrödinger, 1926)

Introduction to Atomic Structure

- An *atom* is the smallest part of a chemical element that demonstrates all of the typical chemical properties of that element.



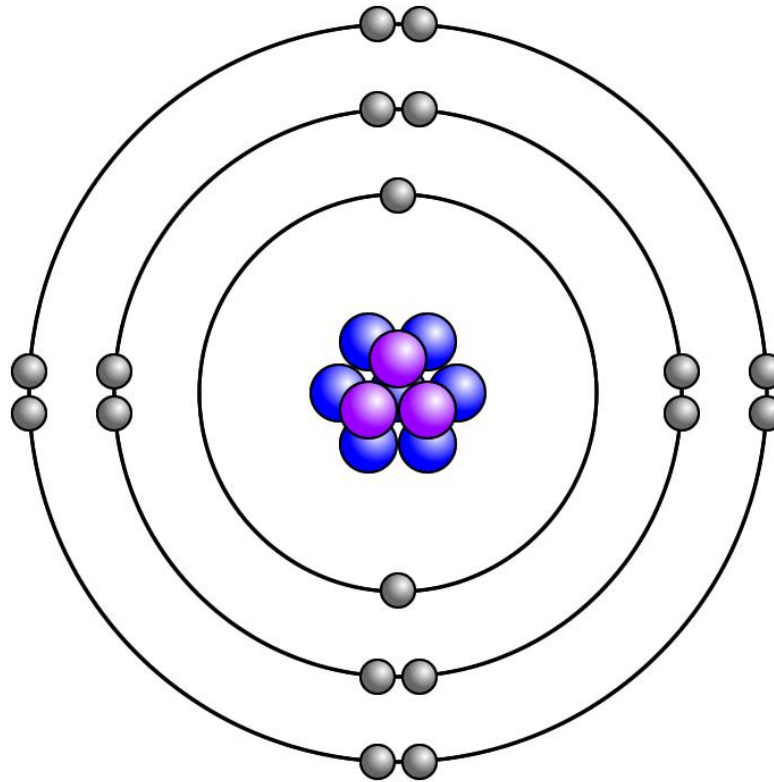
Introduction to Atomic Structure



The atom is composed of a small, dense *nucleus* that contains *protons* and *neutrons*.

- Relative mass of a *proton* = 1
- Relative mass of a *neutron* = 1
- Relative charge on a *proton* = +1
- Relative charge on a *neutron* = 0

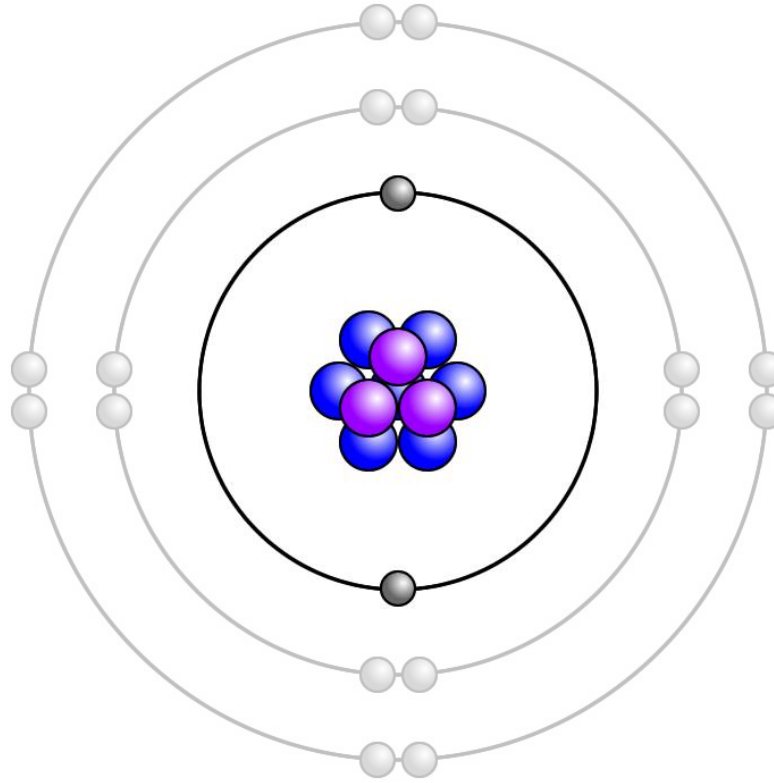
Introduction to Atomic Structure



Electrons orbit the nucleus in specific energy levels called *electron shells*. The outermost electron shell is called the *valence electron shell*.

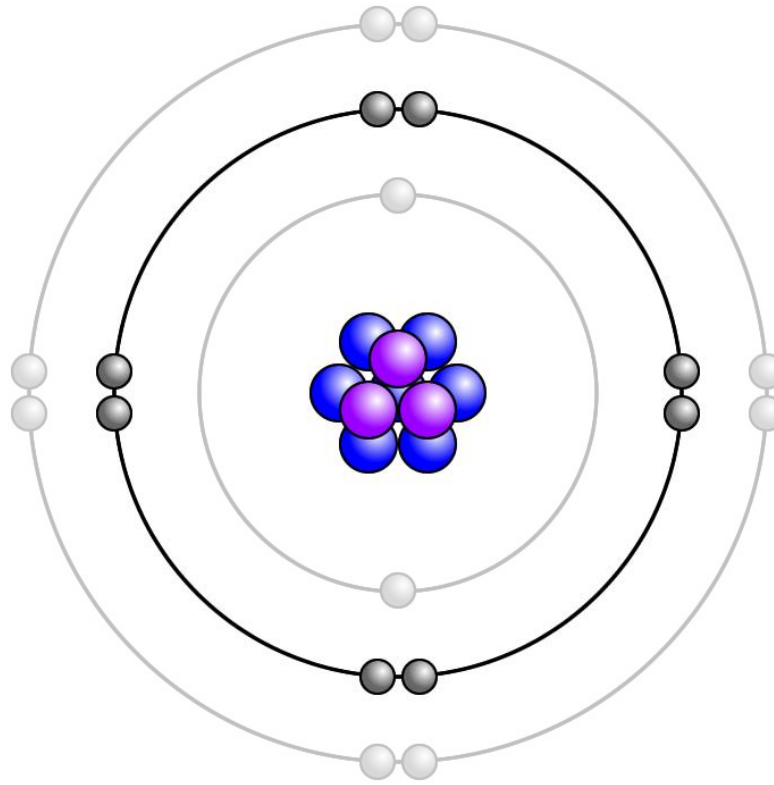
- Relative mass of an **electron** = $\frac{1}{1836}$
- Relative charge on an **electron** = -1

Introduction to Atomic Structure



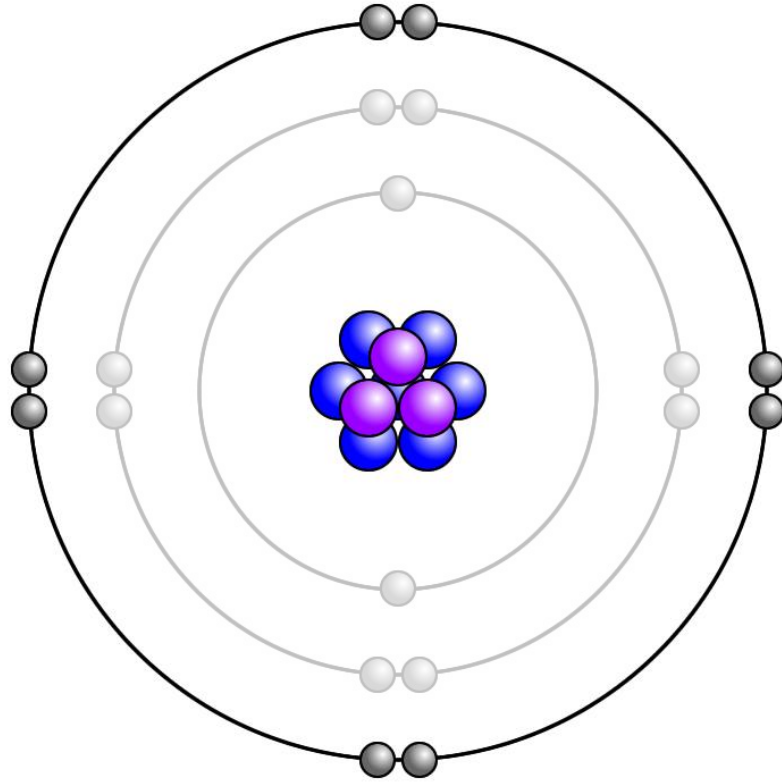
- The *inner electron shell* can hold a maximum number of **2** electrons.
- The *second electron shell* can hold a maximum number of **8** electrons.
- The *third electron shell* can hold a maximum number of **8** electrons.

Introduction to Atomic Structure



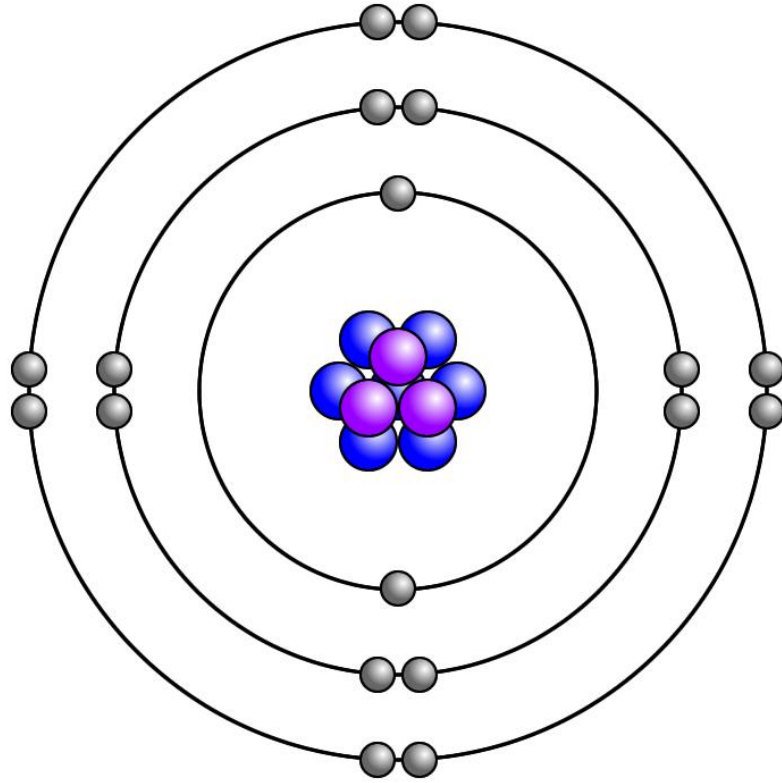
- The *inner electron shell* can hold a maximum number of **2** electrons.
- The *second electron shell* can hold a maximum number of **8** electrons.
- The *third electron shell* can hold a maximum number of **8** electrons.

Introduction to Atomic Structure



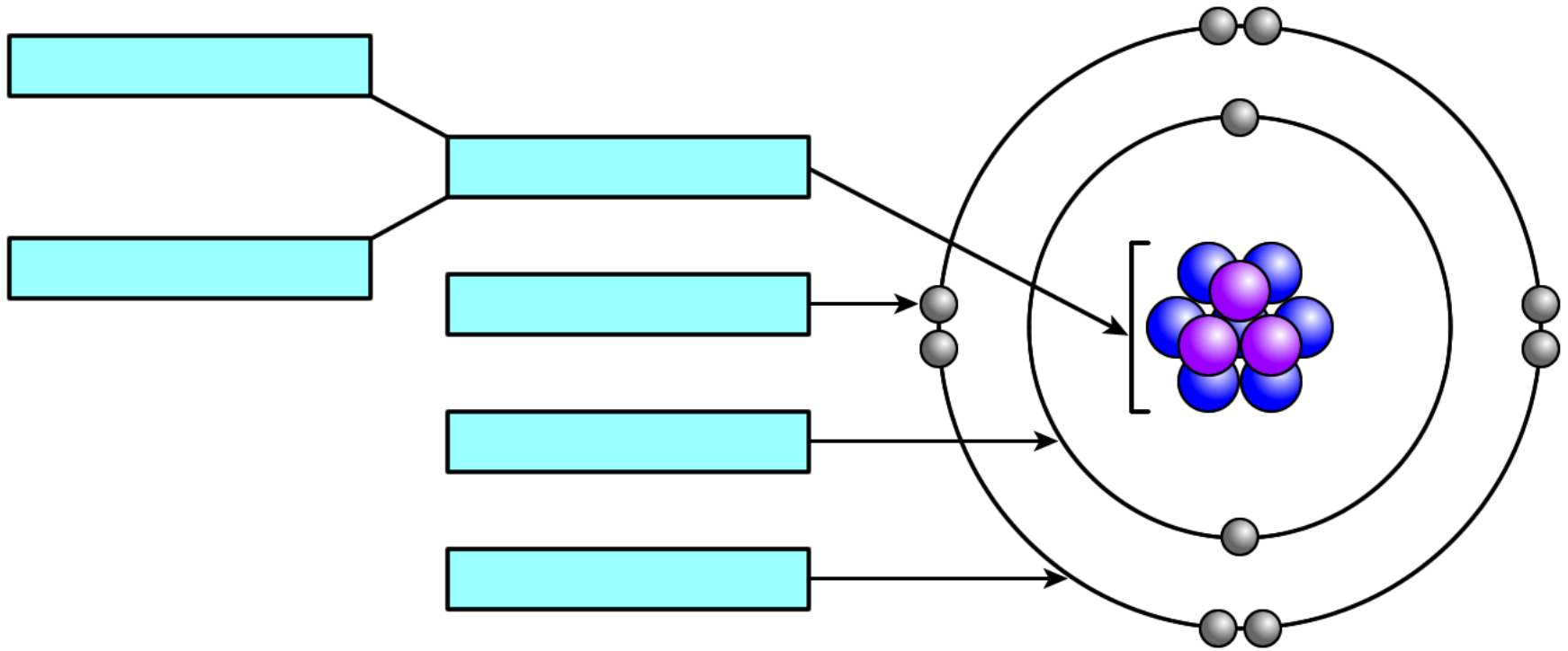
- The *inner electron shell* can hold a maximum number of **2** electrons.
- The *second electron shell* can hold a maximum number of **8** electrons.
- The *third electron shell* can hold a maximum number of **8** electrons.

Introduction to Atomic Structure

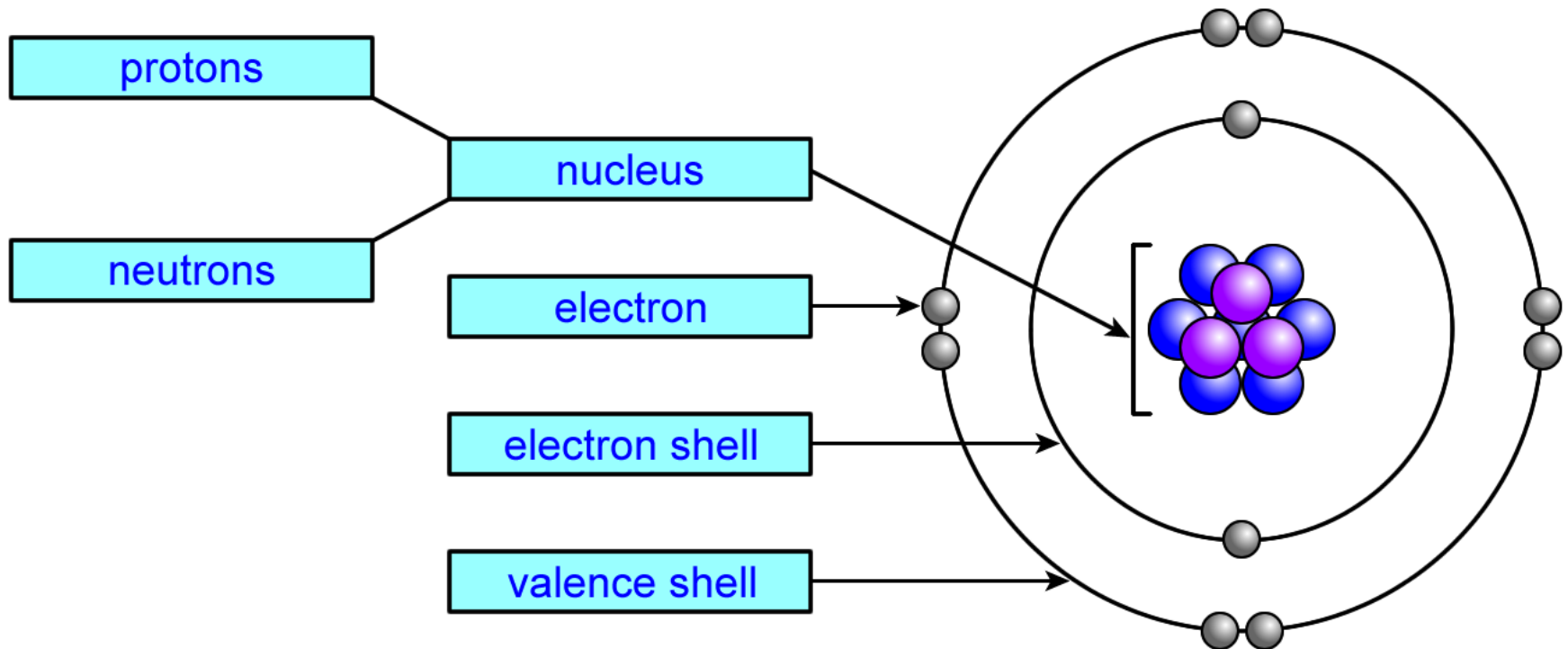


- Atoms are *electrically neutral* because the number of electrons (which carry a charge of -1) orbiting the nucleus equals the number of protons (which carry a charge of $+1$) located in the nucleus.

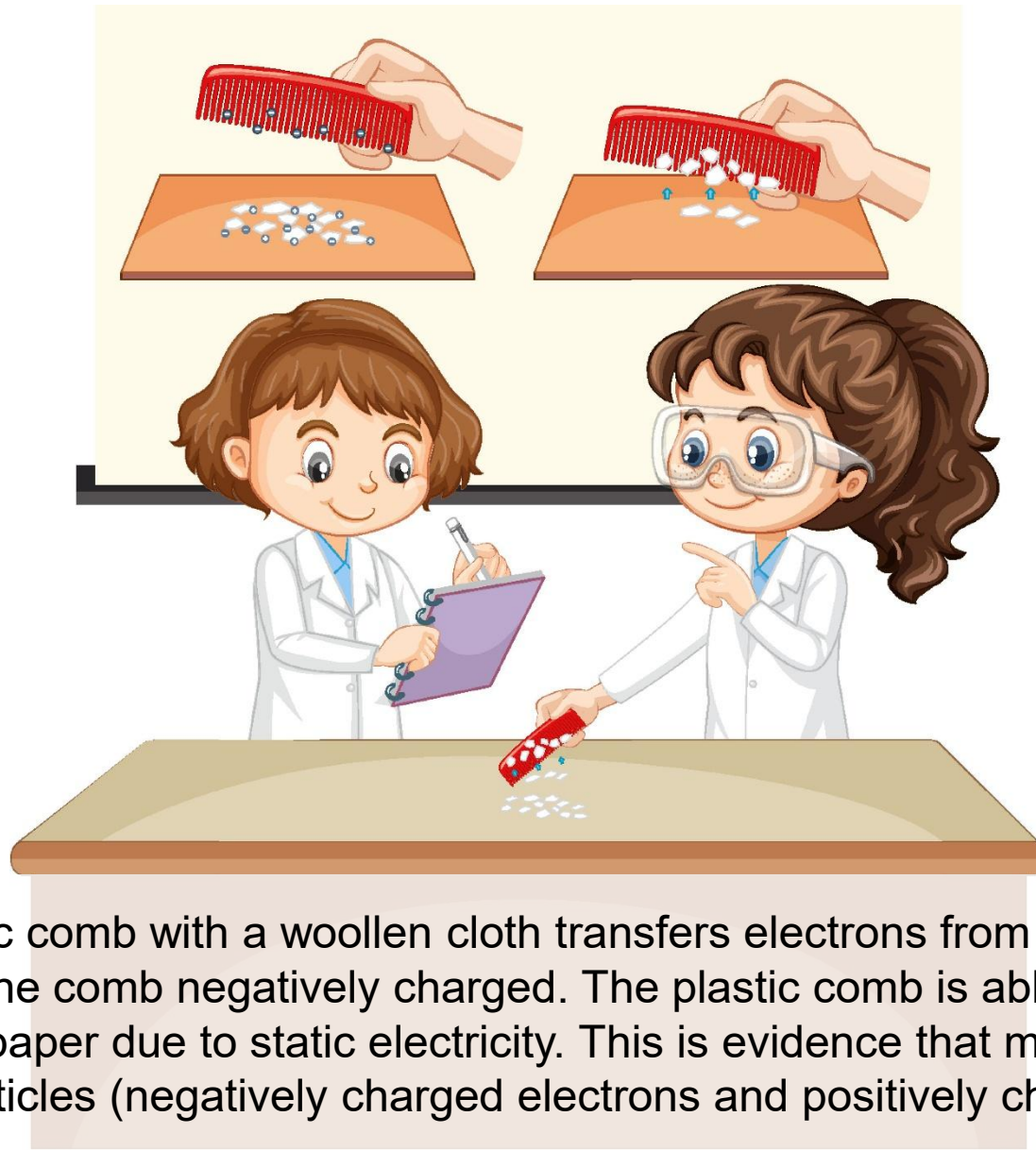
Introduction to Atomic Structure



Introduction to Atomic Structure



Introduction to Atomic Structure



Rubbing a plastic comb with a woollen cloth transfers electrons from the cloth onto the comb, making the comb negatively charged. The plastic comb is able to attract small pieces of tissue paper due to static electricity. This is evidence that matter is composed of charged particles (negatively charged electrons and positively charged protons).

Introduction to Atomic Structure

Particle	Particle's Location in the Atom	Particle's Relative Charge	Particle's Relative Mass*
Proton			
Neutron			
Electron			

Introduction to Atomic Structure

Particle	Particle's Location in the Atom	Particle's Relative Charge	Particle's Relative Mass*
Proton	Nucleus		
Neutron	Nucleus		
Electron	Orbits the Nucleus		

Introduction to Atomic Structure

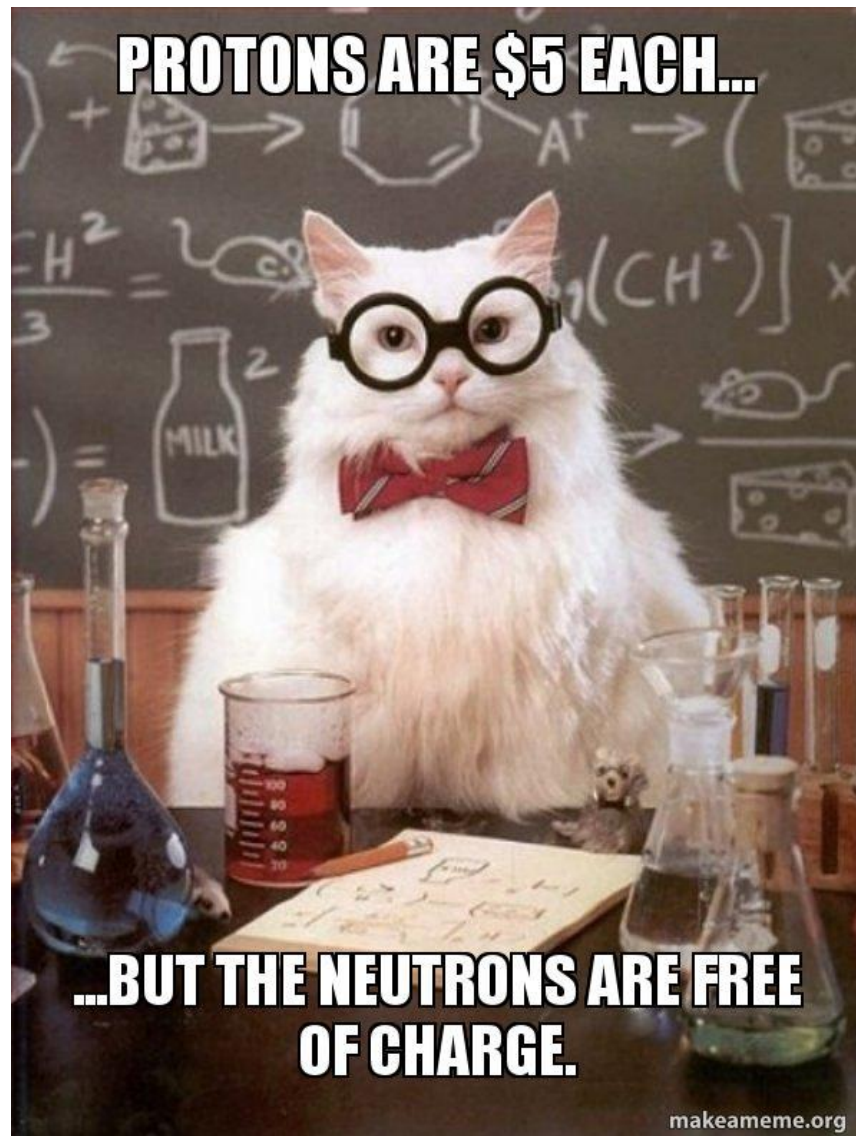
Particle	Particle's Location in the Atom	Particle's Relative Charge	Particle's Relative Mass*
Proton	Nucleus	+1	
Neutron	Nucleus	0	
Electron	Orbits the Nucleus	-1	

Introduction to Atomic Structure

Particle	Particle's Location in the Atom	Particle's Relative Charge	Particle's Relative Mass*
Proton	Nucleus	+1	1
Neutron	Nucleus	0	1
Electron	Orbits the Nucleus	-1	$\frac{1}{1836}$

*The mass of the proton, neutron and electron are sometimes given the units *a.m.u.*
a.m.u. is an abbreviation for *atomic mass unit*.

Introduction to Atomic Structure



Introduction to Atomic Structure



- Build an Atom PHET Simulation

https://phet.colorado.edu/sims/html/build-an-atom/latest/build-an-atom_en.html

Introduction to Atomic Structure

- Mass (Nucleon) Number

number of protons + neutrons

Nuclide
Notation

↓
7

3
↑

Li

- Symbol of the element
Li = *lithium*

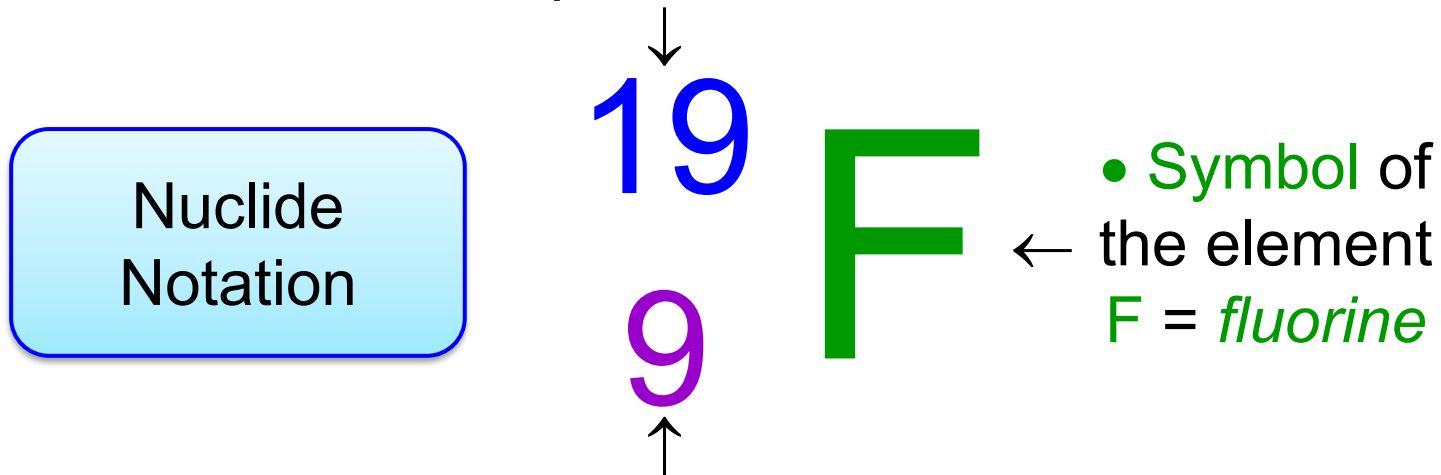
- Atomic (Proton) Number

number of protons

- number of *neutrons* = mass number – atomic number
number of neutrons = $7 - 3 = 4$
- number of *electrons* = number of protons = atomic number
number of electrons = **3**

Introduction to Atomic Structure

- Mass (Nucleon) Number
number of protons + neutrons

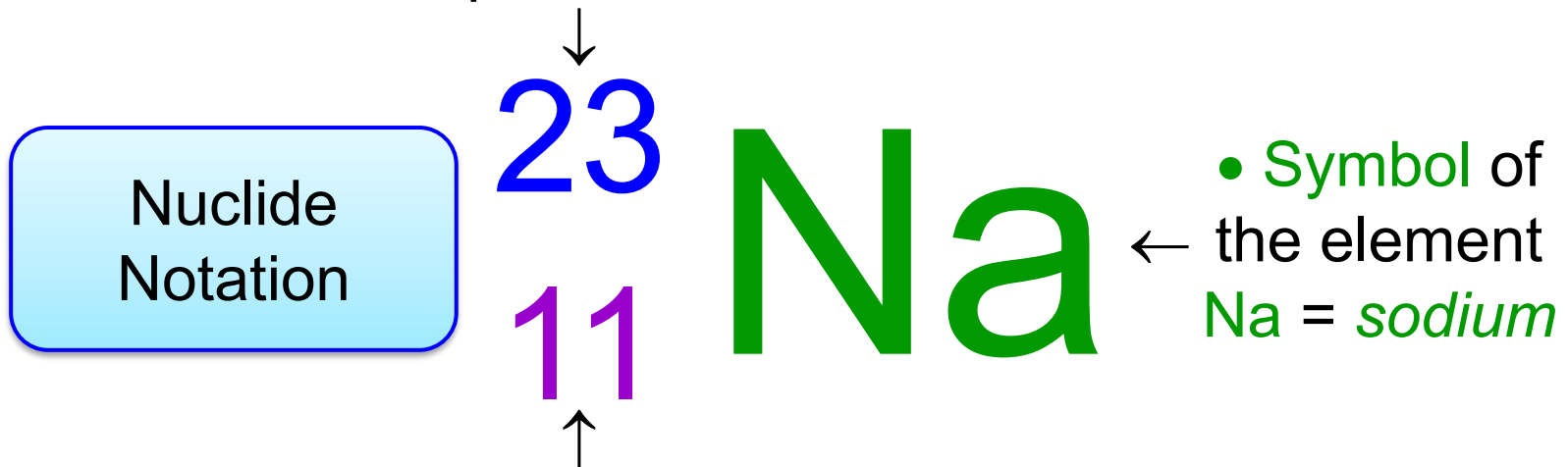


- Atomic (Proton) Number
number of protons

- number of *neutrons* = mass number – atomic number
number of neutrons = $19 - 9 = 10$
- number of *electrons* = number of protons = atomic number
number of electrons = **9**

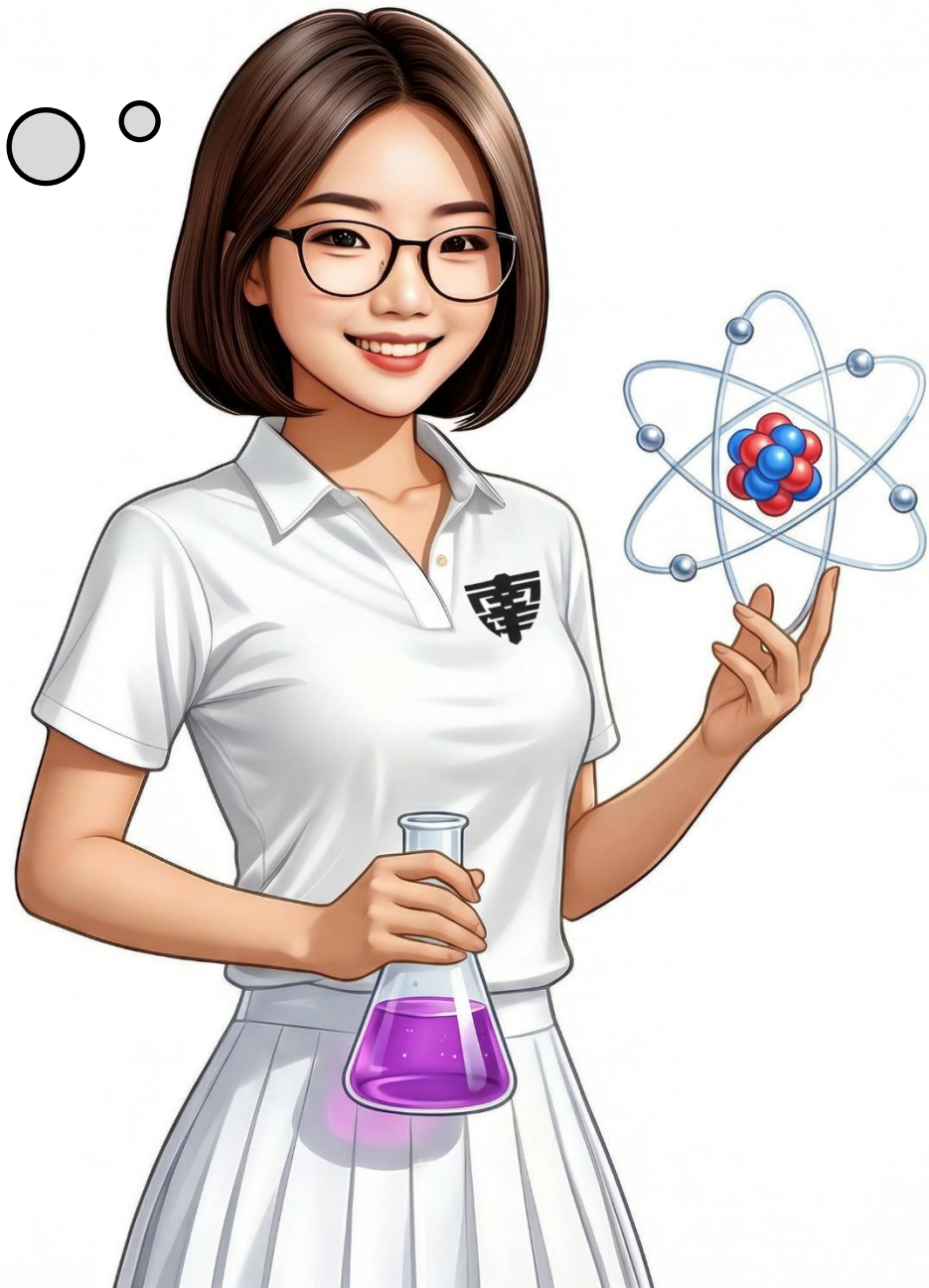
Introduction to Atomic Structure

- Mass (Nucleon) Number
number of protons + neutrons



- Atomic (Proton) Number
number of protons
- number of *neutrons* = mass number – atomic number
number of neutrons = 23 – 11 = 12
- number of *electrons* = number of protons = atomic number
number of electrons = 11

How is atomic structure related to an element's position in the Periodic Table?



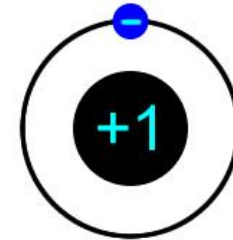
Hydrogen $\begin{matrix} 1 \\ | \\ 1 \end{matrix} \text{H}$

Atomic Number 1

Group

1 2 13 14 15 16 17 18

1	1 Hydrogen H							2 Helium He
2	3 Lithium Li	4 Beryllium Be	5 Boron B	6 Carbon C	7 Nitrogen N	8 Oxygen O	9 Fluorine F	10 Neon Ne
3	11 Sodium Na	12 Magnesium Mg	13 Aluminium Al	14 Silicon Si	15 Phosphorus P	16 Sulphur S	17 Chlorine Cl	18 Argon Ar
4	19 Potassium K	20 Calcium Ca						

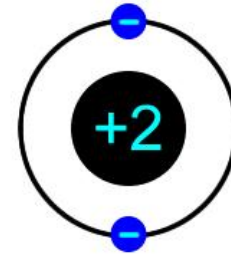


- The Atomic Number equals the number of protons = 1
 - For a neutral atom, the number electrons equals the number of protons = 1
 - The number of electron shell equals the Period number = 1
- The number of valence electrons can be derived from the Group number (1) = 1
 - The electronic configuration of hydrogen = 1
- The Mass Number equals number of protons (1) + number of neutrons (0) = 1

Helium ${}^4_2\text{He}$ Atomic Number 2

Group

	1	2	13	14	15	16	17	18
1	1 Hydrogen H							2 Helium He
2	3 Lithium Li	4 Beryllium Be	5 Boron B	6 Carbon C	7 Nitrogen N	8 Oxygen O	9 Fluorine F	10 Neon Ne
3	11 Sodium Na	12 Magnesium Mg	13 Aluminium Al	14 Silicon Si	15 Phosphorus P	16 Sulphur S	17 Chlorine Cl	18 Argon Ar
4	19 Potassium K	20 Calcium Ca						



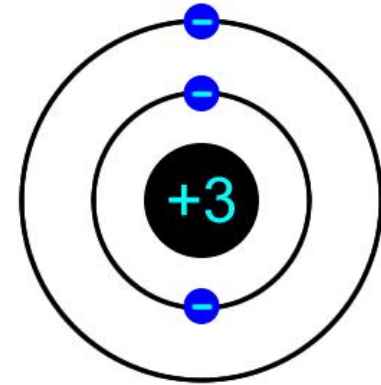
- The Atomic Number equals the number of protons = 2
 - For a neutral atom, the number electrons equals the number of protons = 2
 - The number of electron shell equals the Period number = 1
- The number of electrons in the valence electron shell of helium = 2 (complete)
 - The electronic configuration of helium = 2
- The Mass Number equals number of protons (2) + number of neutrons (2) = 4

Lithium ${}^7_3\text{Li}$ Atomic Number 3

Group

1 2 13 14 15 16 17 18

Period	1	1 Hydrogen H											2 Helium He
	2	3 Lithium Li	4 Beryllium Be	5 Boron B	6 Carbon C	7 Nitrogen N	8 Oxygen O	9 Fluorine F	10 Neon Ne				
	3	11 Sodium Na	12 Magnesium Mg	13 Aluminium Al	14 Silicon Si	15 Phosphorus P	16 Sulphur S	17 Chlorine Cl	18 Argon Ar				
	4	19 Potassium K	20 Calcium Ca										



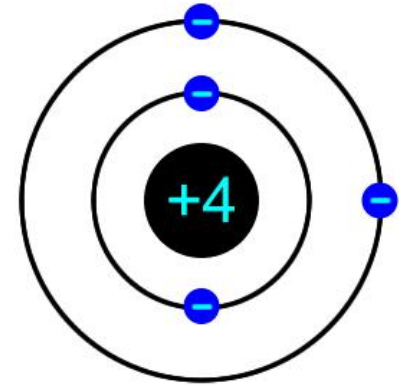
- The Atomic Number equals the number of protons = 3
 - For a neutral atom, the number electrons equals the number of protons = 3
 - The number of electron shell equals the Period number = 2
- The number of valence electrons can be derived from the Group number (1) = 1
 - The electronic configuration of lithium = 2,1
- The Mass Number equals number of protons (3) + number of neutrons (4) = 7

Beryllium ${}^9_4\text{Be}$ Atomic Number 4

Group

1 2 13 14 15 16 17 18

1	1 Hydrogen H							2 Helium He
2	3 Lithium Li	4 Beryllium Be	5 Boron B	6 Carbon C	7 Nitrogen N	8 Oxygen O	9 Fluorine F	10 Neon Ne
3	11 Sodium Na	12 Magnesium Mg	13 Aluminium Al	14 Silicon Si	15 Phosphorus P	16 Sulphur S	17 Chlorine Cl	18 Argon Ar
4	19 Potassium K	20 Calcium Ca						



- The Atomic Number equals the number of protons = 4
 - For a neutral atom, the number electrons equals the number of protons = 4
 - The number of electron shell equals the Period number = 2
- The number of valence electrons can be derived from the Group number (2) = 2
 - The electronic configuration of beryllium = 2,2
- The Mass Number equals number of protons (4) + number of neutrons (5) = 9

Boron

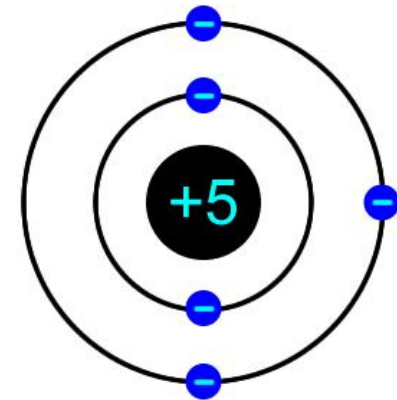


Atomic Number 5

Group

1 2 13 14 15 16 17 18

1	1 Hydrogen H							2 Helium He
2	3 Lithium Li	4 Beryllium Be	5 Boron B	6 Carbon C	7 Nitrogen N	8 Oxygen O	9 Fluorine F	10 Neon Ne
3	11 Sodium Na	12 Magnesium Mg	13 Aluminium Al	14 Silicon Si	15 Phosphorus P	16 Sulphur S	17 Chlorine Cl	18 Argon Ar
4	19 Potassium K	20 Calcium Ca						



- The Atomic Number equals the number of protons = 5
 - For a neutral atom, the number electrons equals the number of protons = 5
 - The number of electron shell equals the Period number = 2
- The number of valence electrons can be derived from the Group number (13) = 3
 - The electronic configuration of boron = 2,3
- The Mass Number equals number of protons (5) + number of neutrons (6) = 11

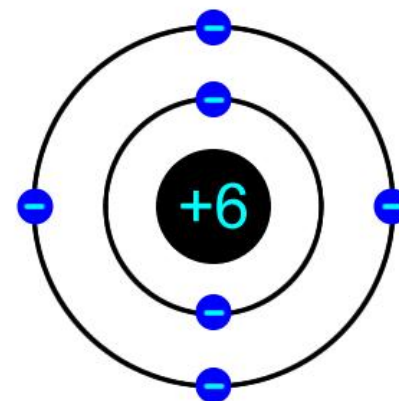
Carbon ${}^{12}_6\text{C}$

Atomic Number 6

Group

1 2 13 14 15 16 17 18

1	1 Hydrogen H							2 Helium He
2	3 Lithium Li	4 Beryllium Be	5 Boron B	6 Carbon C	7 Nitrogen N	8 Oxygen O	9 Fluorine F	10 Neon Ne
3	11 Sodium Na	12 Magnesium Mg	13 Aluminium Al	14 Silicon Si	15 Phosphorus P	16 Sulphur S	17 Chlorine Cl	18 Argon Ar
4	19 Potassium K	20 Calcium Ca						



- The Atomic Number equals the number of protons = 6
 - For a neutral atom, the number electrons equals the number of protons = 6
 - The number of electron shell equals the Period number = 2
- The number of valence electrons can be derived from the Group number (14) = 4
 - The electronic configuration of carbon = 2,4
- The Mass Number equals number of protons (6) + number of neutrons (6) = 12

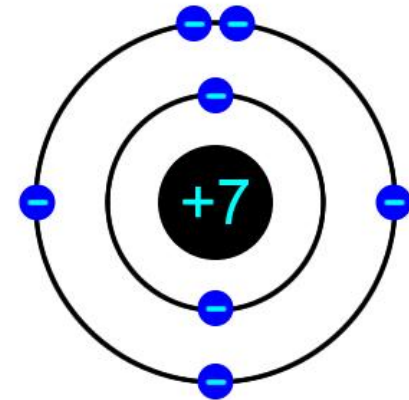
Nitrogen ${}^{14}_7\text{N}$

Atomic Number 7

Group

1 2 13 14 15 16 17 18

1	1 Hydrogen H							2 Helium He
2	3 Lithium Li	4 Beryllium Be	5 Boron B	6 Carbon C	7 Nitrogen N	8 Oxygen O	9 Fluorine F	10 Neon Ne
3	11 Sodium Na	12 Magnesium Mg	13 Aluminium Al	14 Silicon Si	15 Phosphorus P	16 Sulphur S	17 Chlorine Cl	18 Argon Ar
4	19 Potassium K	20 Calcium Ca						



- The Atomic Number equals the number of protons = 7
 - For a neutral atom, the number electrons equals the number of protons = 7
 - The number of electron shell equals the Period number = 2
- The number of valence electrons can be derived from the Group number (15) = 5
 - The electronic configuration of nitrogen = 2,5
- The Mass Number equals number of protons (7) + number of neutrons (7) = 14

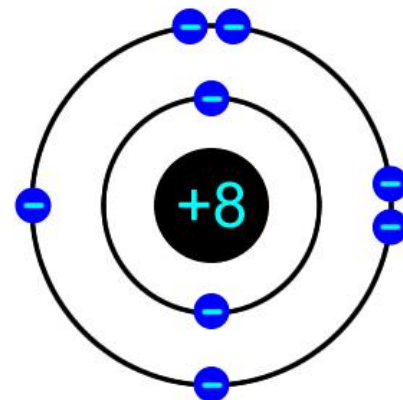
Oxygen $^{16}_8\text{O}$

Atomic Number 8

Group

1 2 13 14 15 16 17 18

1	1 Hydrogen H							2 Helium He
2	3 Lithium Li	4 Beryllium Be	5 Boron B	6 Carbon C	7 Nitrogen N	8 Oxygen O	9 Fluorine F	10 Neon Ne
3	11 Sodium Na	12 Magnesium Mg	13 Aluminium Al	14 Silicon Si	15 Phosphorus P	16 Sulphur S	17 Chlorine Cl	18 Argon Ar
4	19 Potassium K	20 Calcium Ca						



- The Atomic Number equals the number of protons = 8
 - For a neutral atom, the number electrons equals the number of protons = 8
 - The number of electron shell equals the Period number = 2
- The number of valence electrons can be derived from the Group number (16) = 6
 - The electronic configuration of oxygen = 2,6
- The Mass Number equals number of protons (8) + number of neutrons (8) = 16

Fluorine

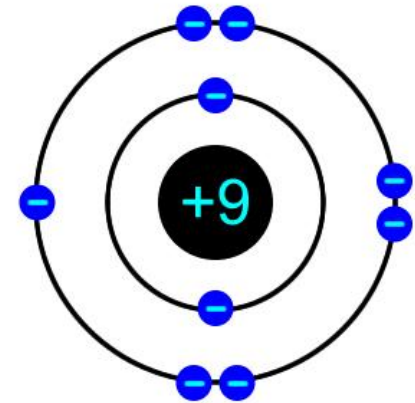


Atomic Number 9

Group

1 2 13 14 15 16 17 18

Period	1	1 Hydrogen H											2 Helium He					
	2	3 Lithium Li	4 Beryllium Be	5 Boron B	6 Carbon C	7 Nitrogen N	8 Oxygen O	9 Fluorine F	10 Neon Ne									
	3	11 Sodium Na	12 Magnesium Mg	13 Aluminium Al	14 Silicon Si	15 Phosphorus P	16 Sulphur S	17 Chlorine Cl	18 Argon Ar									
	4	19 Potassium K	20 Calcium Ca															



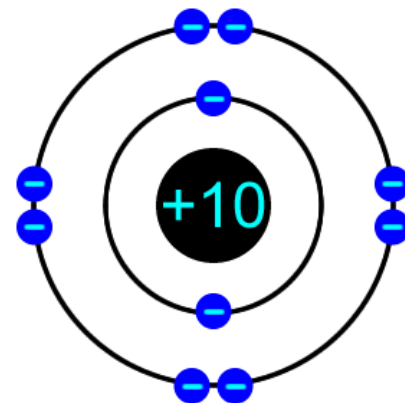
- The Atomic Number equals the number of protons = 9
 - For a neutral atom, the number electrons equals the number of protons = 9
 - The number of electron shell equals the Period number = 2
- The number of valence electrons can be derived from the Group number (17) = 7
 - The electronic configuration of fluorine = 2,7
- The Mass Number equals number of protons (9) + number of neutrons (10) = 19

Neon ${}_{10}^{20}\text{Ne}$ Atomic Number 10

Group

1 2 13 14 15 16 17 18

Period	1	1 Hydrogen H							2 Helium He
	2	3 Lithium Li	4 Beryllium Be	5 Boron B	6 Carbon C	7 Nitrogen N	8 Oxygen O	9 Fluorine F	10 Neon Ne
	3	11 Sodium Na	12 Magnesium Mg	13 Aluminium Al	14 Silicon Si	15 Phosphorus P	16 Sulphur S	17 Chlorine Cl	18 Argon Ar
	4	19 Potassium K	20 Calcium Ca						



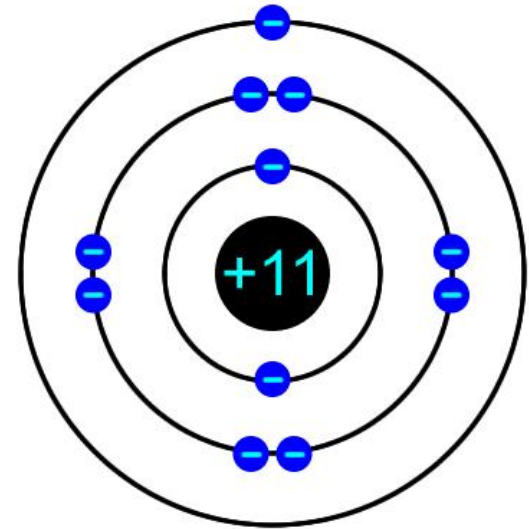
- The Atomic Number equals the number of protons = 10
- For a neutral atom, the number electrons equals the number of protons = 10
- The number of electron shell equals the Period number = 2
- The number of valence electrons can be derived from the Group number (18) = 8 (complete)
 - The electronic configuration of neon = 2,8
- The Mass Number equals number of protons (10) + number of neutrons (10) = 20

Sodium ${}_{11}^{23}\text{Na}$ Atomic Number 11

Group

1 2 13 14 15 16 17 18

Period	1	1 Hydrogen H							2 Helium He
	2	3 Lithium Li	4 Beryllium Be	5 Boron B	6 Carbon C	7 Nitrogen N	8 Oxygen O	9 Fluorine F	10 Neon Ne
	3	11 Sodium Na	12 Magnesium Mg	13 Aluminium Al	14 Silicon Si	15 Phosphorus P	16 Sulphur S	17 Chlorine Cl	18 Argon Ar
	4	19 Potassium K	20 Calcium Ca						



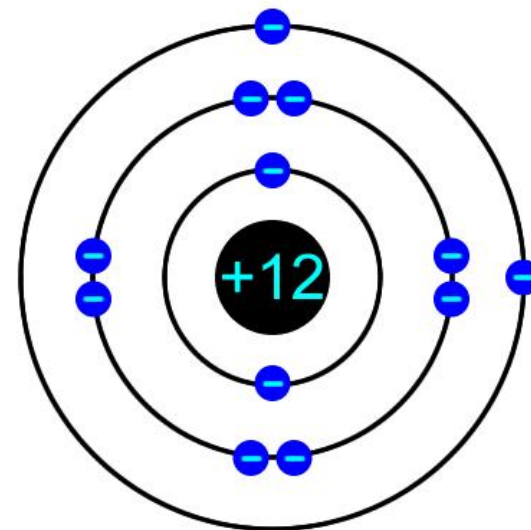
- The Atomic Number equals the number of protons = 11
- For a neutral atom, the number electrons equals the number of protons = 11
- The number of electron shell equals the Period number = 3
- The number of valence electrons can be derived from the Group number (1) = 1
 - The electronic configuration of sodium = 2,8,1
- The Mass Number equals number of protons (11) + number of neutrons (12) = 23

Magnesium ${}_{12}^{24}\text{Mg}$ Atomic Number 12

Group

1 2 13 14 15 16 17 18

1	1 Hydrogen H							2 Helium He
2	3 Lithium Li	4 Beryllium Be	5 Boron B	6 Carbon C	7 Nitrogen N	8 Oxygen O	9 Fluorine F	10 Neon Ne
3	11 Sodium Na	12 Magnesium Mg	13 Aluminium Al	14 Silicon Si	15 Phosphorus P	16 Sulphur S	17 Chlorine Cl	18 Argon Ar
4	19 Potassium K	20 Calcium Ca						



- The Atomic Number equals the number of protons = 12
- For a neutral atom, the number electrons equals the number of protons = 12
- The number of electron shell equals the Period number = 3
- The number of valence electrons can be derived from the Group number (2) = 2
 - The electronic configuration of magnesium = 2,8,2
- The Mass Number equals number of protons (12) + number of neutrons (12) = 24

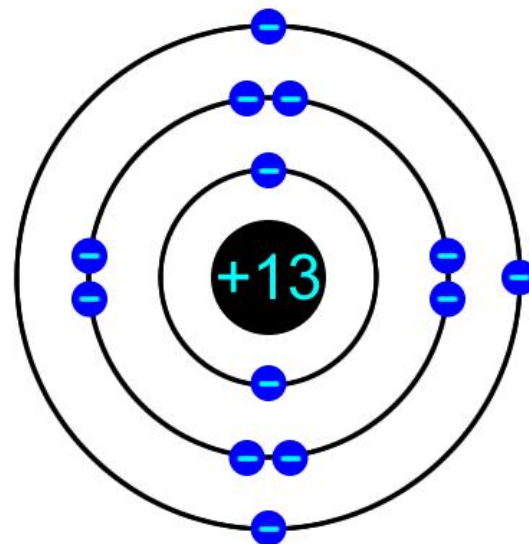
Aluminium ${}_{13}^{27}\text{Al}$

Atomic Number 13

Group

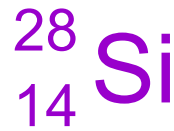
1 2 13 14 15 16 17 18

1	Hydrogen H							2 Helium He
2	3 Lithium Li	4 Beryllium Be	5 Boron B	6 Carbon C	7 Nitrogen N	8 Oxygen O	9 Fluorine F	10 Neon Ne
3	11 Sodium Na	12 Magnesium Mg	13 Aluminium Al	14 Silicon Si	15 Phosphorus P	16 Sulphur S	17 Chlorine Cl	18 Argon Ar
4	19 Potassium K	20 Calcium Ca						



- The Atomic Number equals the number of protons = 13
- For a neutral atom, the number electrons equals the number of protons = 13
- The number of electron shell equals the Period number = 3
- The number of valence electrons can be derived from the Group number (13) = 3
 - The electronic configuration of aluminium = 2,8,3
- The Mass Number equals number of protons (13) + number of neutrons (14) = 27

Silicon

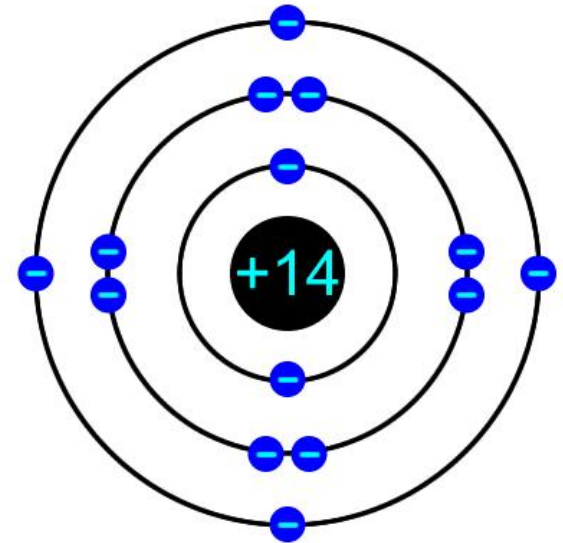


Atomic Number 14

Group

1 2 13 14 15 16 17 18

Period	1	1 Hydrogen H											2 Helium He					
	2	3 Lithium Li	4 Beryllium Be	5 Boron B	6 Carbon C	7 Nitrogen N	8 Oxygen O	9 Fluorine F	10 Neon Ne									
	3	11 Sodium Na	12 Magnesium Mg	13 Aluminium Al	14 Silicon Si	15 Phosphorus P	16 Sulphur S	17 Chlorine Cl	18 Argon Ar									
	4	19 Potassium K	20 Calcium Ca															



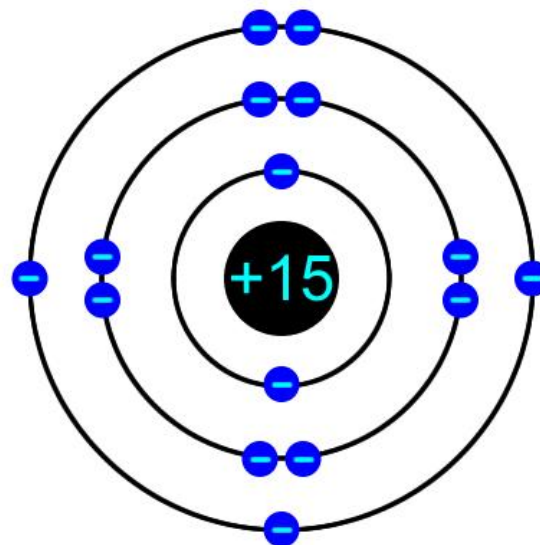
- The Atomic Number equals the number of protons = 14
- For a neutral atom, the number electrons equals the number of protons = 14
- The number of electron shell equals the Period number = 3
- The number of valence electrons can be derived from the Group number (14) = 4
 - The electronic configuration of silicon = 2,8,4
- The Mass Number equals number of protons (14) + number of neutrons (14) = 28

Phosphorus ${}_{15}^{31}\text{P}$

Atomic Number 15

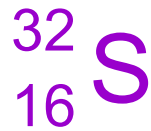
Group

	1	2	13	14	15	16	17	18
1	1 Hydrogen H							2 Helium He
2	3 Lithium Li	4 Beryllium Be	5 Boron B	6 Carbon C	7 Nitrogen N	8 Oxygen O	9 Fluorine F	10 Neon Ne
3	11 Sodium Na	12 Magnesium Mg	13 Aluminium Al	14 Silicon Si	15 Phosphorus P	16 Sulphur S	17 Chlorine Cl	18 Argon Ar
4	19 Potassium K	20 Calcium Ca						



- The Atomic Number equals the number of protons = 15
- For a neutral atom, the number electrons equals the number of protons = 15
- The number of electron shell equals the Period number = 3
- The number of valence electrons can be derived from the Group number (15) = 5
 - The electronic configuration of phosphorus = 2,8,5
- The Mass Number equals number of protons (15) + number of neutrons (16) = 31

Sulfur

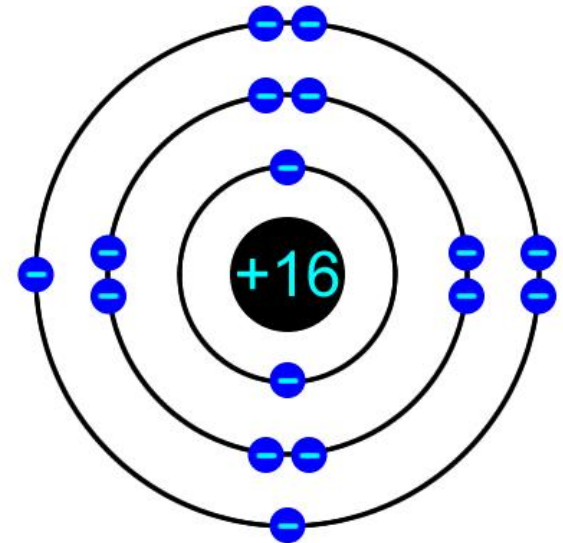


Atomic Number 16

Group

1 2 13 14 15 16 17 18

Period	1	1 Hydrogen H											2 Helium He				
	2	3 Lithium Li	4 Beryllium Be	5 Boron B	6 Carbon C	7 Nitrogen N	8 Oxygen O	9 Fluorine F	10 Neon Ne								
	3	11 Sodium Na	12 Magnesium Mg	13 Aluminium Al	14 Silicon Si	15 Phosphorus P	16 Sulphur S	17 Chlorine Cl	18 Argon Ar								
	4	19 Potassium K	20 Calcium Ca														



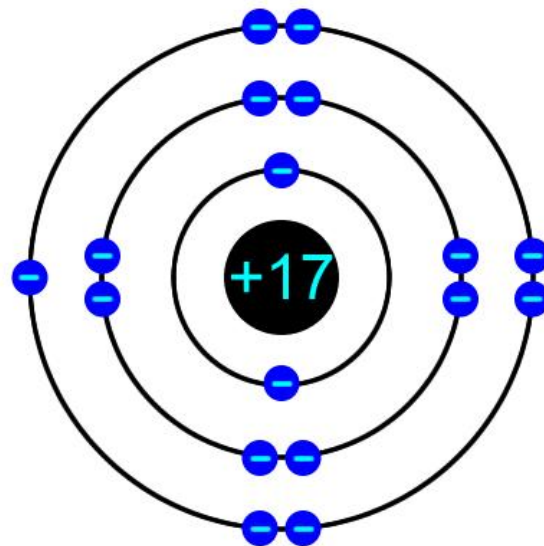
- The Atomic Number equals the number of protons = 16
- For a neutral atom, the number electrons equals the number of protons = 16
- The number of electron shell equals the Period number = 3
- The number of valence electrons can be derived from the Group number (16) = 6
 - The electronic configuration of sulfur = 2,8,6
- The Mass Number equals number of protons (16) + number of neutrons (16) = 32

Chlorine ${}_{17}^{35}\text{Cl}$ Atomic Number 17

Group

1 2 13 14 15 16 17 18

Period	1	1 Hydrogen H											2 Helium He
	2	3 Lithium Li	4 Beryllium Be	5 Boron B	6 Carbon C	7 Nitrogen N	8 Oxygen O	9 Fluorine F	10 Neon Ne				
	3	11 Sodium Na	12 Magnesium Mg	13 Aluminium Al	14 Silicon Si	15 Phosphorus P	16 Sulphur S	17 Chlorine Cl	18 Argon Ar				
	4	19 Potassium K	20 Calcium Ca										



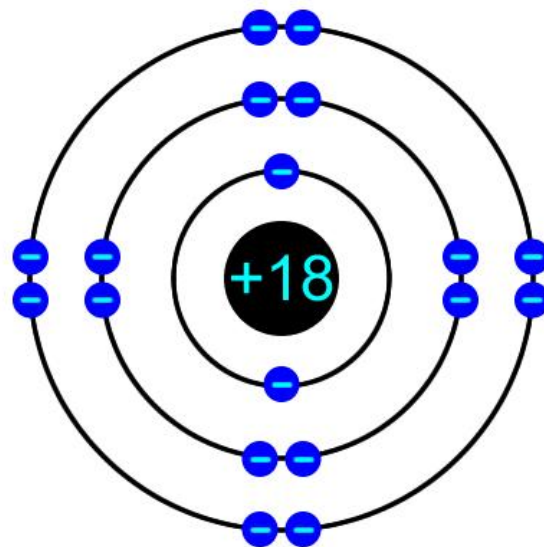
- The Atomic Number equals the number of protons = 17
- For a neutral atom, the number electrons equals the number of protons = 17
- The number of electron shell equals the Period number = 3
- The number of valence electrons can be derived from the Group number (17) = 7
 - The electronic configuration of chlorine = 2,8,7
- The Mass Number equals number of protons (17) + number of neutrons (18) = 35

Argon ${}_{18}^{40}\text{Ar}$ Atomic Number 18

Group

1 2 13 14 15 16 17 18

Period	1	1 Hydrogen H											2 Helium He
	2	3 Lithium Li	4 Beryllium Be	5 Boron B	6 Carbon C	7 Nitrogen N	8 Oxygen O	9 Fluorine F	10 Neon Ne				
	3	11 Sodium Na	12 Magnesium Mg	13 Aluminium Al	14 Silicon Si	15 Phosphorus P	16 Sulphur S	17 Chlorine Cl	18 Argon Ar				
	4	19 Potassium K	20 Calcium Ca										



- The Atomic Number equals the number of protons = 18
- For a neutral atom, the number electrons equals the number of protons = 18
- The number of electron shell equals the Period number = 3
- The number of valence electrons can be derived from the Group number (18) = 8 (complete)
 - The electronic configuration of argon = 2,8,8
- The Mass Number equals number of protons (18) + number of neutrons (22) = 40

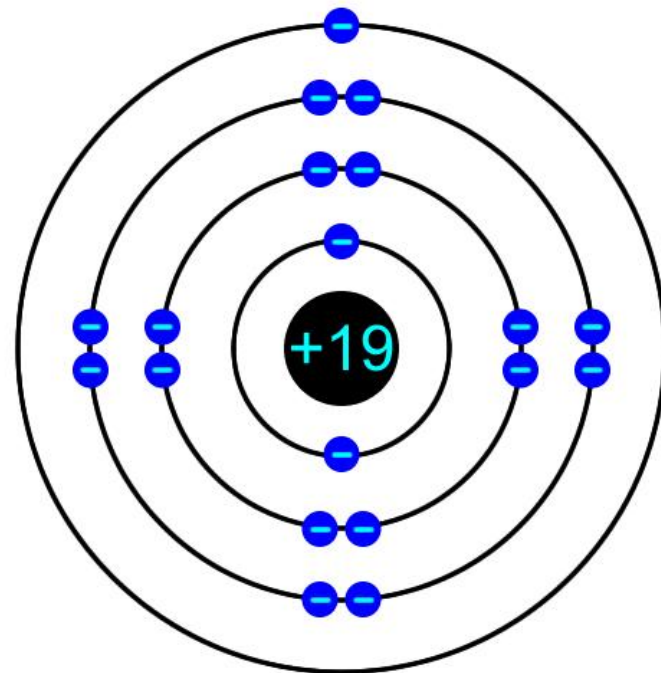
Potassium ${}_{19}^{39}\text{K}$

Atomic Number 19

Group

1 2 13 14 15 16 17 18

1	Hydrogen H							2 Helium He
2	3 Lithium Li	4 Beryllium Be	5 Boron B	6 Carbon C	7 Nitrogen N	8 Oxygen O	9 Fluorine F	10 Neon Ne
3	11 Sodium Na	12 Magnesium Mg	13 Aluminium Al	14 Silicon Si	15 Phosphorus P	16 Sulphur S	17 Chlorine Cl	18 Argon Ar
4	19 Potassium K	20 Calcium Ca						



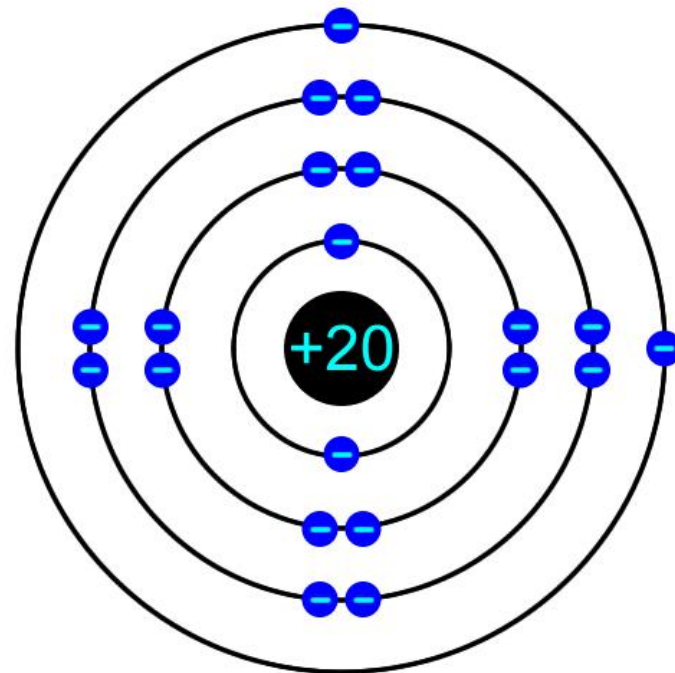
- The Atomic Number equals the number of protons = 19
- For a neutral atom, the number electrons equals the number of protons = 19
- The number of electron shell equals the Period number = 4
- The number of valence electrons can be derived from the Group number (1) = 1
 - The electronic configuration of potassium = 2,8,8,1
- The Mass Number equals number of protons (19) + number of neutrons (20) = 39

Calcium ${}_{20}^{40}\text{Ca}$ Atomic Number 20

Group

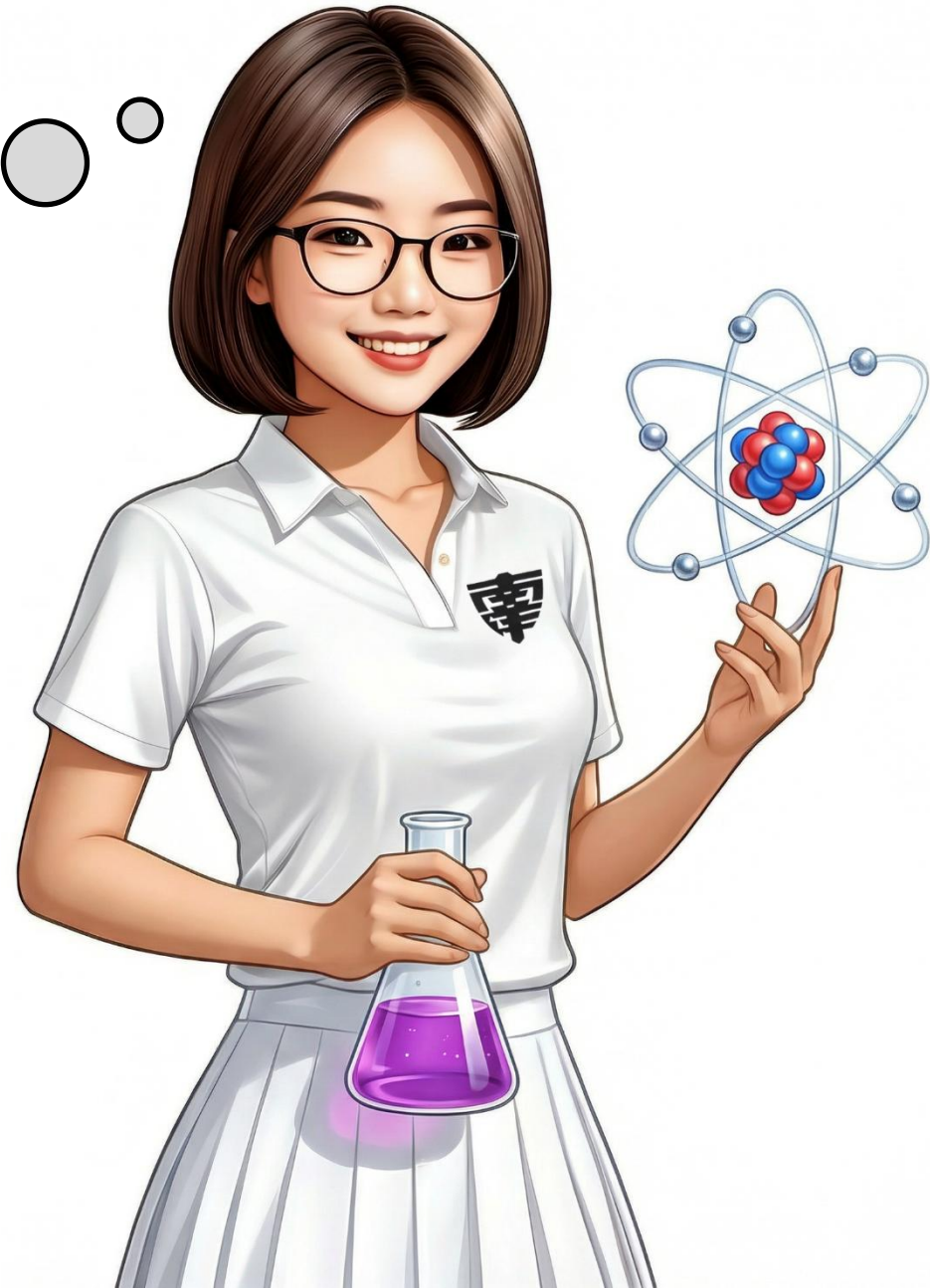
1 2 13 14 15 16 17 18

Period	1	1 Hydrogen H															2 Helium He		
	2	3 Lithium Li	4 Beryllium Be	5 Boron B	6 Carbon C	7 Nitrogen N	8 Oxygen O	9 Fluorine F	10 Neon Ne										
	3	11 Sodium Na	12 Magnesium Mg	13 Aluminium Al	14 Silicon Si	15 Phosphorus P	16 Sulphur S	17 Chlorine Cl	18 Argon Ar										
	4	19 Potassium K	20 Calcium Ca																

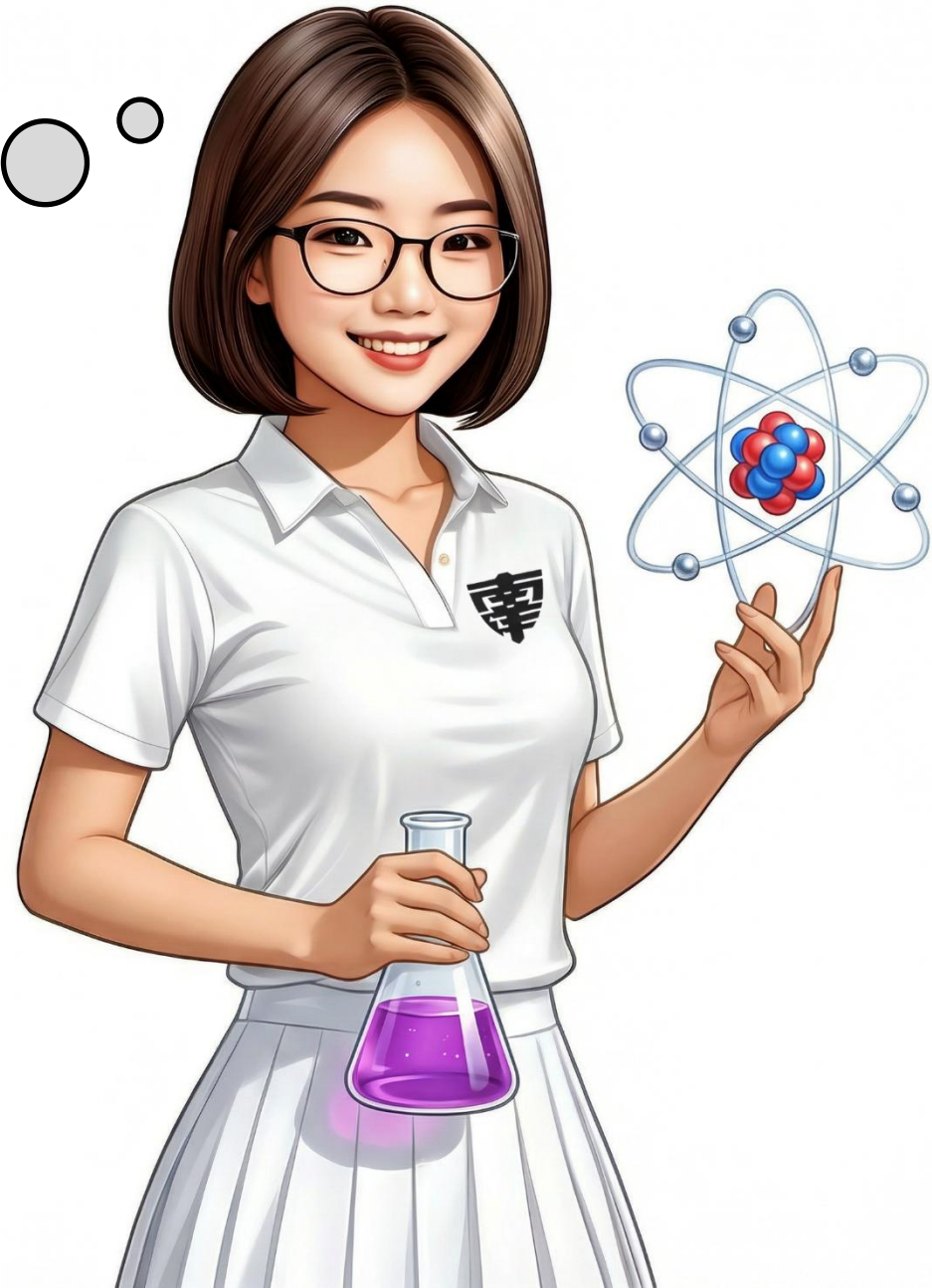


- The Atomic Number equals the number of protons = 20
- For a neutral atom, the number electrons equals the number of protons = 20
- The number of electron shell equals the Period number = 4
- The number of valence electrons can be derived from the Group number (2) = 2
 - The electronic configuration of calcium = 2,8,8,2
- The Mass Number equals number of protons (20) + number of neutrons (20) = 40

So the number
of **protons** in
the nucleus of
the atom
defines which
chemical
element the
atom belongs to!

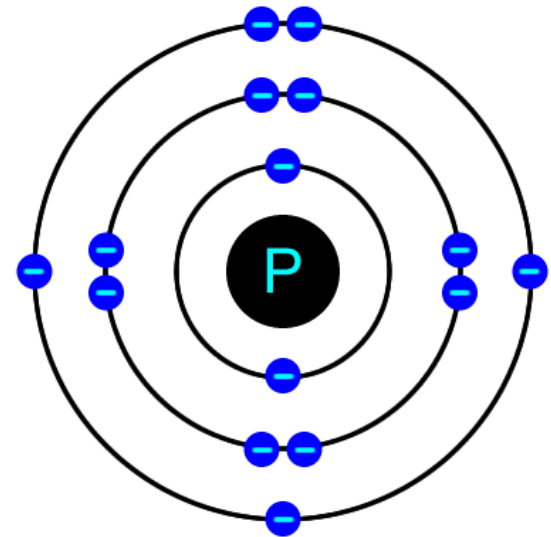
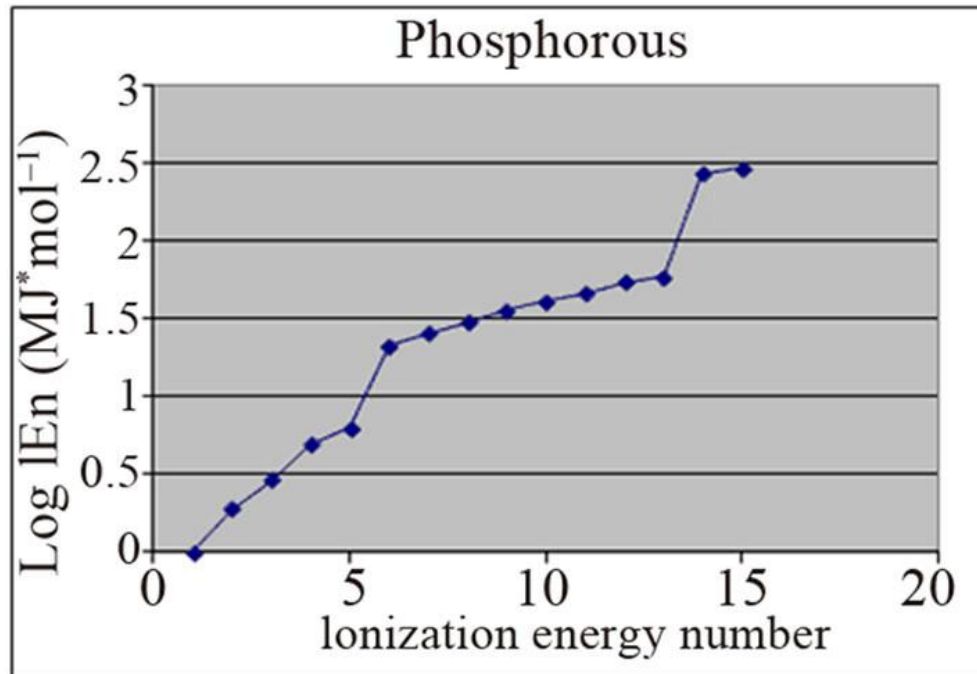


What evidence
is there to
support this
model of atomic
structure?



Introduction to Atomic Structure

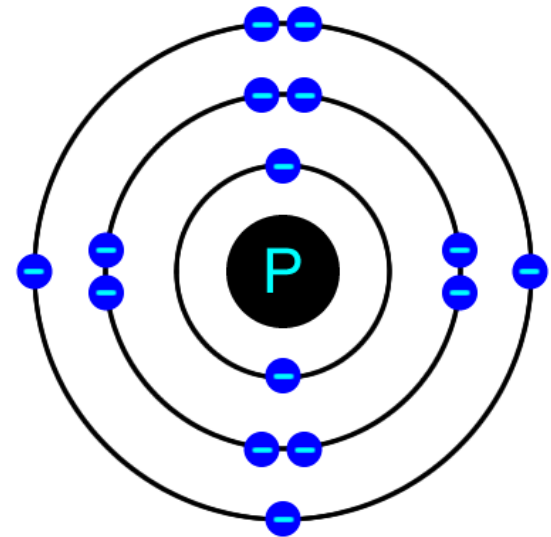
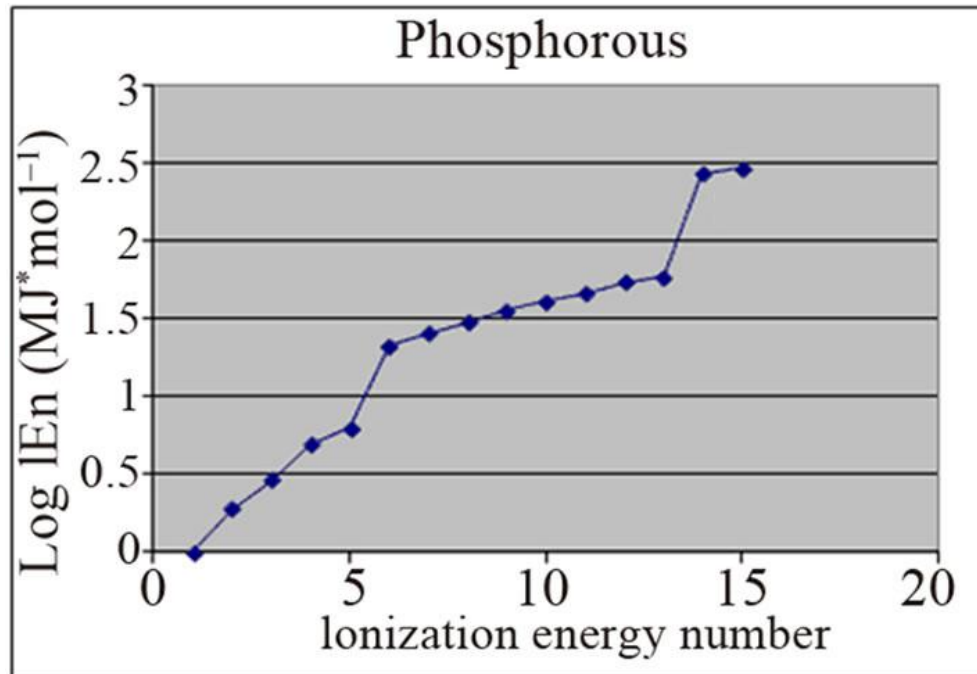
Evidence of Atomic Structure – Successive Ionisation Energies



- Energy is *required* to *remove* an electron from an atom. The energy is used to overcome the *electrostatic force of attraction* between the negatively charged electron and positively charged nucleus.

Introduction to Atomic Structure

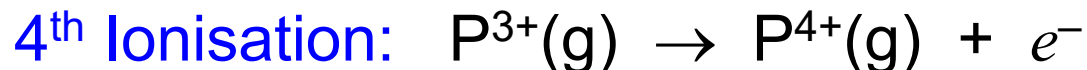
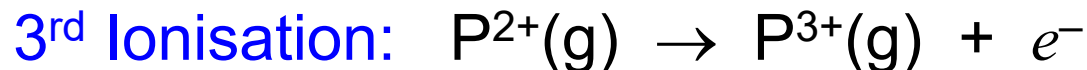
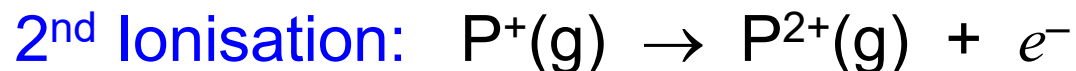
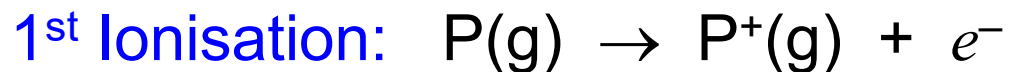
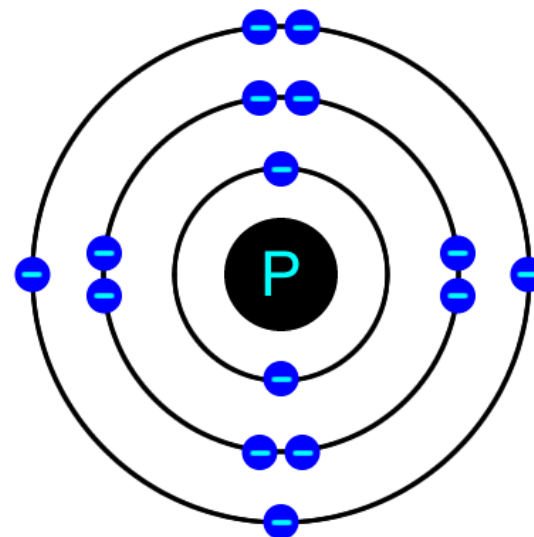
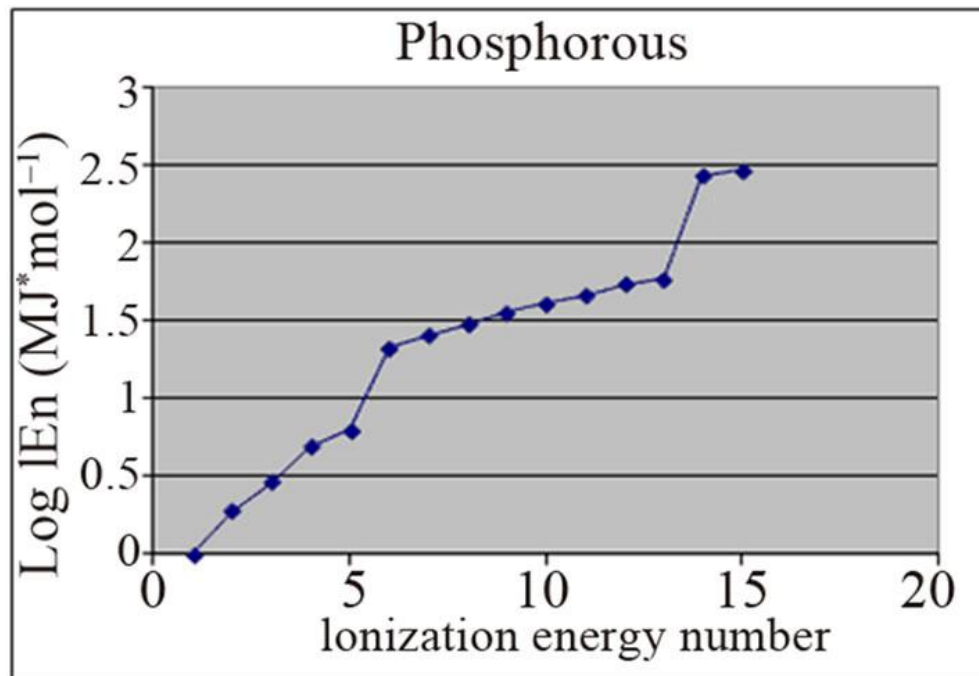
Evidence of Atomic Structure – Successive Ionisation Energies



- If *all* of the electrons are removed from a single atom of *phosphorus*, one after the other, and the energy required to remove each electron is plotted on a graph, we obtain the result shown above.

Introduction to Atomic Structure

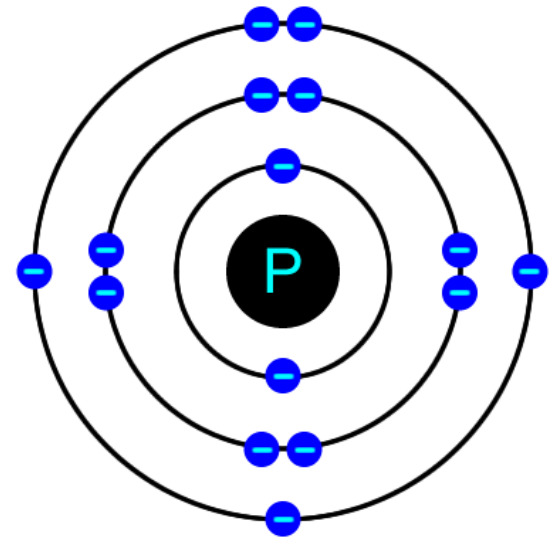
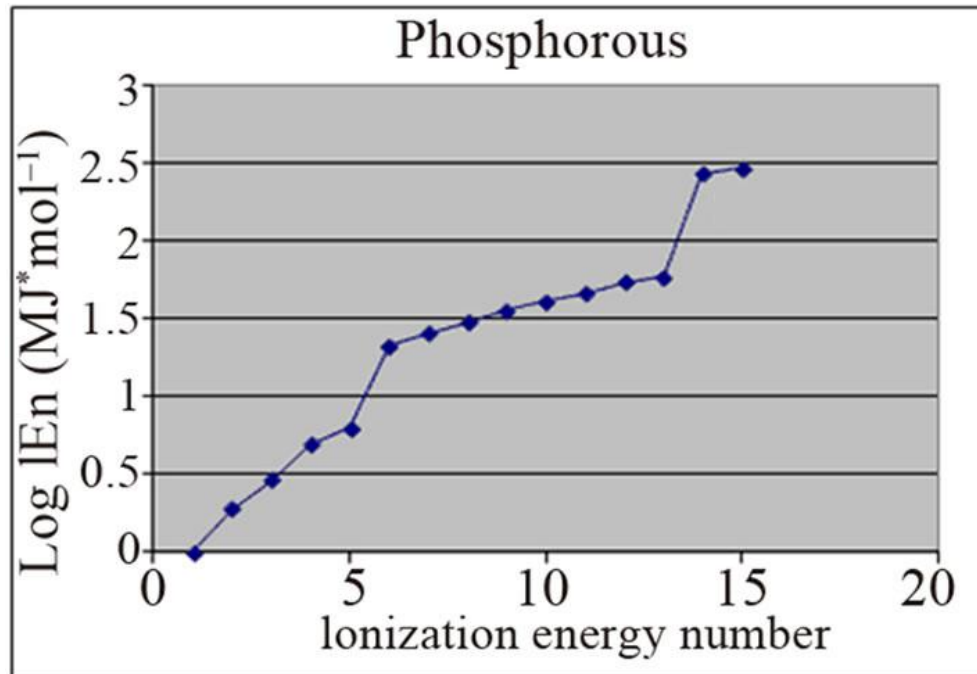
Evidence of Atomic Structure – Successive Ionisation Energies



...and so on...

Introduction to Atomic Structure

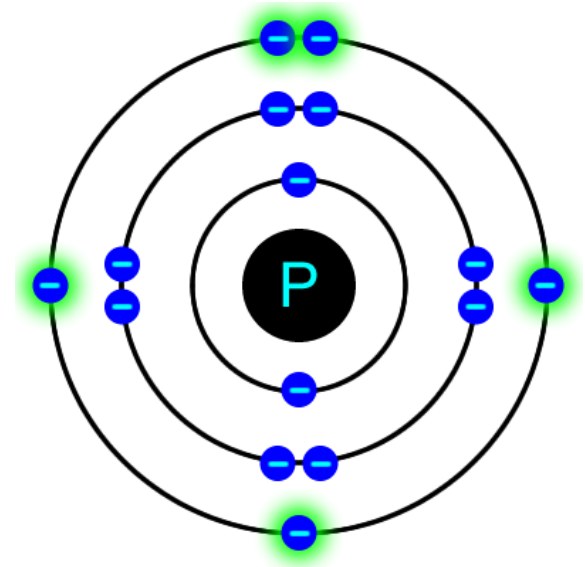
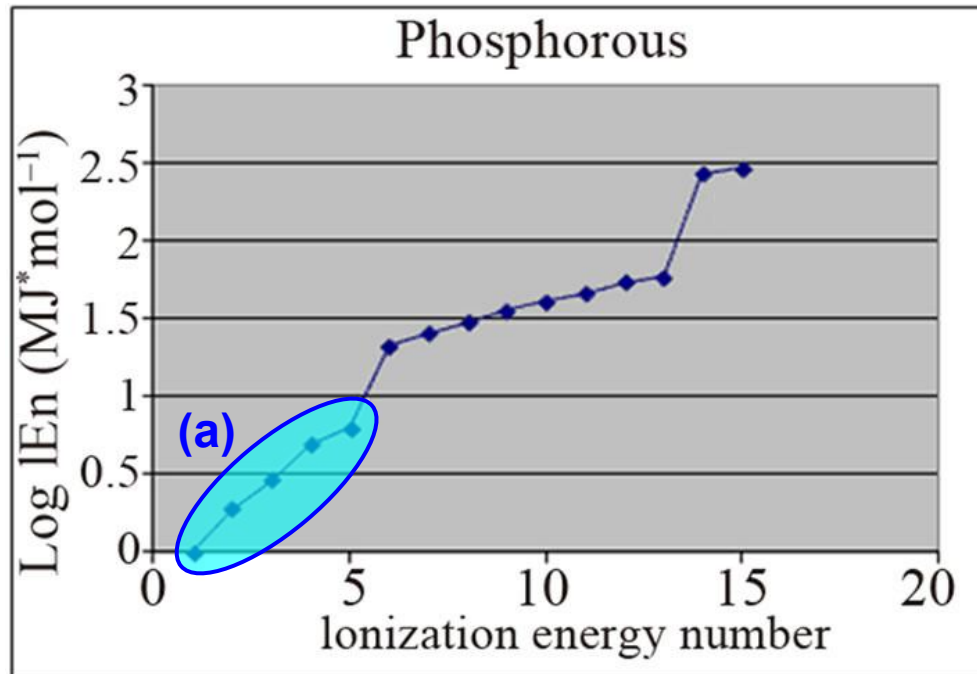
Evidence of Atomic Structure – Successive Ionisation Energies



- The graph is not linear, but can be divided into *three* distinct regions, which reflect the *three* electron shells:

Introduction to Atomic Structure

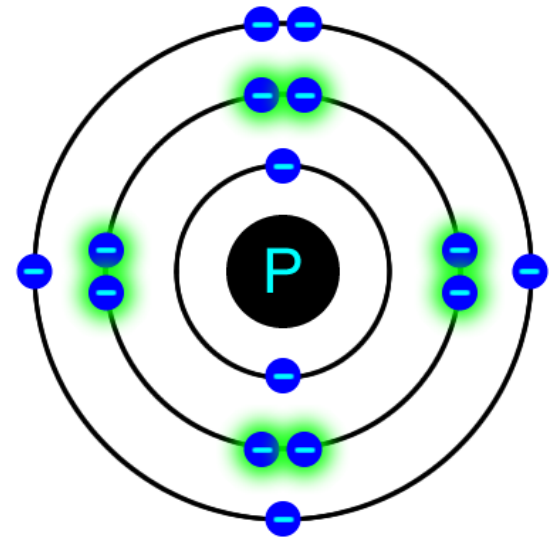
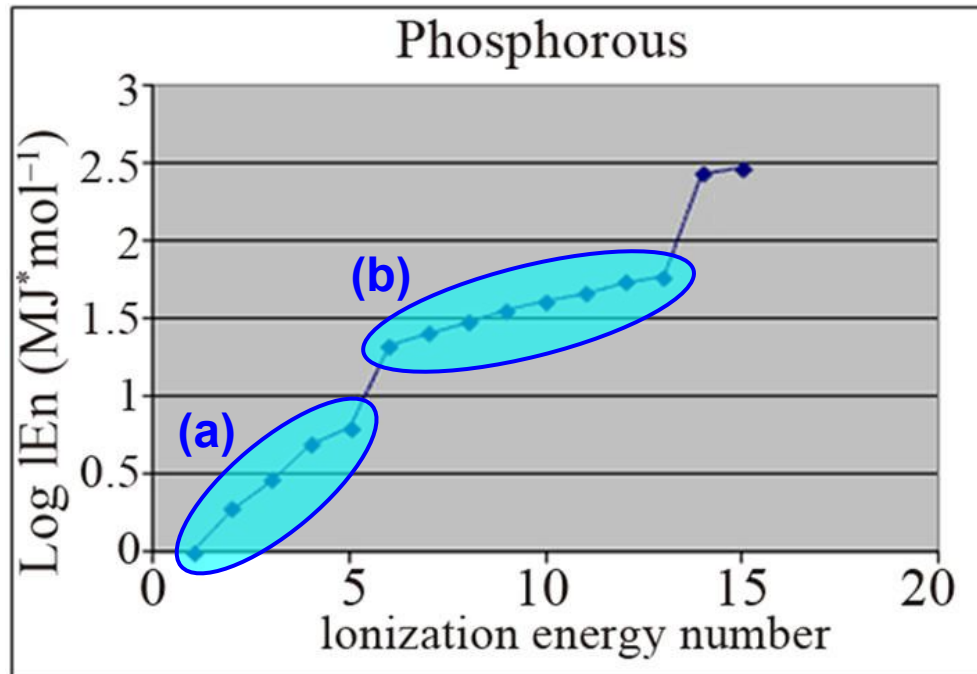
Evidence of Atomic Structure – Successive Ionisation Energies



- The graph is not linear, but can be divided into *three* distinct regions, which reflect the *three* electron shells:
(a) Removal of the *five electrons* in the *valence electron shell*.

Introduction to Atomic Structure

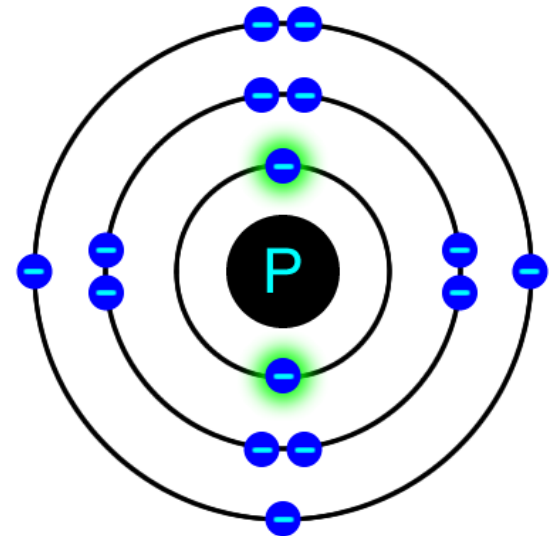
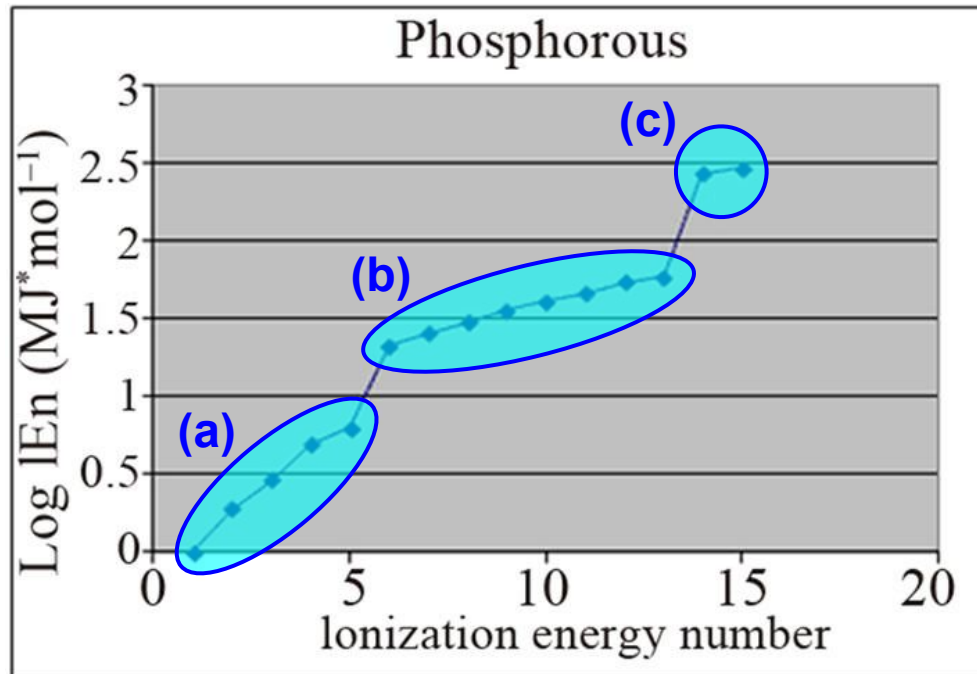
Evidence of Atomic Structure – Successive Ionisation Energies



- The graph is not linear, but can be divided into *three* distinct regions, which reflect the *three* electron shells:
 - (a) Removal of the *five electrons* in the *valence electron shell*.
 - (b) Removal of the *eight electrons* in the *second shell*.

Introduction to Atomic Structure

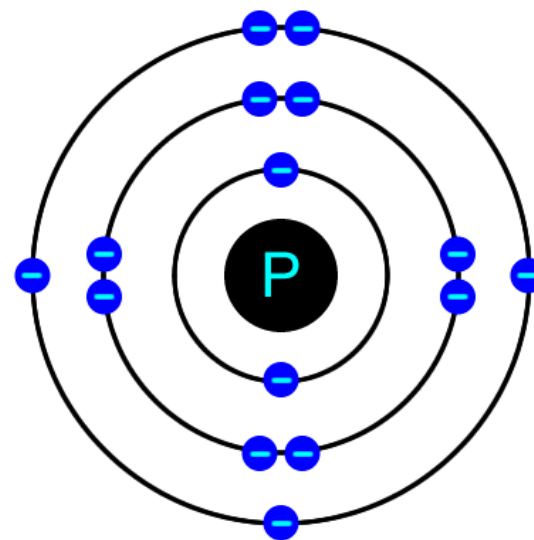
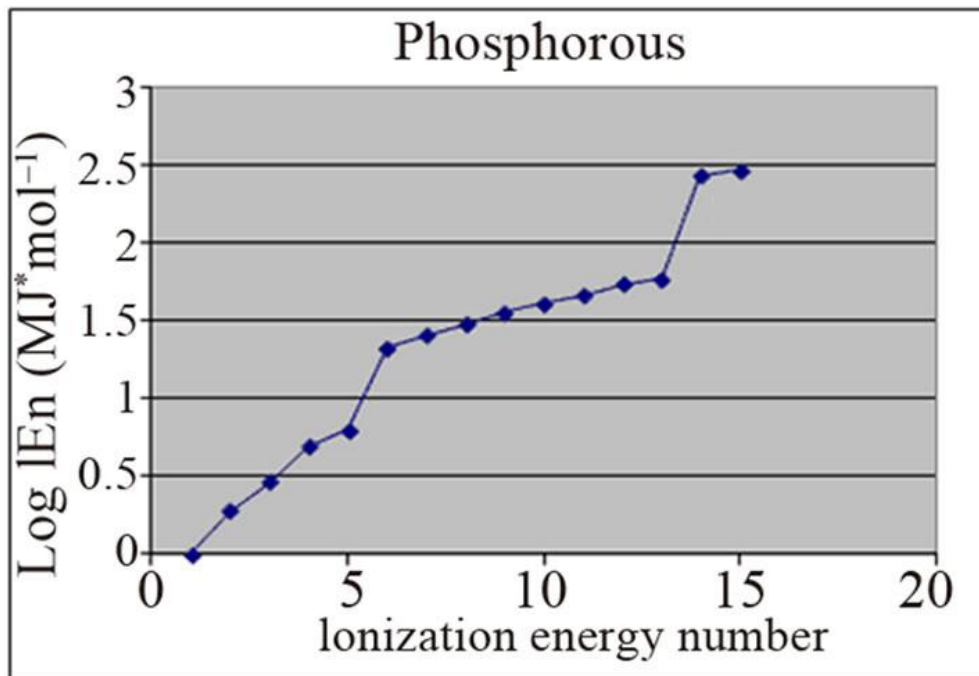
Evidence of Atomic Structure – Successive Ionisation Energies



- The graph is not linear, but can be divided into *three* distinct regions, which reflect the *three* electron shells:
 - (a) Removal of the *five electrons* in the *valence electron shell*.
 - (b) Removal of the *eight electrons* in the *second shell*.
 - (c) Removal of the *two electrons* in the *inner shell*.

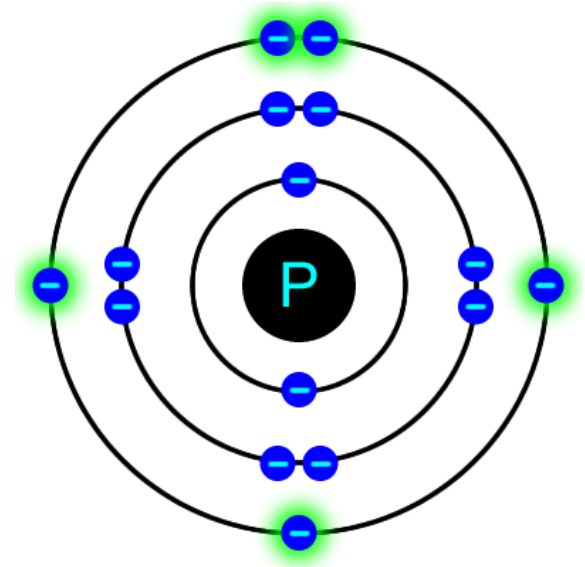
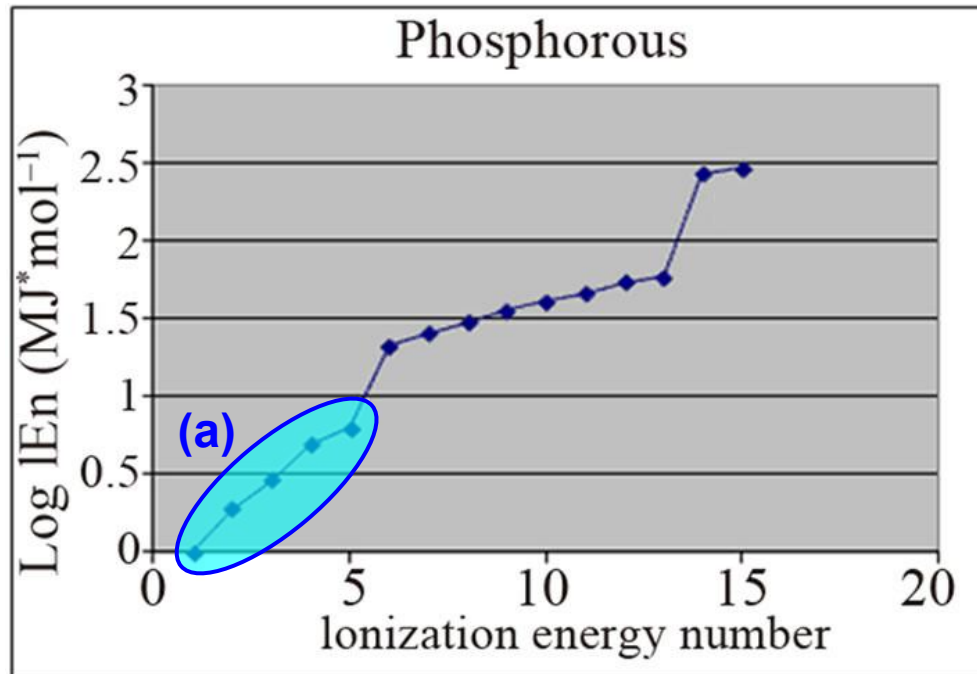
Introduction to Atomic Structure

Evidence of Atomic Structure – Successive Ionisation Energies



Introduction to Atomic Structure

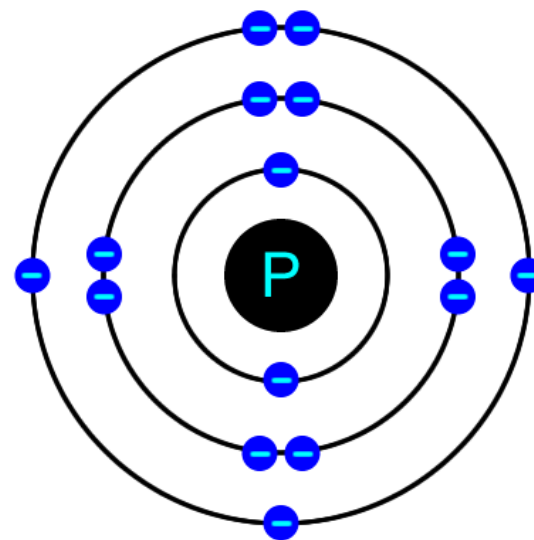
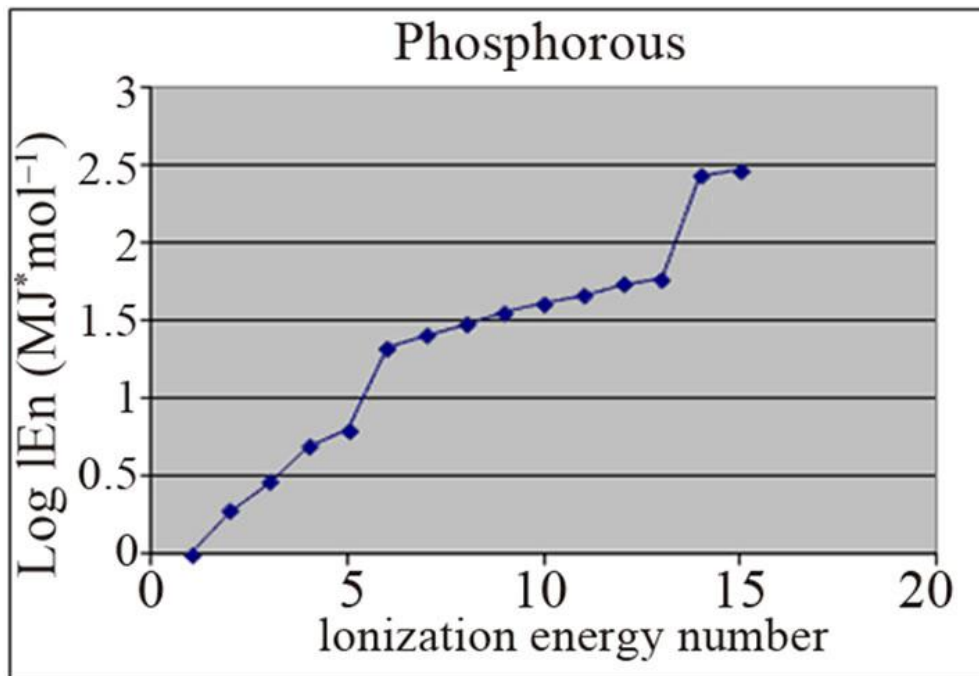
Evidence of Atomic Structure – Successive Ionisation Energies



(a) The *five electrons* in the *valence electron shell* are *furthest* from the nucleus, and so the electrostatic force of attraction between the nucleus and negatively charged electrons is *weakest*. A relatively *small amount of energy* is required to remove the five valence electrons.

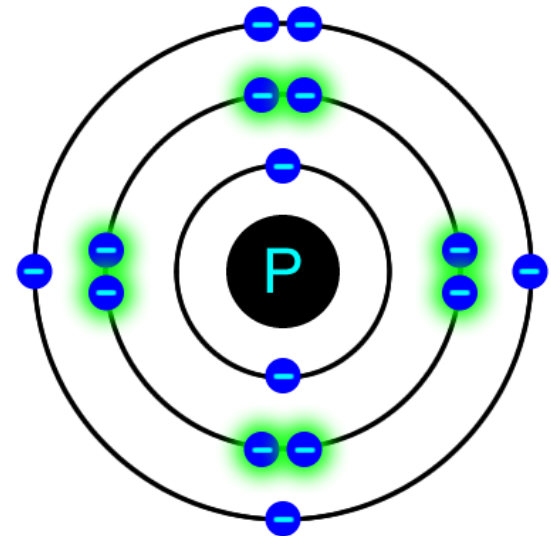
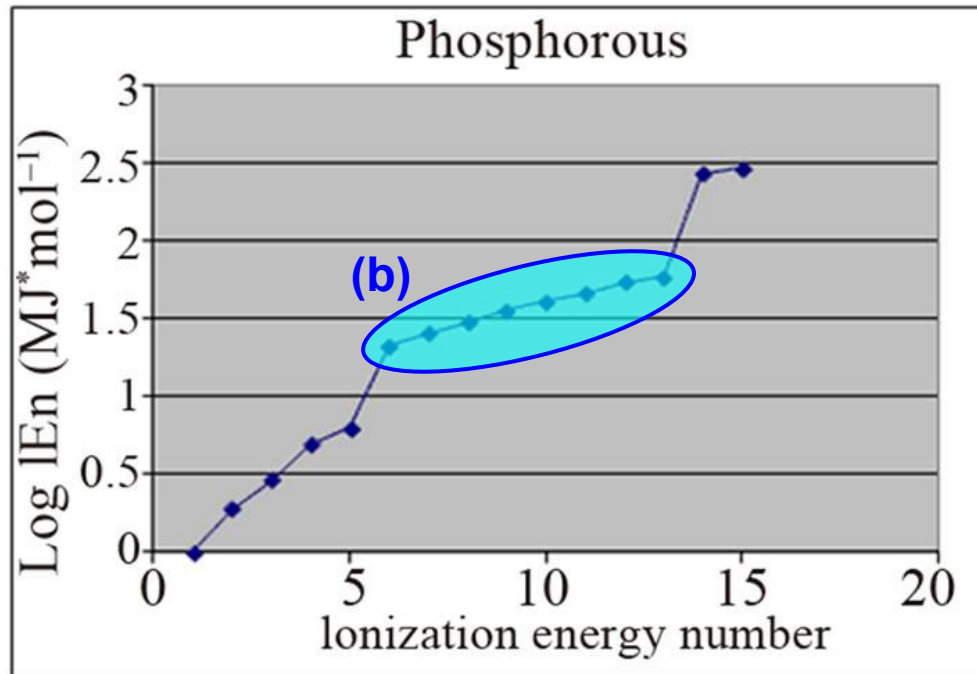
Introduction to Atomic Structure

Evidence of Atomic Structure – Successive Ionisation Energies



Introduction to Atomic Structure

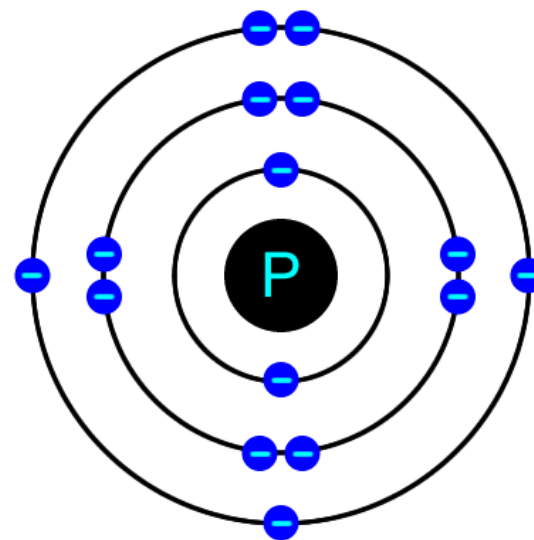
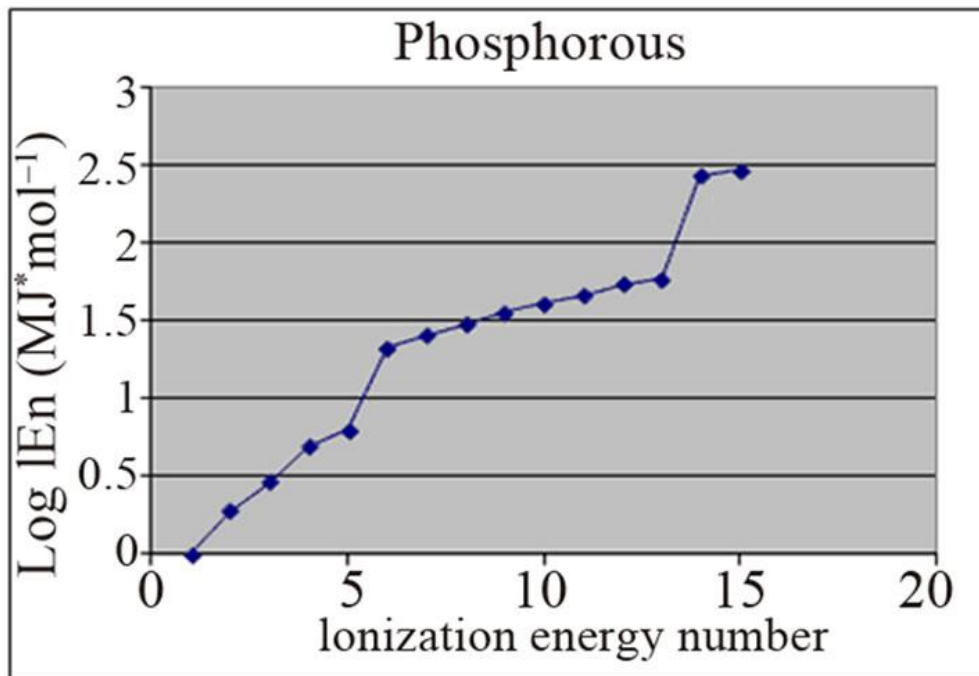
Evidence of Atomic Structure – Successive Ionisation Energies



(b) There is a significant *increase in energy* between **(a)** and **(b)** as electrons are now being removed from the *second* electron shell which is *closer to the nucleus*, hence electrostatic forces of attraction are *stronger*, and need *more energy* to overcome.

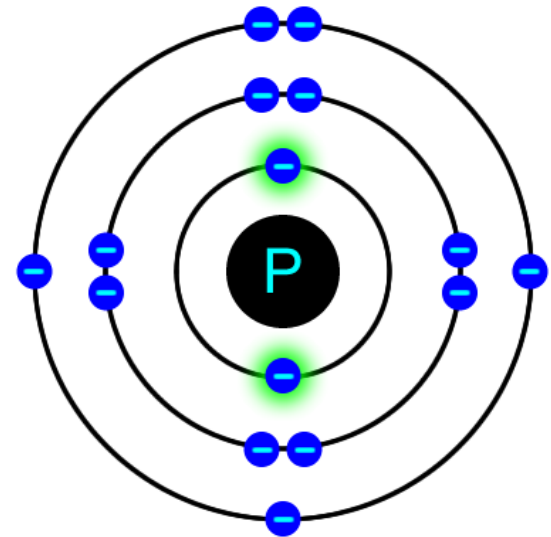
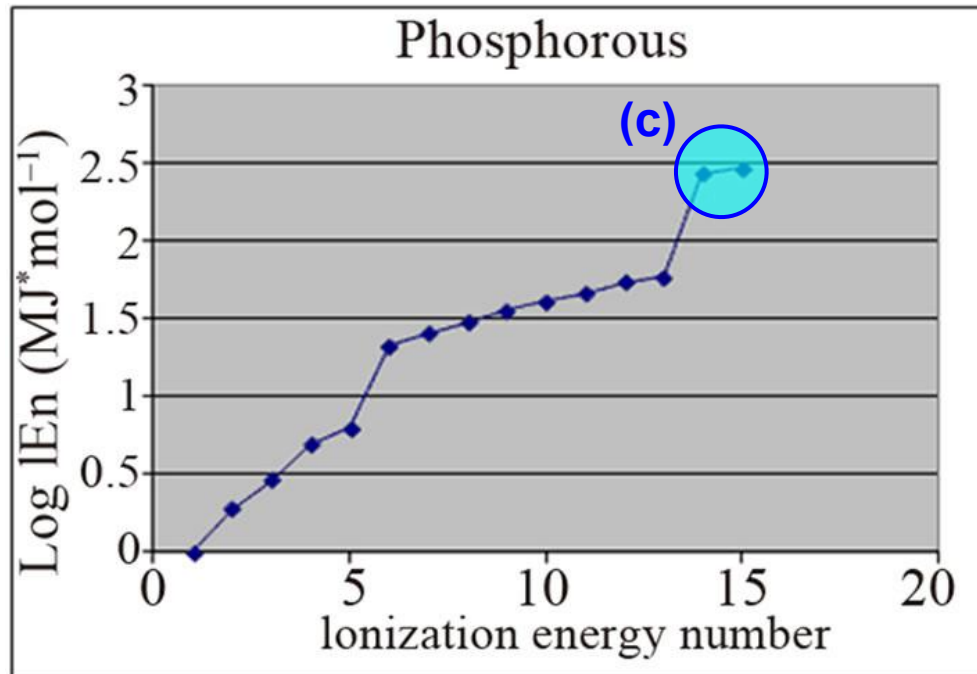
Introduction to Atomic Structure

Evidence of Atomic Structure – Successive Ionisation Energies



Introduction to Atomic Structure

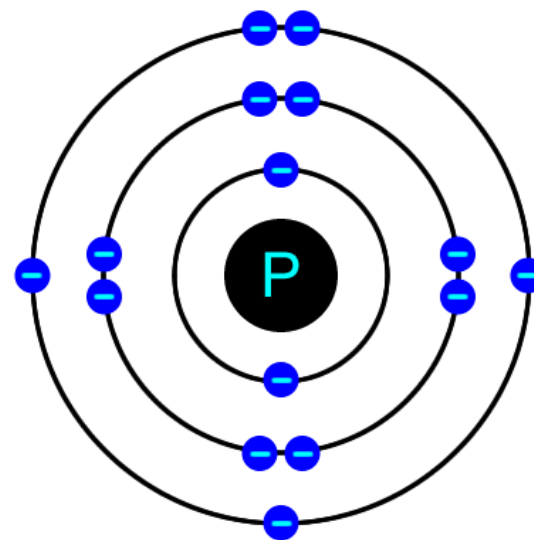
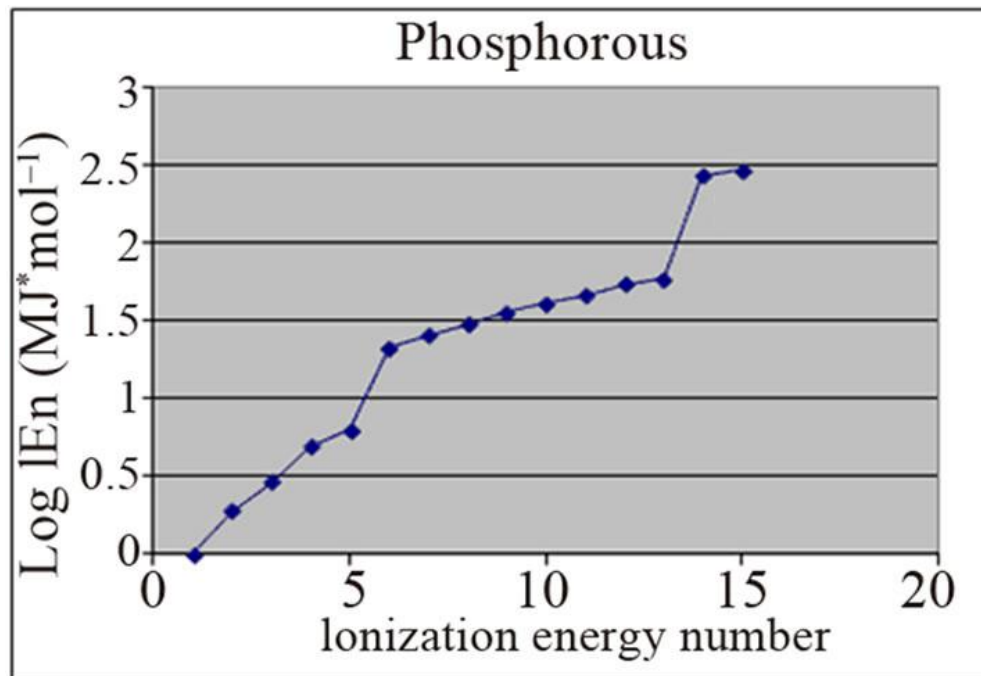
Evidence of Atomic Structure – Successive Ionisation Energies



(c) There is a substantial *increase in energy* between **(b)** and **(c)** as electrons are now being removed from the *innermost* electron shell which is *closest to the nucleus*, hence electrostatic forces of attraction are *very strong*, and need *the most energy* to overcome.

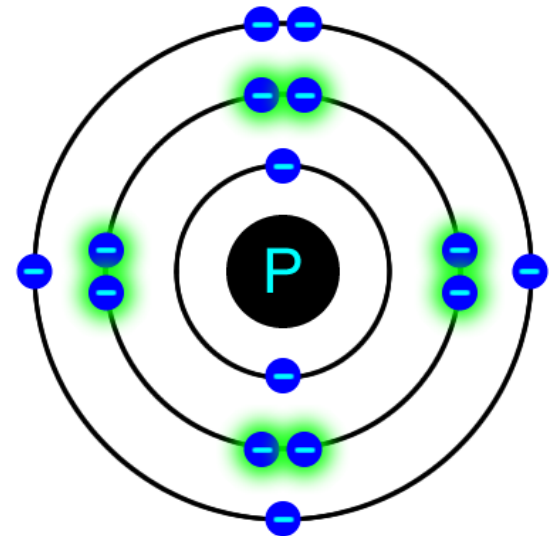
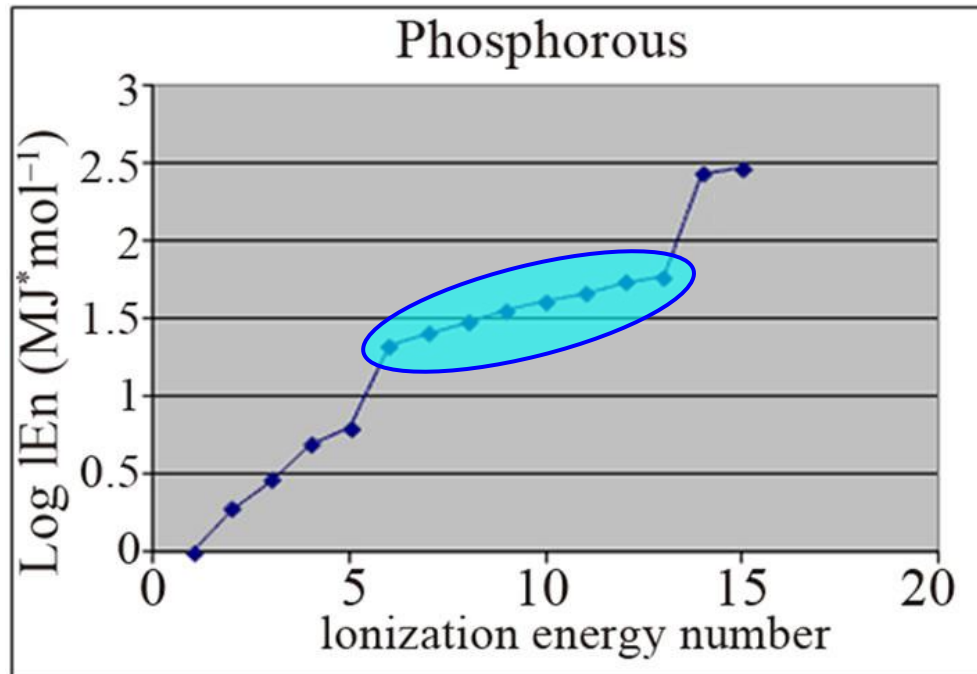
Introduction to Atomic Structure

Evidence of Atomic Structure – Successive Ionisation Energies



Introduction to Atomic Structure

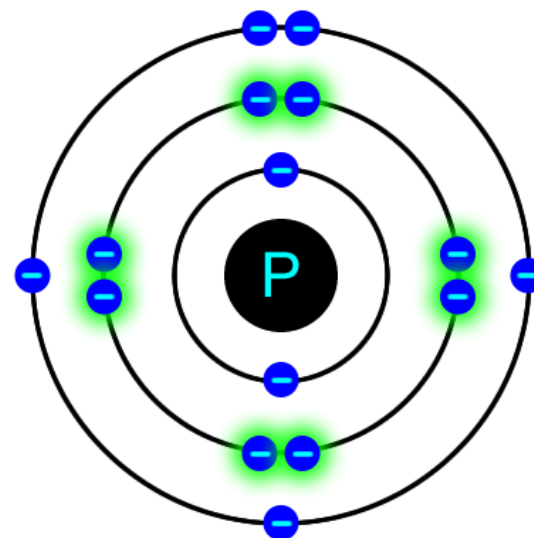
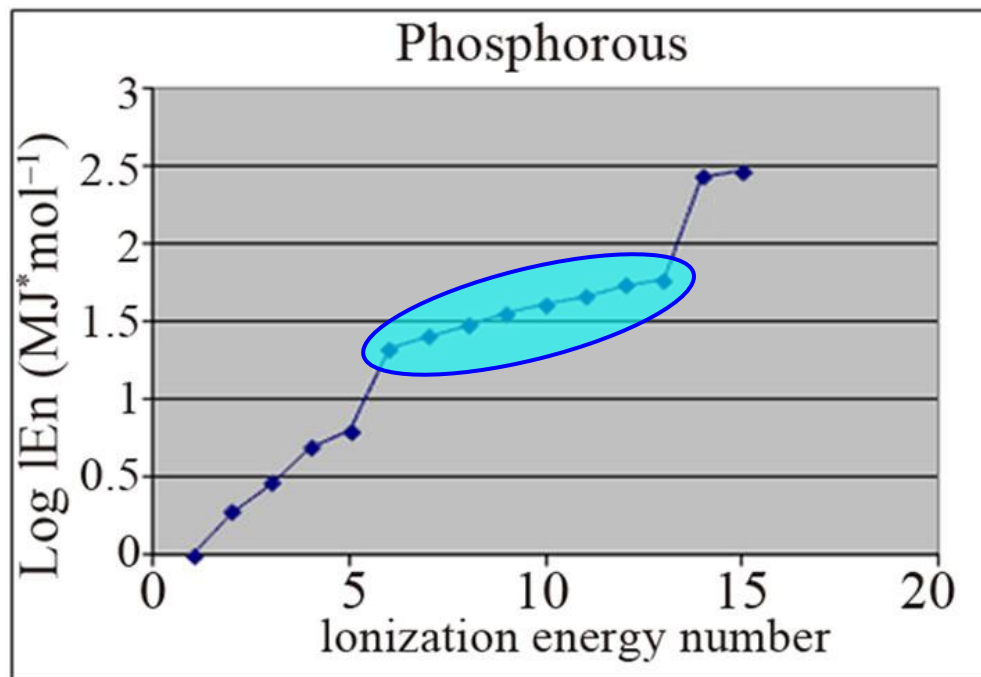
Evidence of Atomic Structure – Successive Ionisation Energies



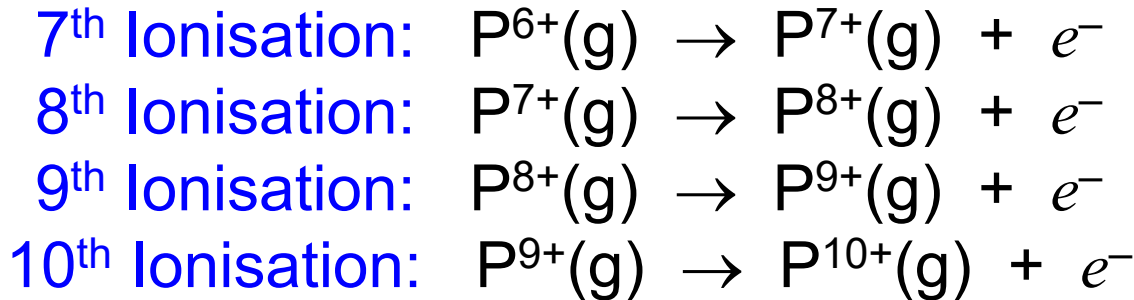
Note: The energy required to remove electrons from the *same* electron shell *gradually increases*. This is because the phosphorus ion is steadily becoming *more positively charged*, hence *more energy is required* to overcome the *stronger electrostatic force of attraction* between the negatively charged electron and *increasingly positive phosphorus ion*.

Introduction to Atomic Structure

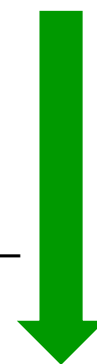
Evidence of Atomic Structure – Successive Ionisation Energies



Increasing amount of positive charge on phosphorus ion.



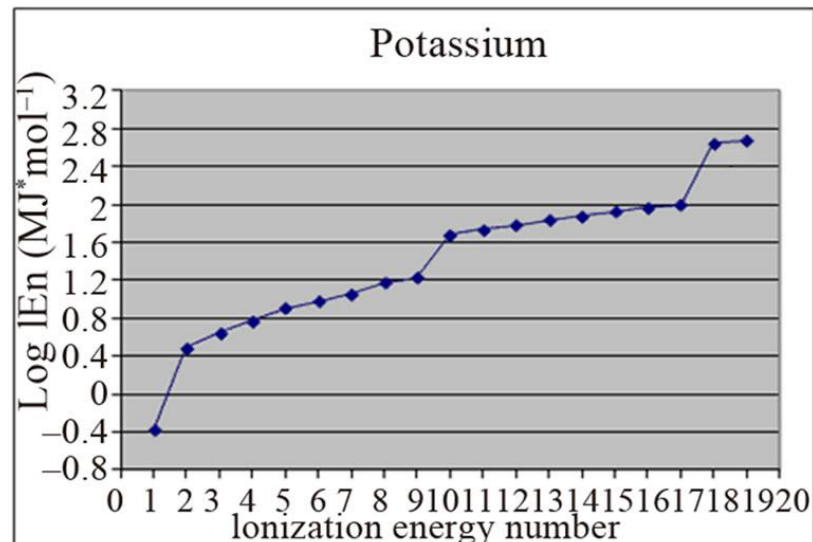
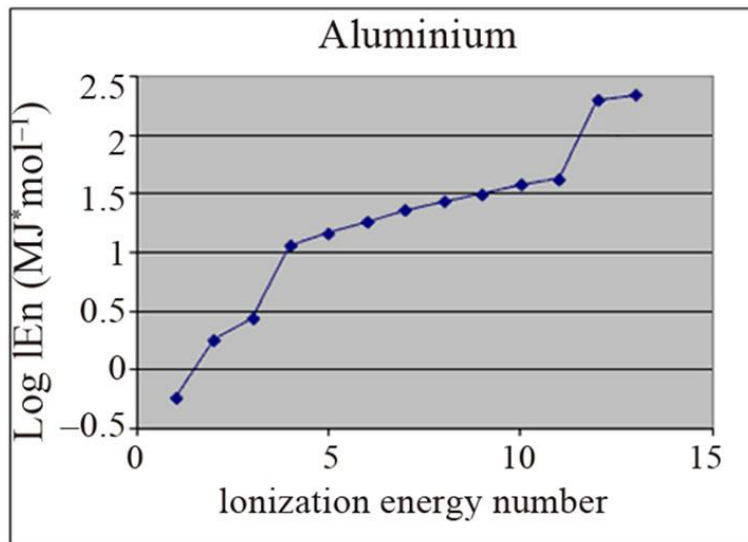
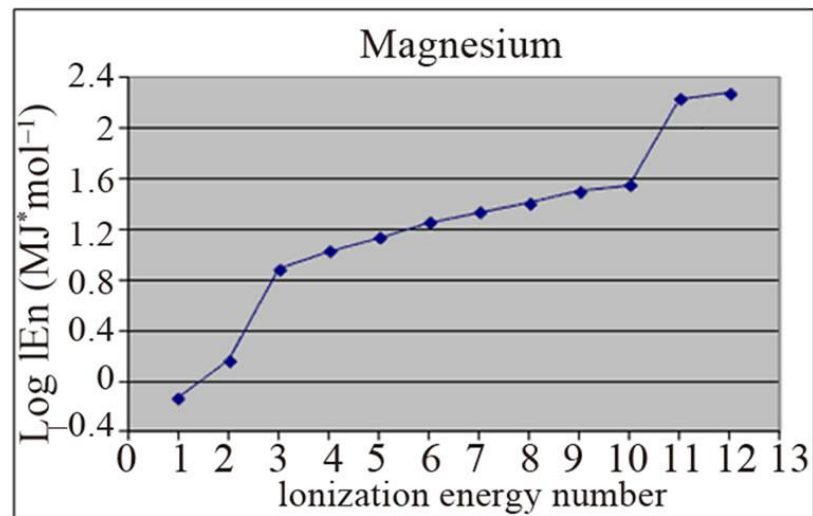
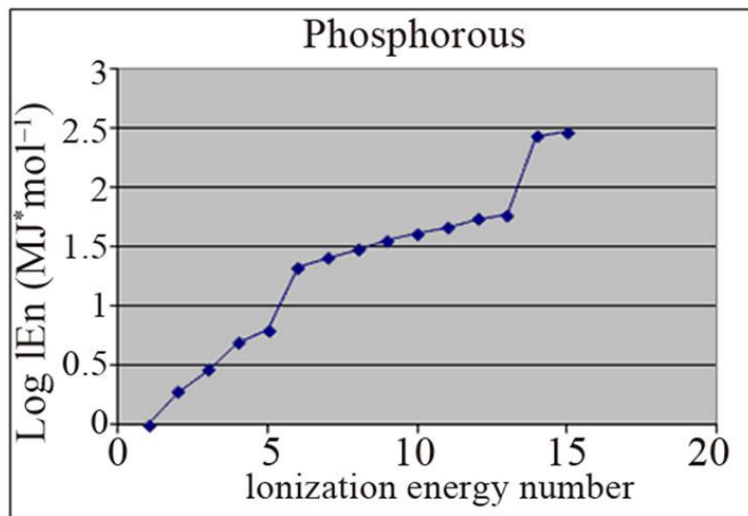
...and so on...



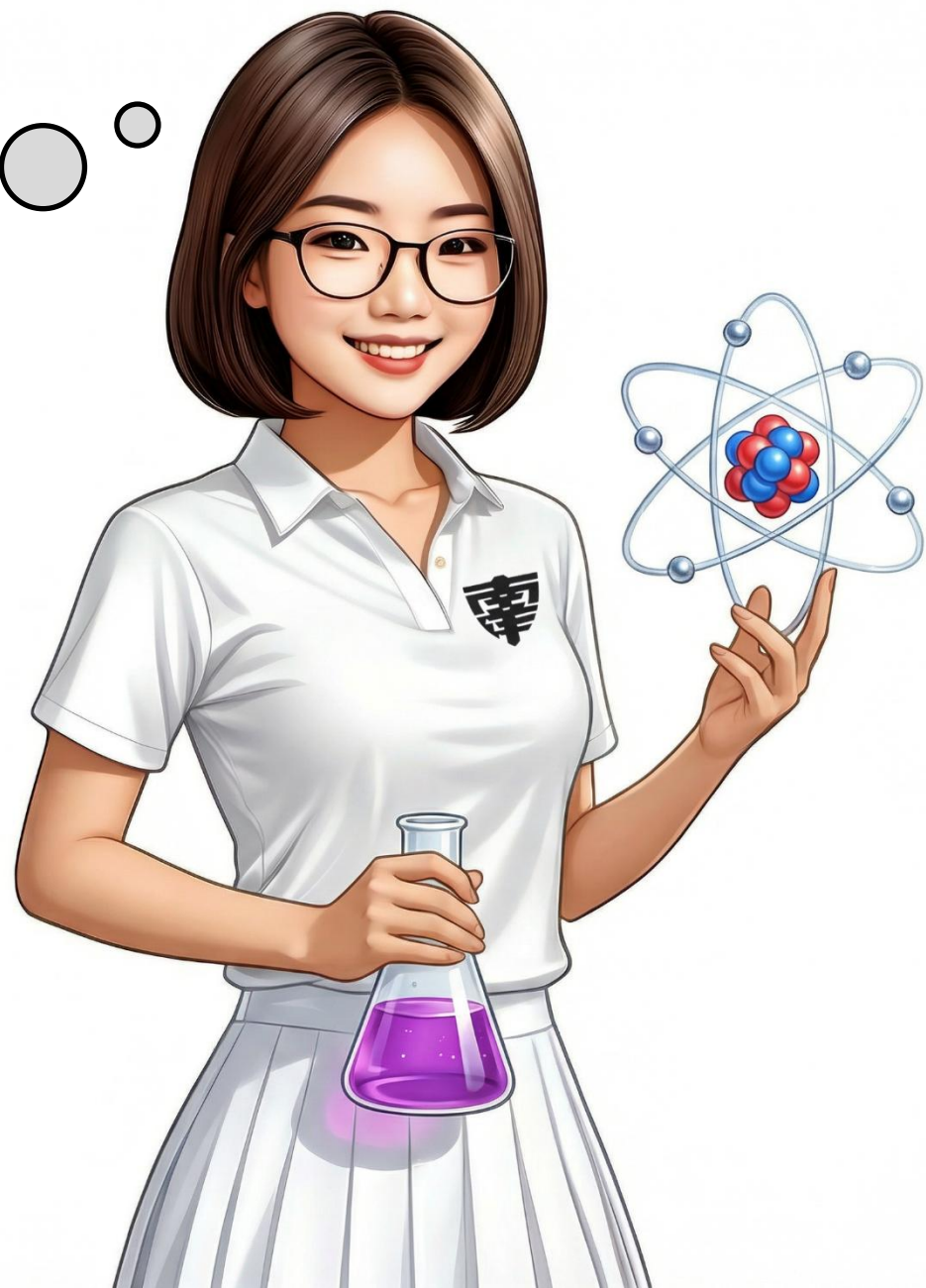
Increasing electrostatic force of attraction between e^{-} and ion.

Introduction to Atomic Structure

Evidence of Atomic Structure – Successive Ionisation Energies



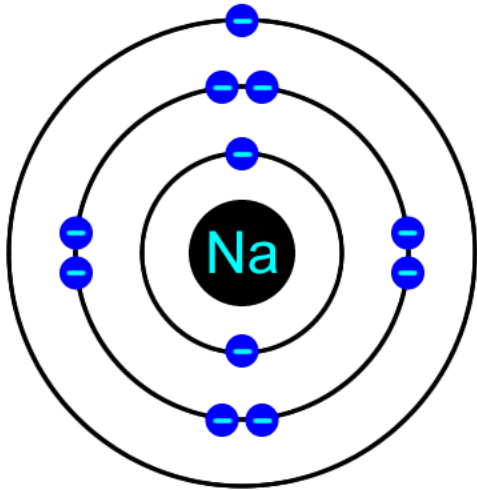
Can I tell whether an atom belongs to a metallic or a non-metallic element from its electronic configuration?



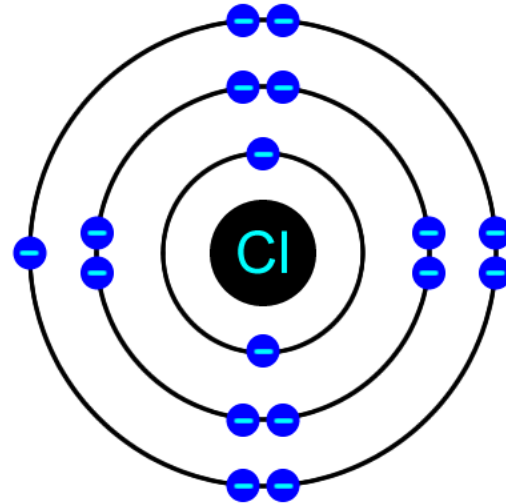
Introduction to Atomic Structure

- Atoms of *metallic elements* tend to have only *one*, *two* or *three* electrons in their valence electron shell.
- Atoms of *non-metallic elements* tend to have *four*, *five*, *six*, *seven* or *eight* electrons in their valence electron shell.
- **Note:** This is only a *general rule* and there are some exceptions.

Introduction to Atomic Structure

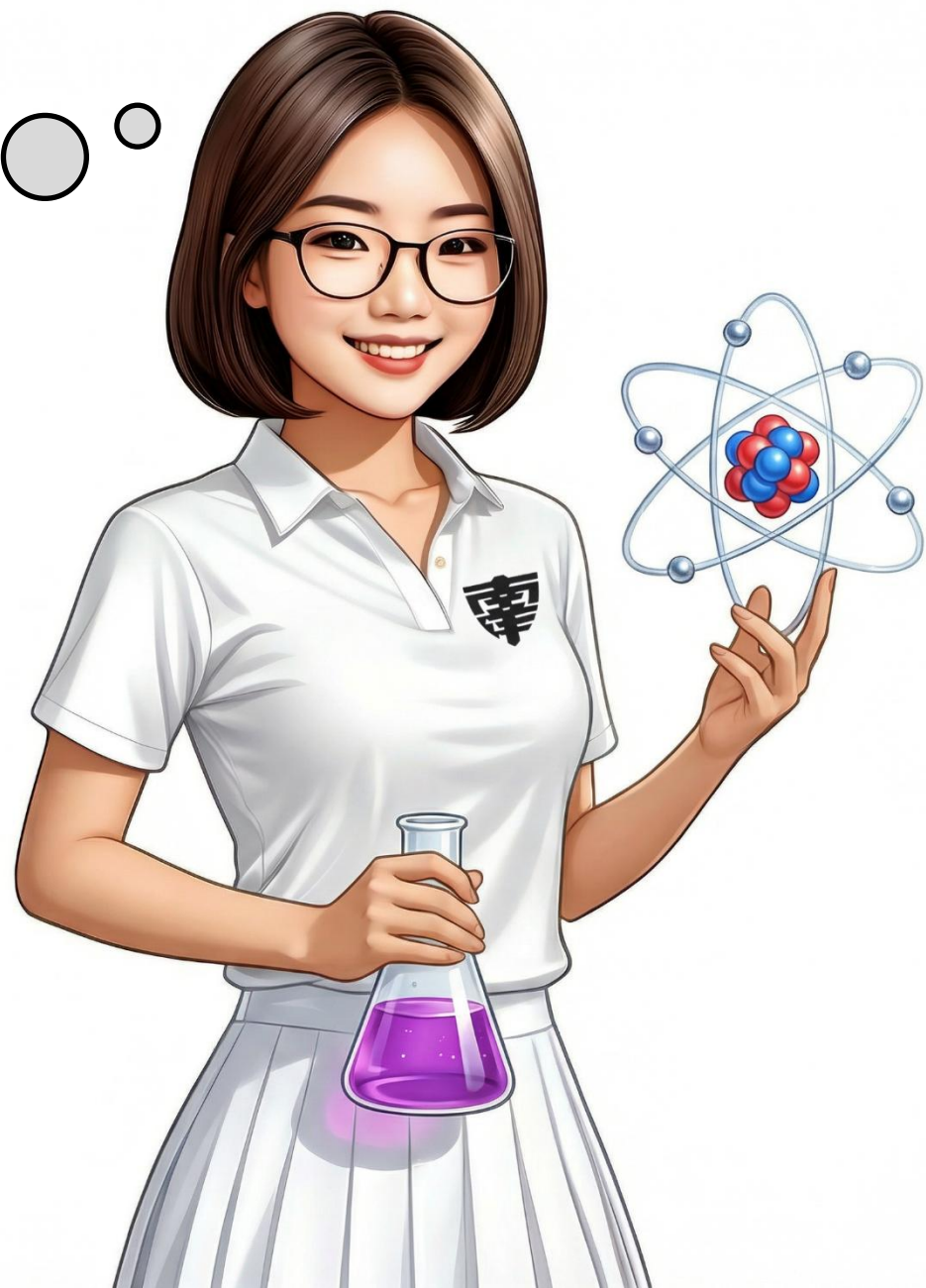


- An atom of *sodium* only has *one* electron in its valence electron shell, making it a *metal*.



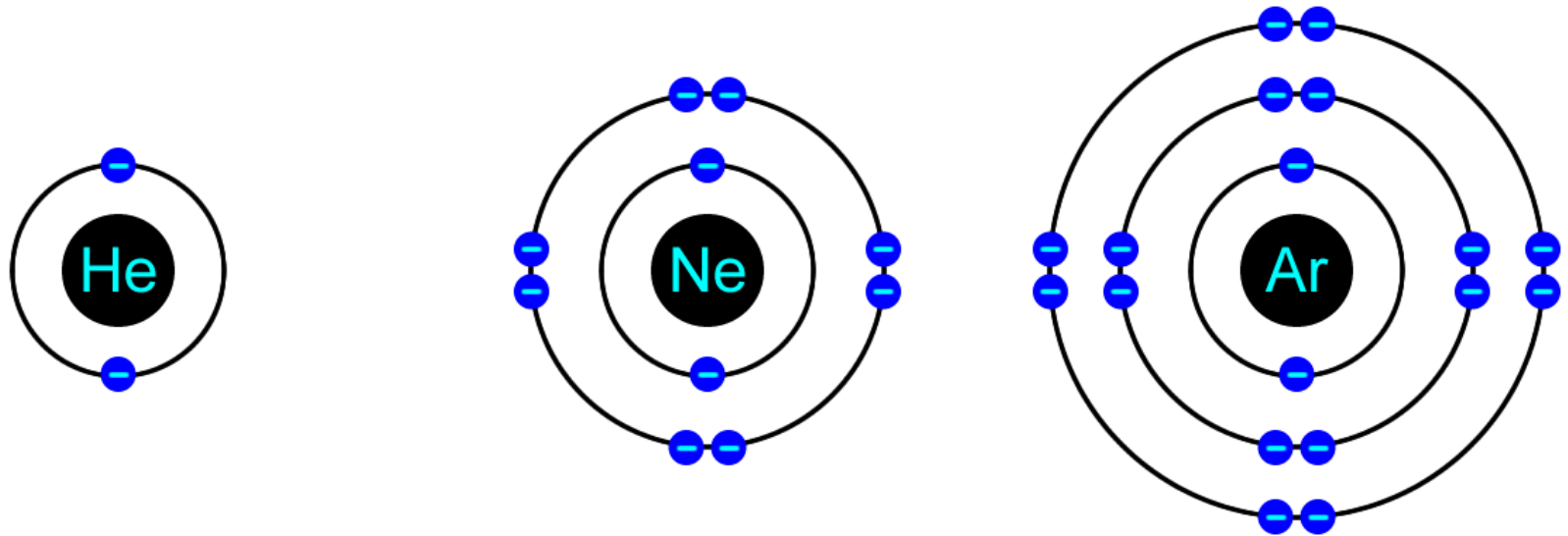
- An atom of *chlorine* has *seven* electrons in its valence shell, making it a *non-metal*.

Is there anything special about atoms that have a fully filled valence electron shell?



Introduction to Atomic Structure

- *Helium* (He), *neon* (Ne) and *argon* (Ar) are noble gases in Group 18 of the Periodic Table. Their atoms all have a complete or fully filled valence electron shell which is referred to as the *noble gas electronic configuration*.



- Chemical elements whose atoms have noble gas electronic configurations are *very stable* and *very unreactive*, and are often described as being *chemically inert*.

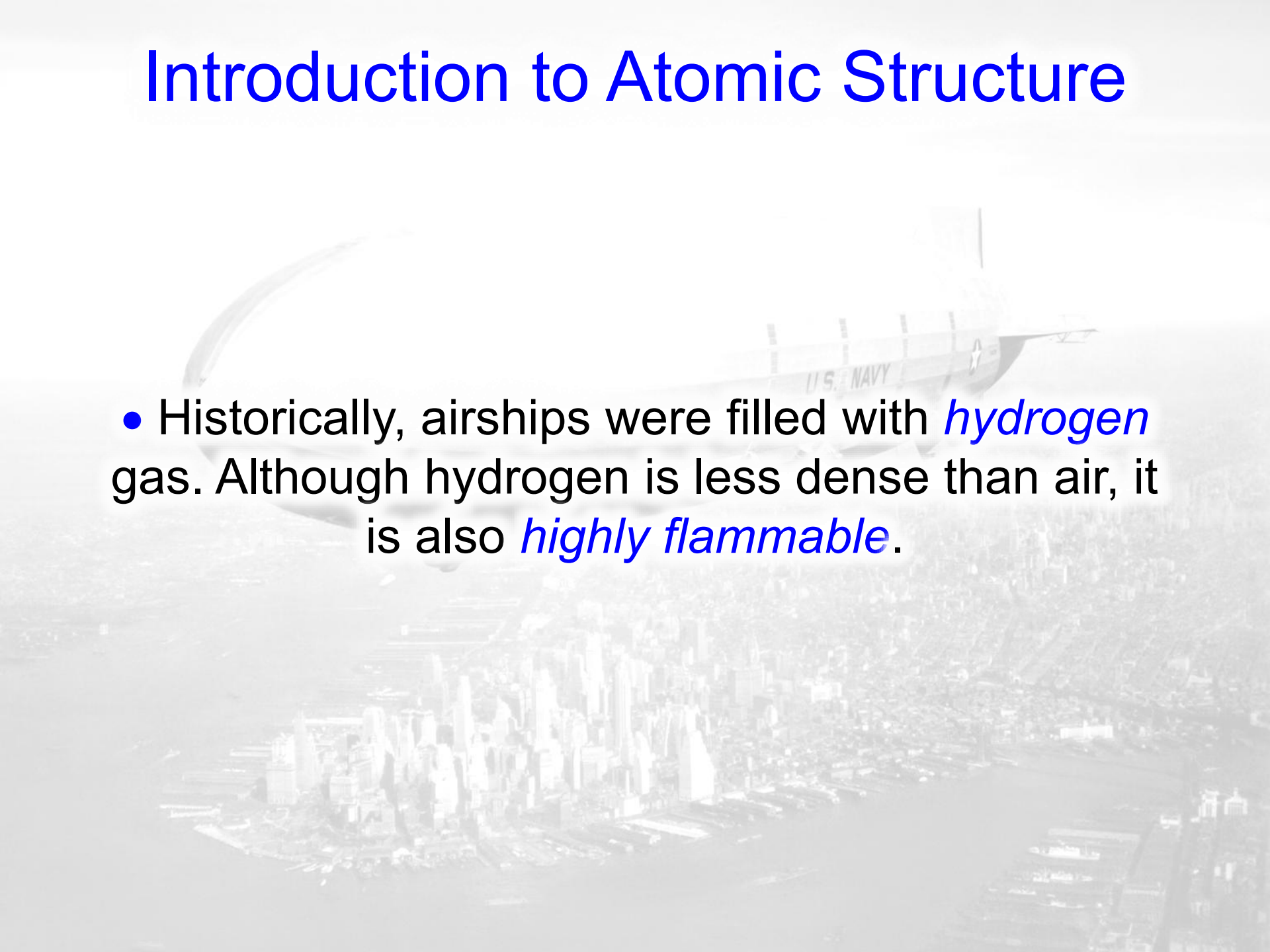
Introduction to Atomic Structure



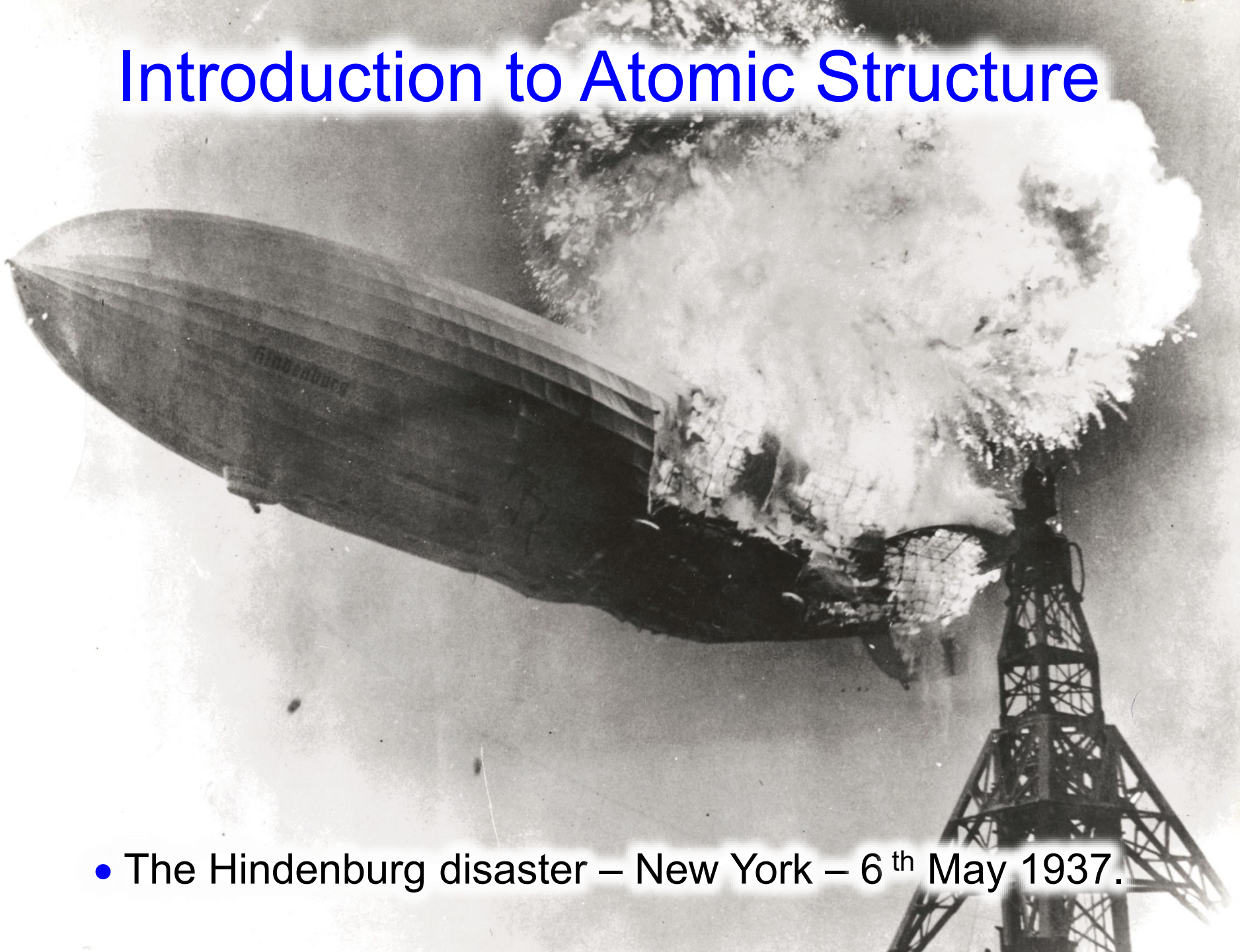
- USS Macon flying over New York harbour *ca.* 1933.

Introduction to Atomic Structure

- Historically, airships were filled with *hydrogen* gas. Although hydrogen is less dense than air, it is also *highly flammable*.



Introduction to Atomic Structure

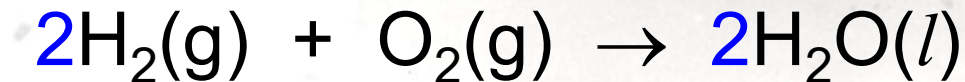


- The Hindenburg disaster – New York – 6th May 1937.

Introduction to Atomic Structure

- There were many disasters when airships filled with hydrogen gas caught fire. The highly flammable hydrogen burned very quickly and the entire airship was engulfed in flames in a matter of seconds.

hydrogen + oxygen → water

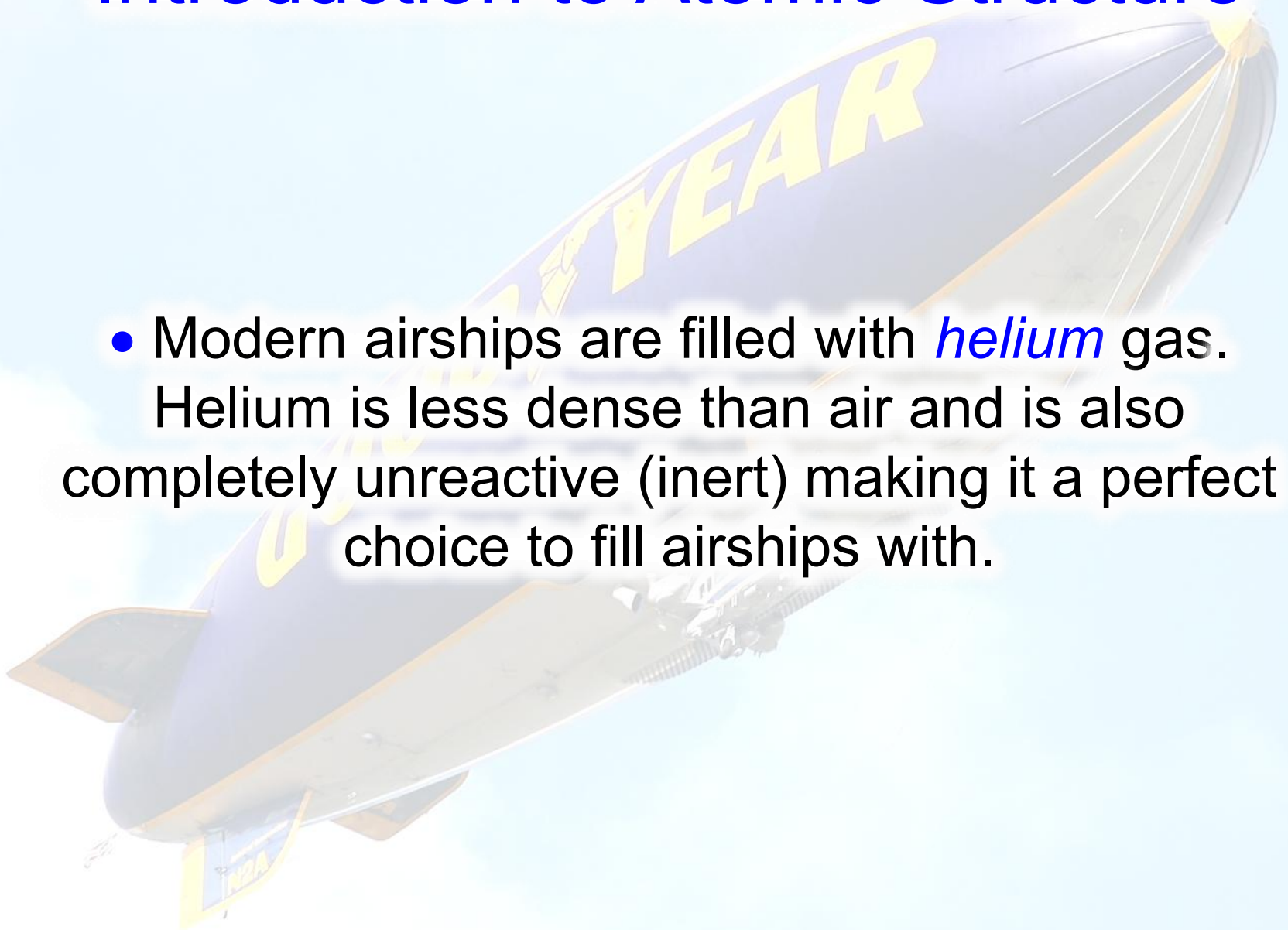


Introduction to Atomic Structure

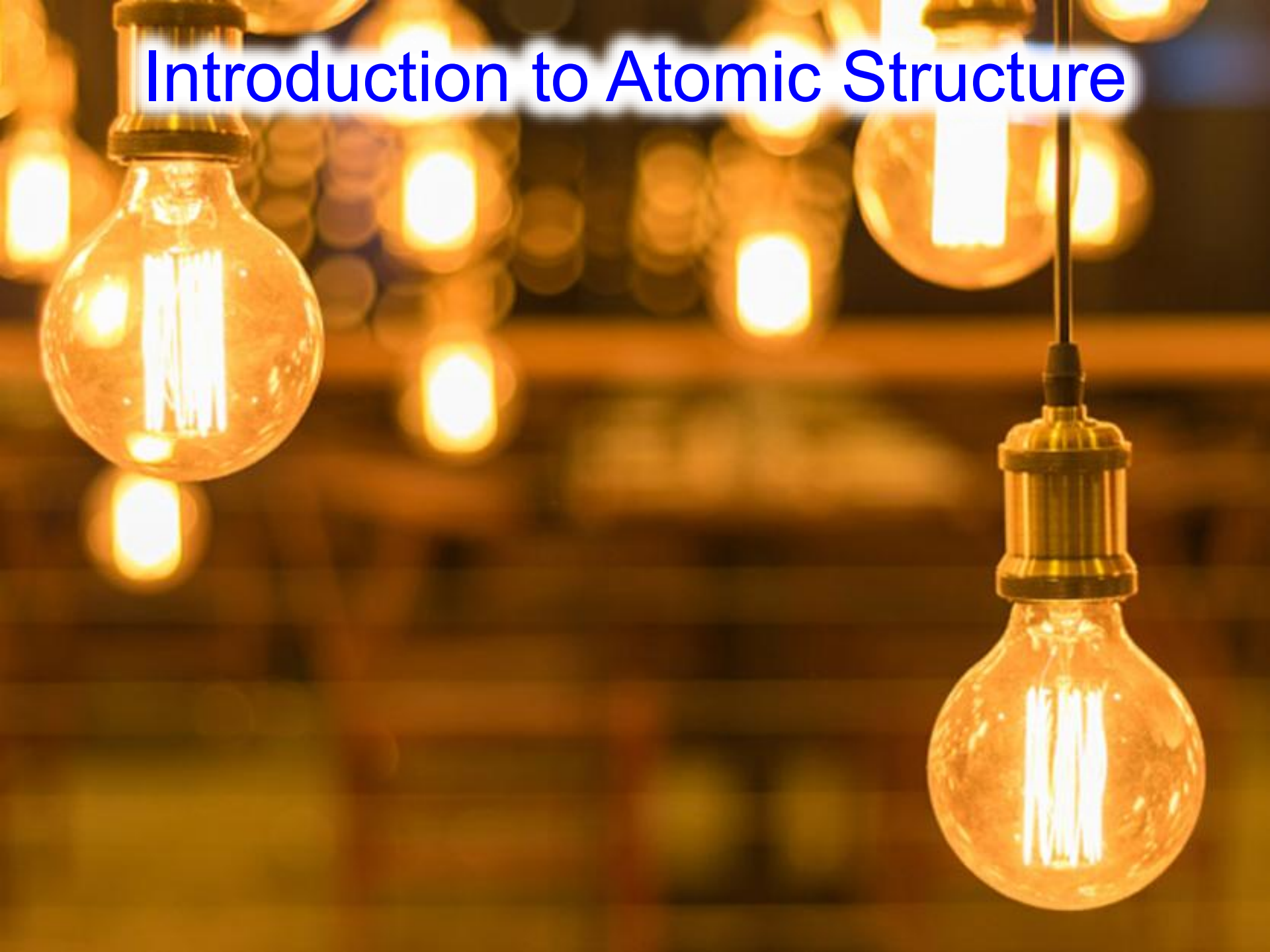


Introduction to Atomic Structure

- Modern airships are filled with *helium* gas. Helium is less dense than air and is also completely unreactive (inert) making it a perfect choice to fill airships with.



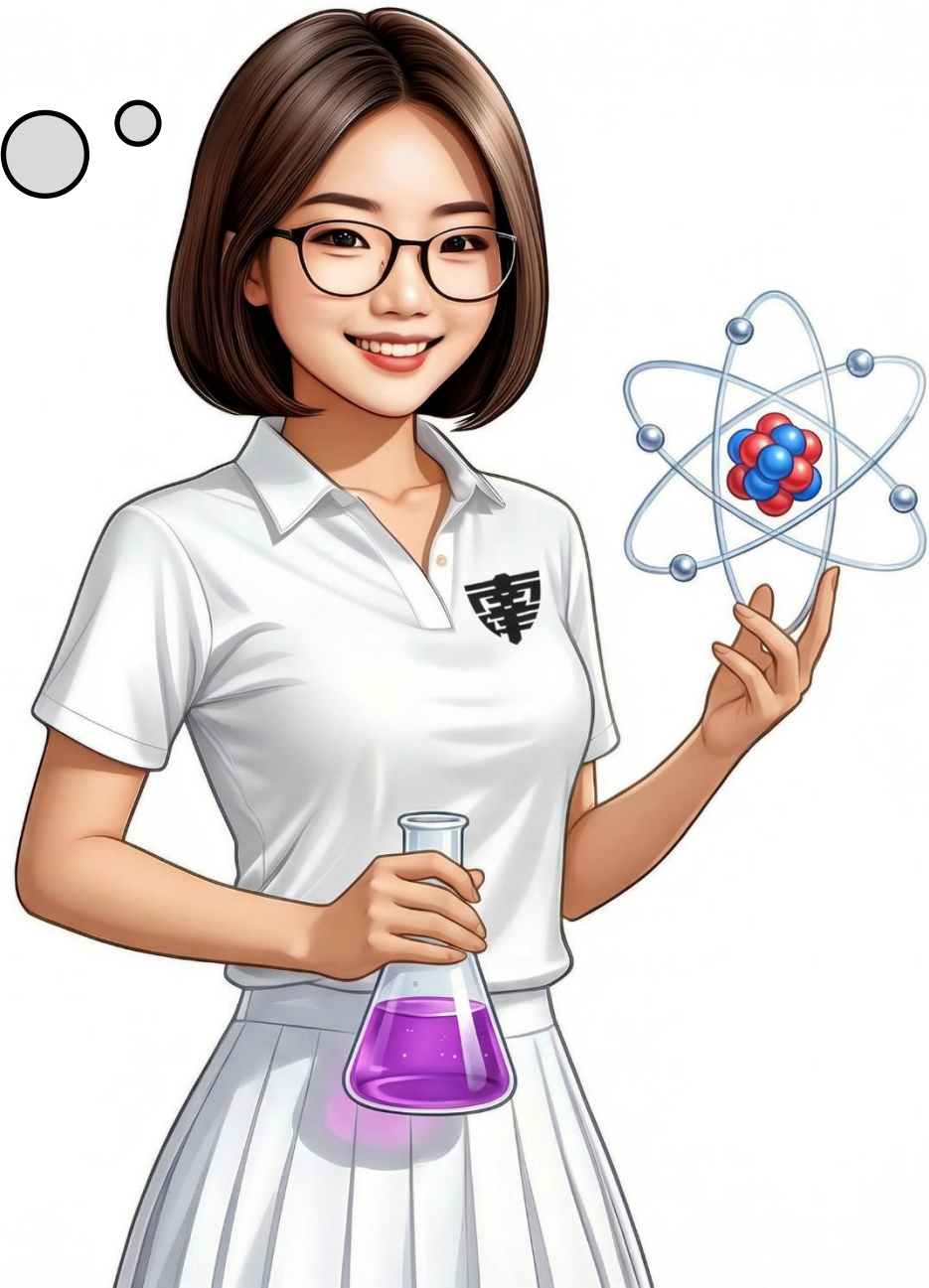
Introduction to Atomic Structure



Introduction to Atomic Structure

- Modern lightbulbs use energy efficient light emitting diodes (LEDs).
- Before these were widely used, lightbulbs contained a coiled filament of metal which became hot and glowed brightly when electricity passed through it.
- To prevent the hot filament reacting with oxygen in the air and 'burning out' these lightbulbs were filled with the Noble gas, argon.

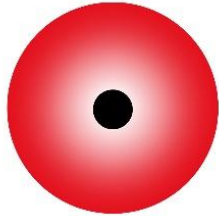
Can atoms of
the same
chemical
element have a
different
number of
neutrons?



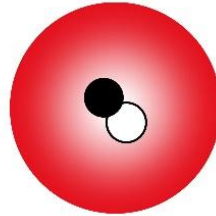
Isotopes

- *Isotopes* are atoms of the same chemical element with the *same Atomic Number* (the same number of protons) but a *different Mass Number* (a different number of neutrons).
- Because they have identical electronic configurations, isotopes will have the *same chemical properties*, *i.e.* they will react in the same way.
- Because they have different mass numbers, isotopes will have *different physical properties*, *e.g.* different densities, different melting points and different rates of diffusion.

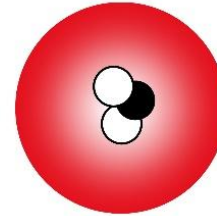
Isotopes



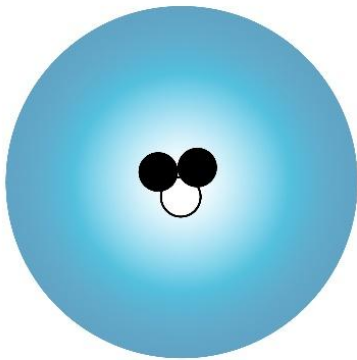
${}^1_1\text{H}$



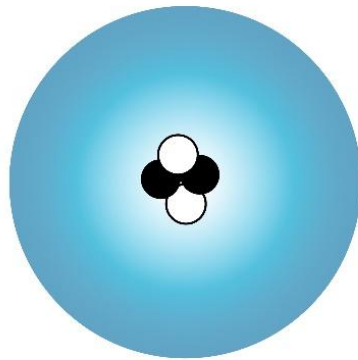
${}^2_1\text{H}$ = Deuterium



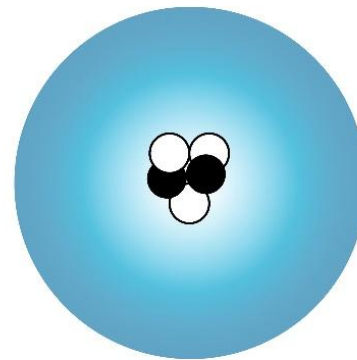
${}^3_1\text{H}$ = Tritium



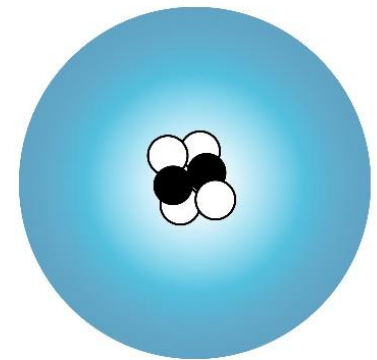
${}^3_2\text{He}$



${}^4_2\text{He}$

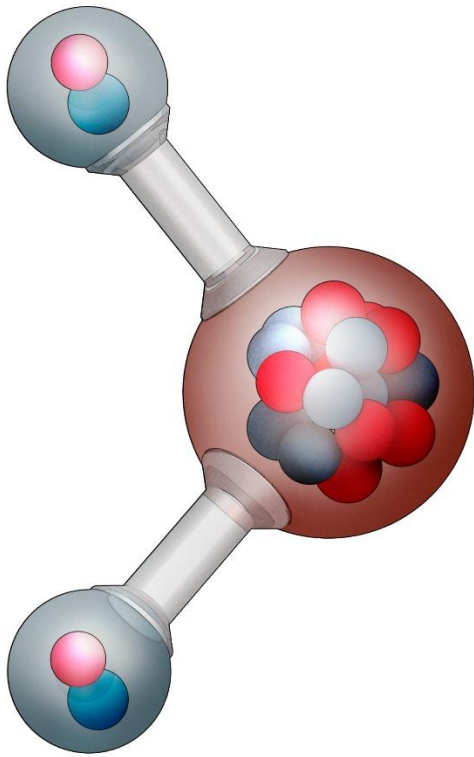


${}^5_2\text{He}$

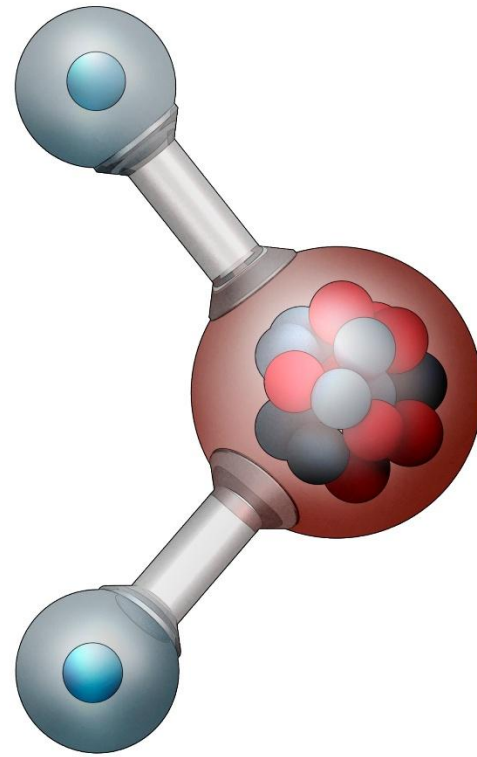


${}^6_2\text{He}$

Isotopes



- A molecule of *heavy water* that contains the heavy isotope of hydrogen known as *deuterium*. Deuterium contains a proton and a neutron. The relative molecular mass of heavy water = $2 + 2 + 16 = 20$.

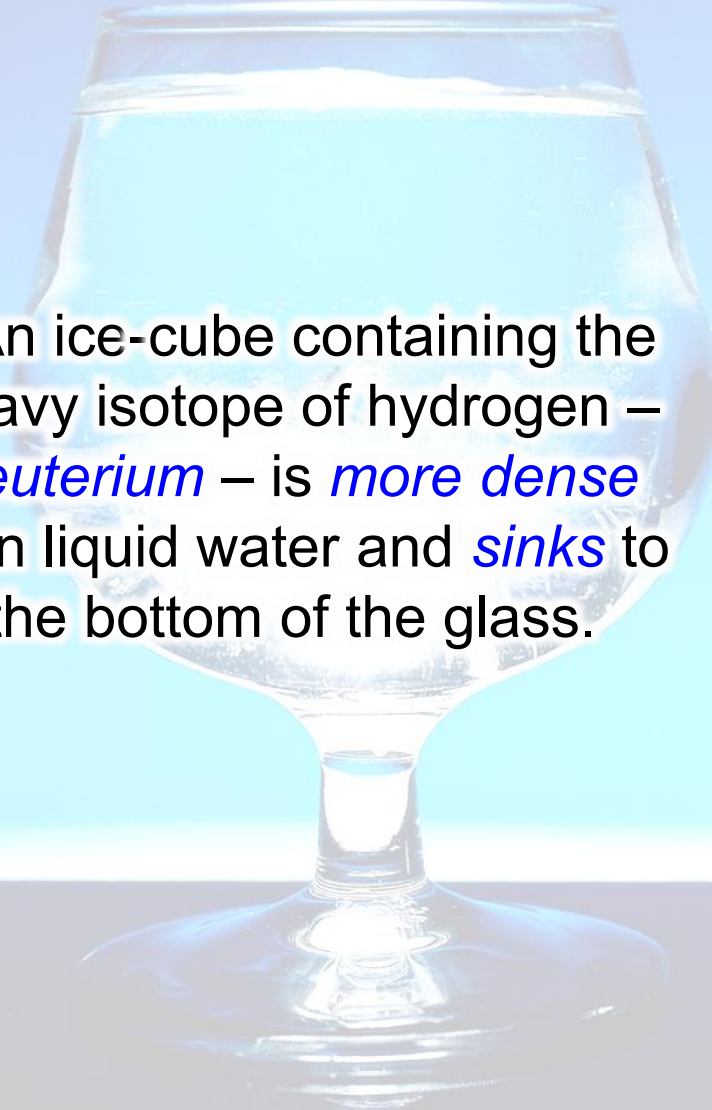


- A molecule of *common water* that contains the most common isotope of hydrogen. The nucleus of hydrogen contains a single proton. The relative molecular mass of common water = $1 + 1 + 16 = 18$.

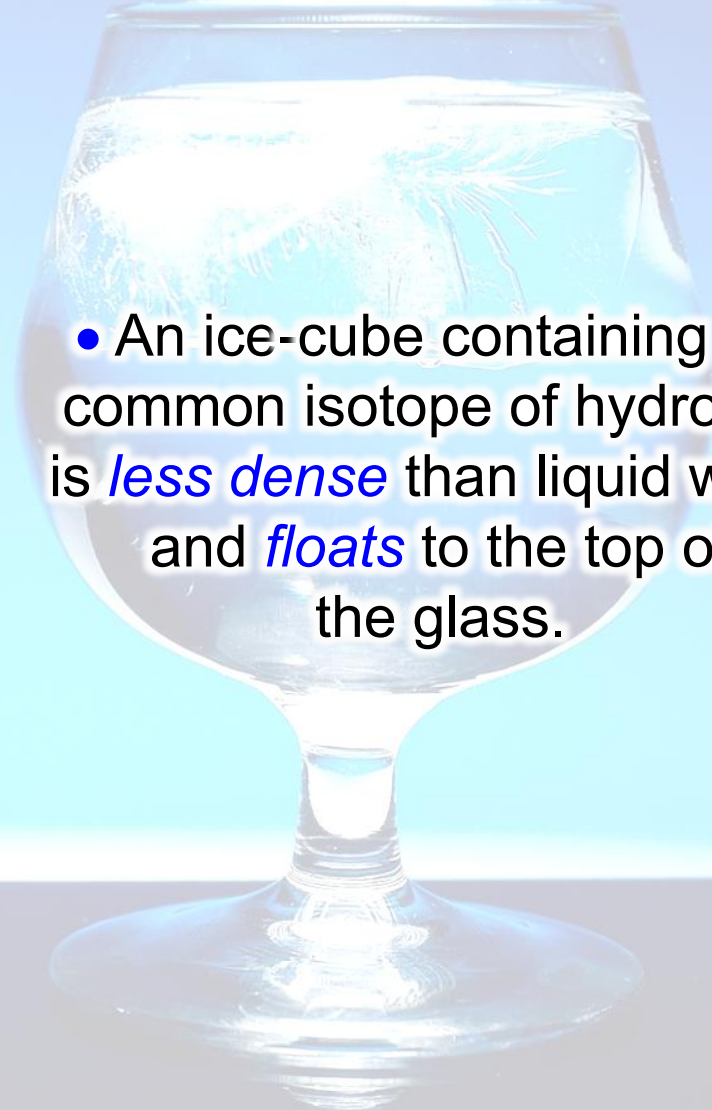
Isotopes



Isotopes



• An ice-cube containing the heavy isotope of hydrogen – *deuterium* – is *more dense* than liquid water and *sinks* to the bottom of the glass.



• An ice-cube containing the common isotope of hydrogen is *less dense* than liquid water and *floats* to the top of the glass.

Isotopes

Isotope of Carbon	Number of Protons	Number of Electrons	Number of Neutrons
$^{12}_6\text{C}$			
$^{13}_6\text{C}$			
$^{14}_6\text{C}$			

Isotopes

Isotope of Carbon	Number of Protons	Number of Electrons	Number of Neutrons
$^{12}_6\text{C}$	6		
$^{13}_6\text{C}$	6		
$^{14}_6\text{C}$	6		

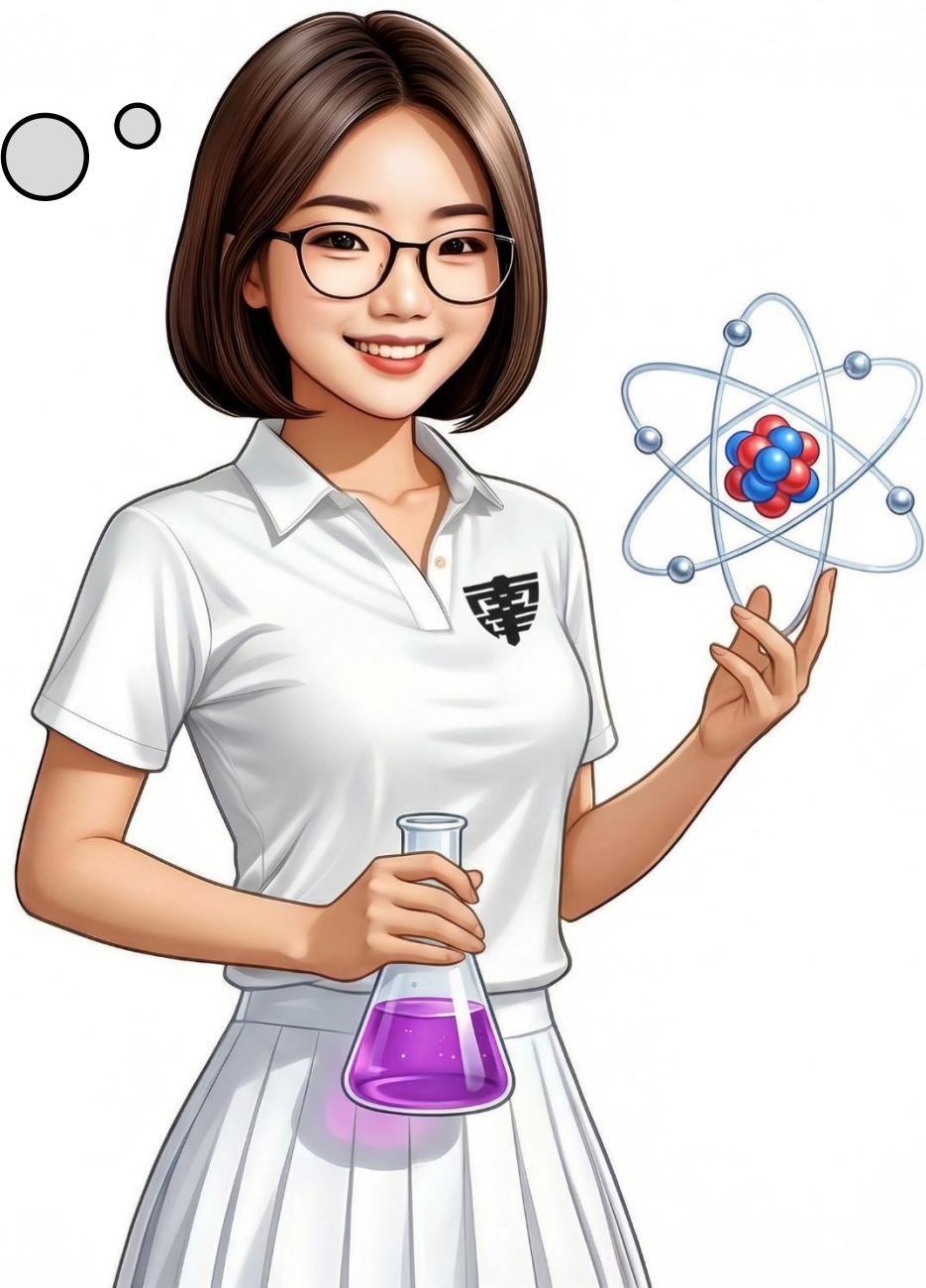
Isotopes

Isotope of Carbon	Number of Protons	Number of Electrons	Number of Neutrons
$^{12}_6\text{C}$	6	6	
$^{13}_6\text{C}$	6	6	
$^{14}_6\text{C}$	6	6	

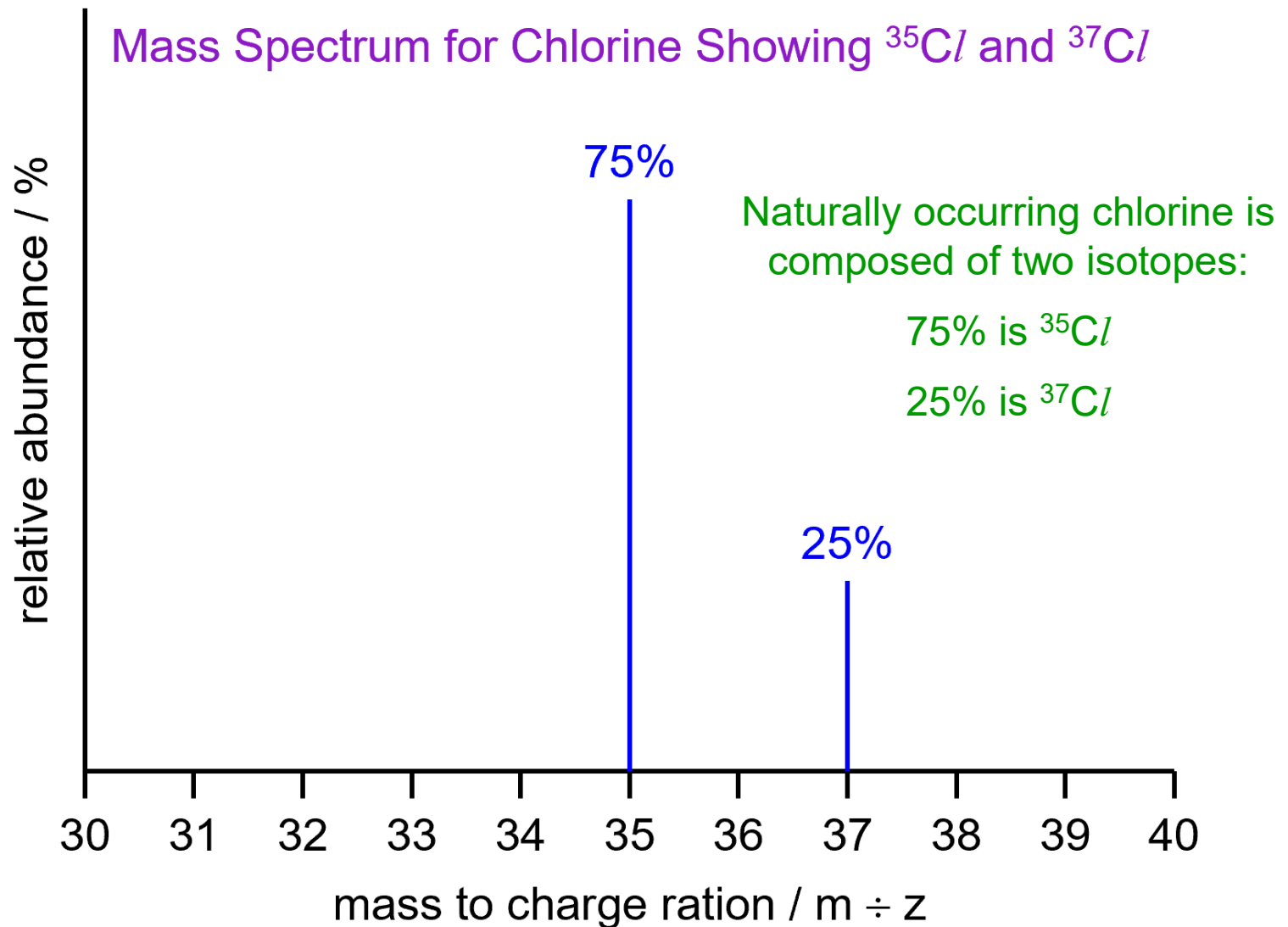
Isotopes

Isotope of Carbon	Number of Protons	Number of Electrons	Number of Neutrons
$\begin{matrix} 12 \\ 6 \end{matrix} \text{C}$	6	6	$12 - 6 = 6$
$\begin{matrix} 13 \\ 6 \end{matrix} \text{C}$	6	6	$13 - 6 = 7$
$\begin{matrix} 14 \\ 6 \end{matrix} \text{C}$	6	6	$14 - 6 = 8$

Why does chlorine have a Mass Number of 35.5? Is there such a thing as 0.5 of a proton or neutron?



Calculation of Relative Atomic Mass



Calculation of Relative Atomic Mass

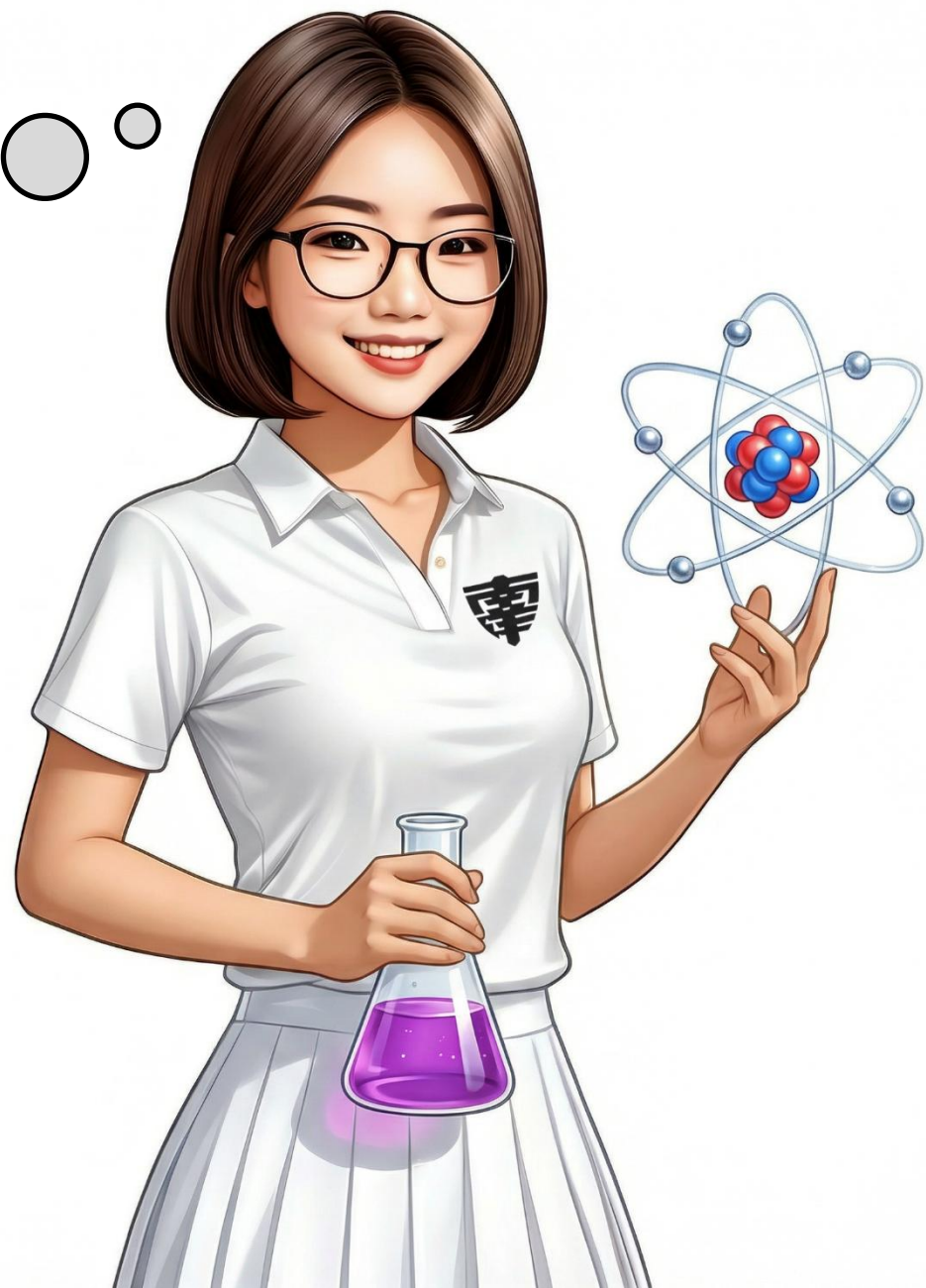
- Naturally occurring chlorine has two common isotopes, $^{35}_{17}\text{Cl}$ and $^{37}_{17}\text{Cl}$.
- $^{35}_{17}\text{Cl}$ contains 17 protons and $35 - 17 = 18$ neutrons while $^{37}_{17}\text{Cl}$ contains 17 protons and $37 - 17 = 20$ neutrons.
- The two isotopes of chlorine do not occur in equal quantities. 75% is $^{35}_{17}\text{Cl}$ and 25% is $^{37}_{17}\text{Cl}$.

Calculation of Relative Atomic Mass

- Naturally occurring chlorine has two common isotopes, $^{35}_{17}\text{Cl}$ and $^{37}_{17}\text{Cl}$.
- $^{35}_{17}\text{Cl}$ contains 17 protons and $35 - 17 = 18$ neutrons while $^{37}_{17}\text{Cl}$ contains 17 protons and $37 - 17 = 20$ neutrons.
 - The two isotopes of chlorine do not occur in equal quantities. 75% is $^{35}_{17}\text{Cl}$ and 25% is $^{37}_{17}\text{Cl}$.
- This means that if 100 atoms of chlorine were sampled at random, 75 would be $^{35}_{17}\text{Cl}$ and 25 would be $^{37}_{17}\text{Cl}$.
- The mass of 100 chlorine atoms would therefore be:
$$(75 \times 35) + (25 \times 37) = 3550.$$
- The average mass of a single chlorine atom would therefore be $3550 \div 100 = 35.5$ (3 s.f.).

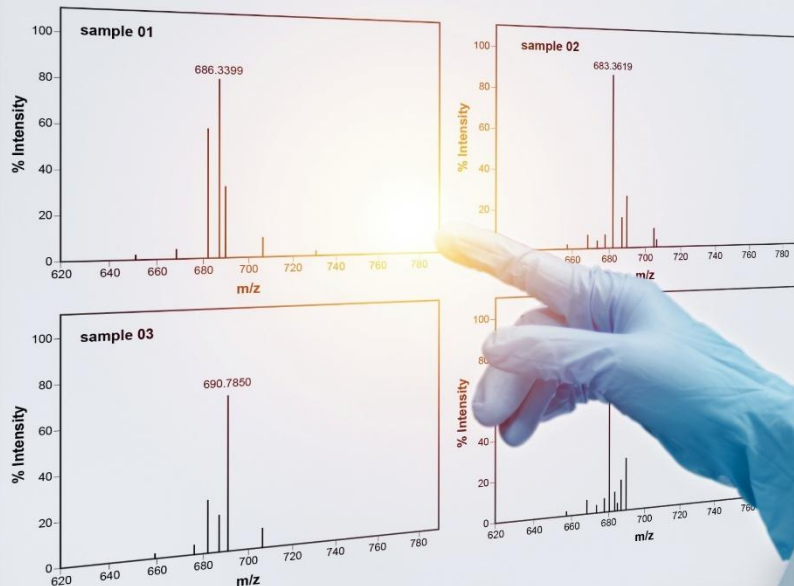
So chlorine's
Mass Number of
35.5 is simply
the average
mass of its two
main isotopes!

- The average mass of the isotopes of a chemical element is correctly known as the *relative atomic mass*.



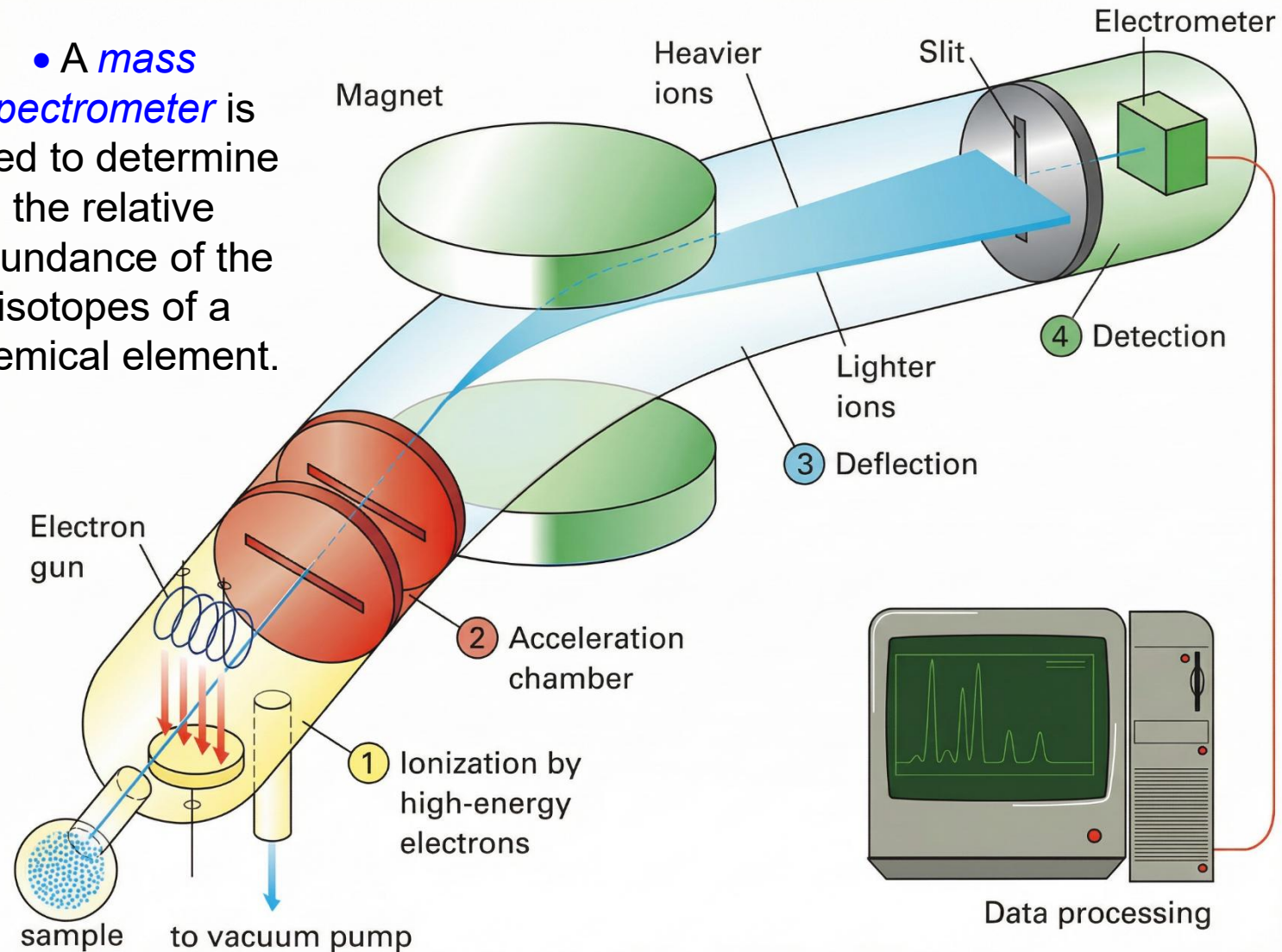
Calculation of Relative Atomic Mass

Mass Spectrometry Analysis Report



Calculation of Relative Atomic Mass

- A *mass spectrometer* is used to determine the relative abundance of the isotopes of a chemical element.

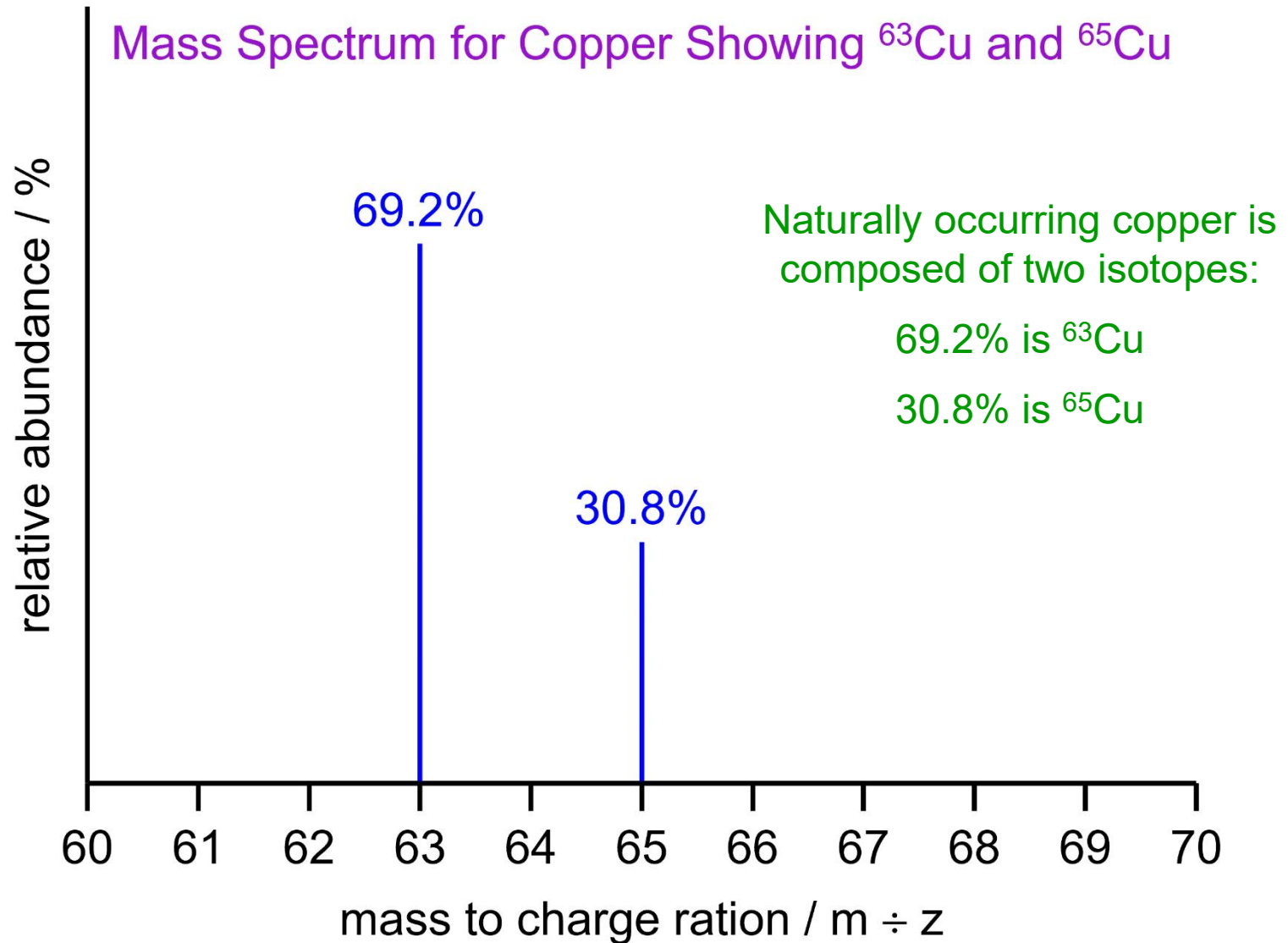


Calculation of Relative Atomic Mass

Definition of Relative Atomic Mass

- *Relative atomic mass* is the average mass of the isotopes of a chemical element compared to $1/12^{\text{th}}$ the mass of a single atom of $^{12}_6\text{C}$.
- Because it is a *ratio*, relative atomic mass is dimensionless *i.e.* it does not have any units.
- The $^{12}_6\text{C}$ isotope has a mass of exactly 12.000. Therefore, $1/12^{\text{th}}$ the mass of a single atom of $^{12}_6\text{C}$ is exactly 1.000. This is known as the *atomic mass unit*.
- **Note:** The ^1_1H isotope is *not* used as the reference used to define relative atomic mass because it actually has a mass of 1.008 and not 1.000!

Calculation of Relative Atomic Mass



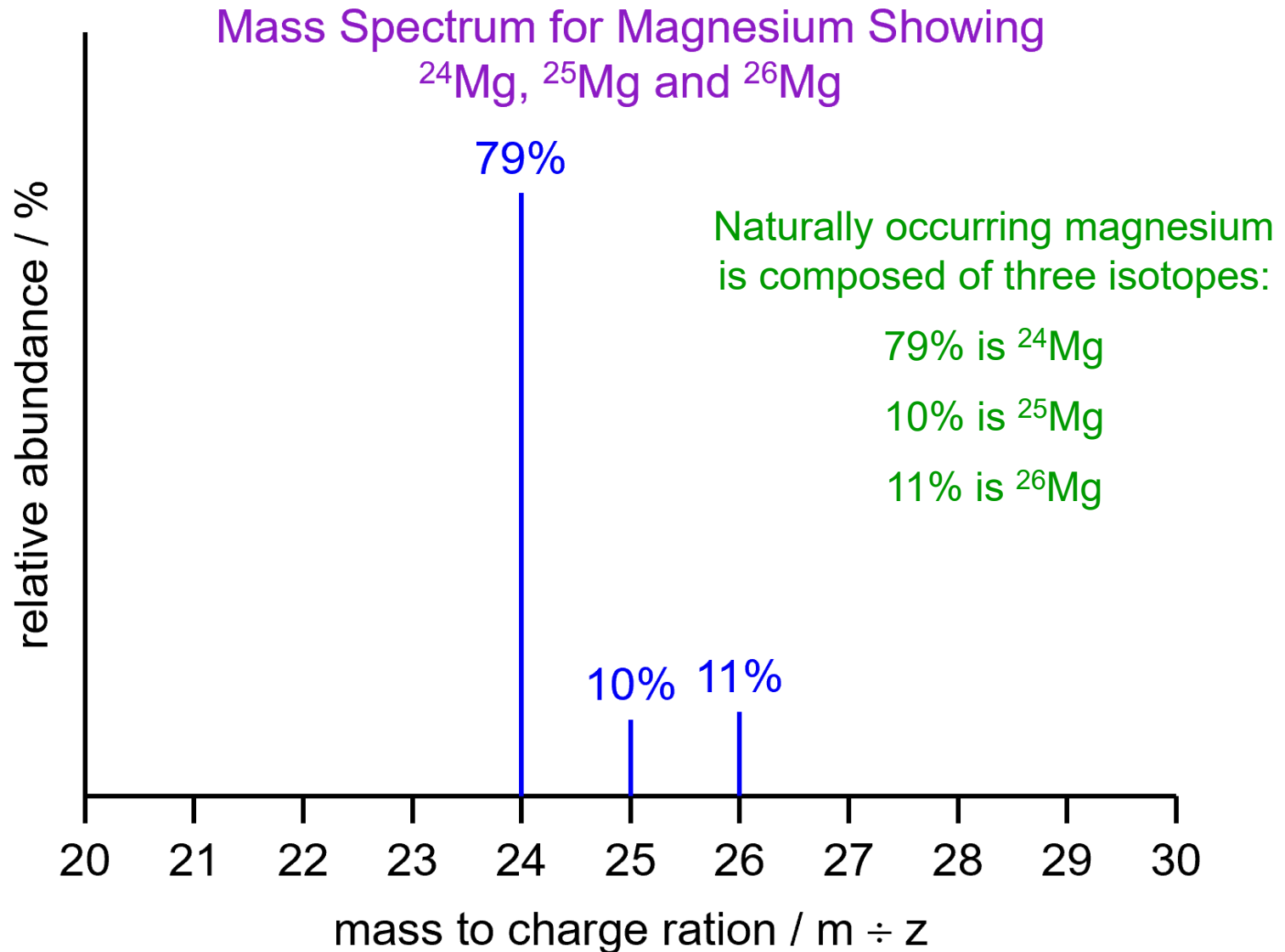
Calculation of Relative Atomic Mass

- Naturally occurring copper has two common isotopes, ${}^{63}_{29}\text{Cu}$ and ${}^{65}_{29}\text{Cu}$.
- ${}^{63}_{29}\text{Cu}$ contains 29 protons and $63 - 29 = 34$ neutrons while ${}^{65}_{29}\text{Cu}$ contains 29 protons and $65 - 29 = 36$ neutrons.
 - The two isotopes of copper do not occur in equal quantities. 69.2% is ${}^{63}_{29}\text{Cu}$ and 30.8% is ${}^{65}_{29}\text{Cu}$.

Calculation of Relative Atomic Mass

- Naturally occurring copper has two common isotopes, $^{63}_{29}\text{Cu}$ and $^{65}_{29}\text{Cu}$.
- $^{63}_{29}\text{Cu}$ contains 29 protons and $63 - 29 = 34$ neutrons while $^{65}_{29}\text{Cu}$ contains 29 protons and $65 - 29 = 36$ neutrons.
 - The two isotopes of copper do not occur in equal quantities. 69.2% is $^{63}_{29}\text{Cu}$ and 30.8% is $^{65}_{29}\text{Cu}$.
- This means that if 100 atoms of copper were sampled at random, 69.2 would be $^{63}_{29}\text{Cu}$ and 30.8 would be $^{65}_{29}\text{Cu}$.
 - The mass of 100 copper atoms would therefore be:
$$(69.2 \times 63) + (30.8 \times 65) = 6361.6.$$
 - The average mass of a single copper atom would therefore be $6361.6 \div 100 = 63.6$ (3 s.f.).

Calculation of Relative Atomic Mass



Calculation of Relative Atomic Mass

- Naturally occurring magnesium has three common isotopes, $^{24}_{12}\text{Mg}$, $^{25}_{12}\text{Mg}$ and $^{26}_{12}\text{Mg}$.
- $^{24}_{12}\text{Mg}$ contains 12 protons and $24 - 12 = 12$ neutrons
 $^{25}_{12}\text{Mg}$ contains 12 protons and $25 - 12 = 13$ neutrons
 $^{26}_{12}\text{Mg}$ contains 12 protons and $26 - 12 = 14$ neutrons
- The three isotopes of magnesium do not occur in equal quantities. 79% is $^{24}_{12}\text{Mg}$, 10% is $^{25}_{12}\text{Mg}$ and 11% is $^{26}_{12}\text{Mg}$.

Calculation of Relative Atomic Mass

- Naturally occurring magnesium has three common isotopes, $^{24}_{12}\text{Mg}$, $^{25}_{12}\text{Mg}$ and $^{26}_{12}\text{Mg}$.
- $^{24}_{12}\text{Mg}$ contains 12 protons and $24 - 12 = 12$ neutrons
 $^{25}_{12}\text{Mg}$ contains 12 protons and $25 - 12 = 13$ neutrons
 $^{26}_{12}\text{Mg}$ contains 12 protons and $26 - 12 = 14$ neutrons
- The three isotopes of magnesium do not occur in equal quantities. 79% is $^{24}_{12}\text{Mg}$, 10% is $^{25}_{12}\text{Mg}$ and 11% is $^{26}_{12}\text{Mg}$.
- This means that if 100 atoms of magnesium were sampled, 79 would be $^{24}_{12}\text{Mg}$, 10 would be $^{25}_{12}\text{Mg}$ and 11 would be $^{26}_{12}\text{Mg}$.
- The mass of 100 magnesium atoms would therefore be:
 $(79 \times 24) + (10 \times 25) + (11 \times 26) = 2432$.
- The average mass of a single magnesium atom would therefore be $2432 \div 100 = 24.3$ (3 s.f.).

Calculation of Relative Abundance

- Boron has two isotopes, ^{10}B and ^{11}B . Given that the relative atomic mass of boron is 10.8, calculate the percentage abundance of ^{10}B and ^{11}B .

Calculation of Relative Abundance

- Boron has two isotopes, ^{10}B and ^{11}B . Given that the relative atomic mass of boron is 10.8, calculate the percentage abundance of ^{10}B and ^{11}B .

- Let the percentage abundance of ^{10}B be x

Then the percentage abundance of ^{11}B is $100 - x$

100 atoms of boron have a total mass of 1080

$$10x + 11(100 - x) = 1080$$

$$10x + 1100 - 11x = 1080$$

$$x = 20$$

Percentage abundance of $^{10}\text{B} = 20.0\%$

Percentage abundance of $^{11}\text{B} = 80.0\%$

Check! $((10 \times 20) + (11 \times 80)) \div 100 = \underline{10.8}$

Calculation of Relative Abundance

- Gallium has two isotopes, ^{69}Ga and ^{71}Ga . Given that the relative atomic mass of gallium is 69.7, calculate the percentage abundance of ^{69}Ga and ^{71}Ga .

Calculation of Relative Abundance

- Gallium has two isotopes, ^{69}Ga and ^{71}Ga . Given that the relative atomic mass of gallium is 69.7 , calculate the percentage abundance of ^{69}Ga and ^{71}Ga .

- Let the percentage abundance of ^{69}Ga be x

Then the percentage abundance of ^{71}Ga is $100 - x$

100 atoms of gallium have a total mass of 6970

$$69x + 71(100 - x) = 6970$$

$$69x + 7100 - 71x = 6970$$

$$x = 65$$

Percentage abundance of $^{69}\text{Ga} = 65.0\%$

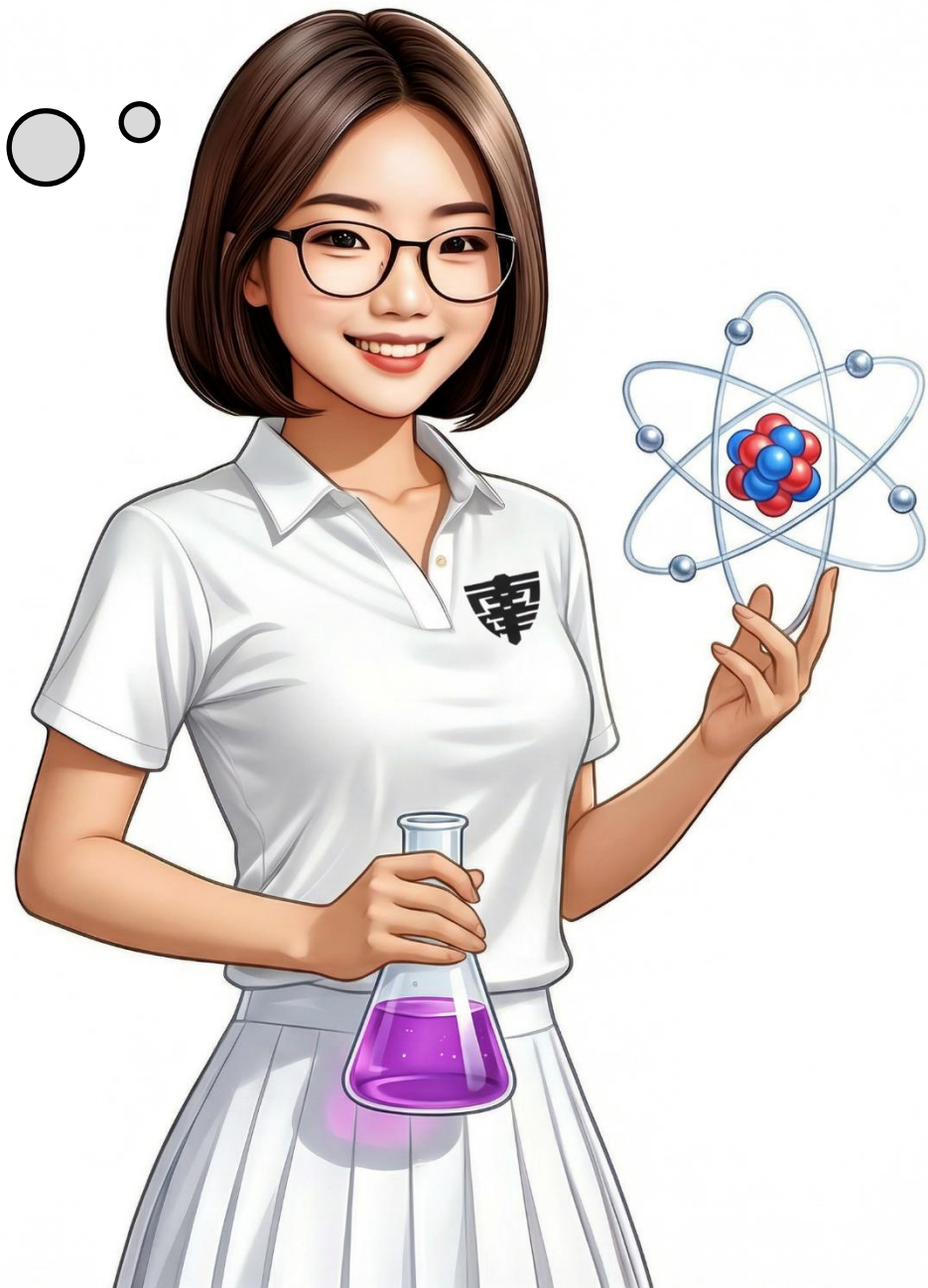
Percentage abundance of $^{71}\text{Ga} = 35.0\%$

Check! $((69 \times 65) + (71 \times 35)) \div 100 = \underline{69.7}$

Why do atoms
react?

How do atoms
react?

- Atoms of the different chemical elements react in order to obtain the *stable electronic configuration* of a *noble gas* (Group 18).



Essential Understanding for Chemistry

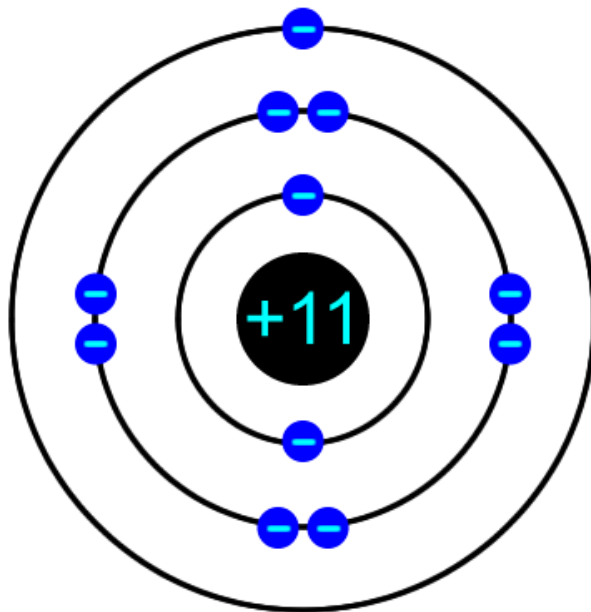
Why do atoms react?

- Atoms react in order to obtain the electronic configuration of a Noble gas (Group 18).

How do atoms react?

- Atoms will *lose* / *gain* / *share* electrons in order to obtain the electronic configuration of a Noble gas (Group 18).

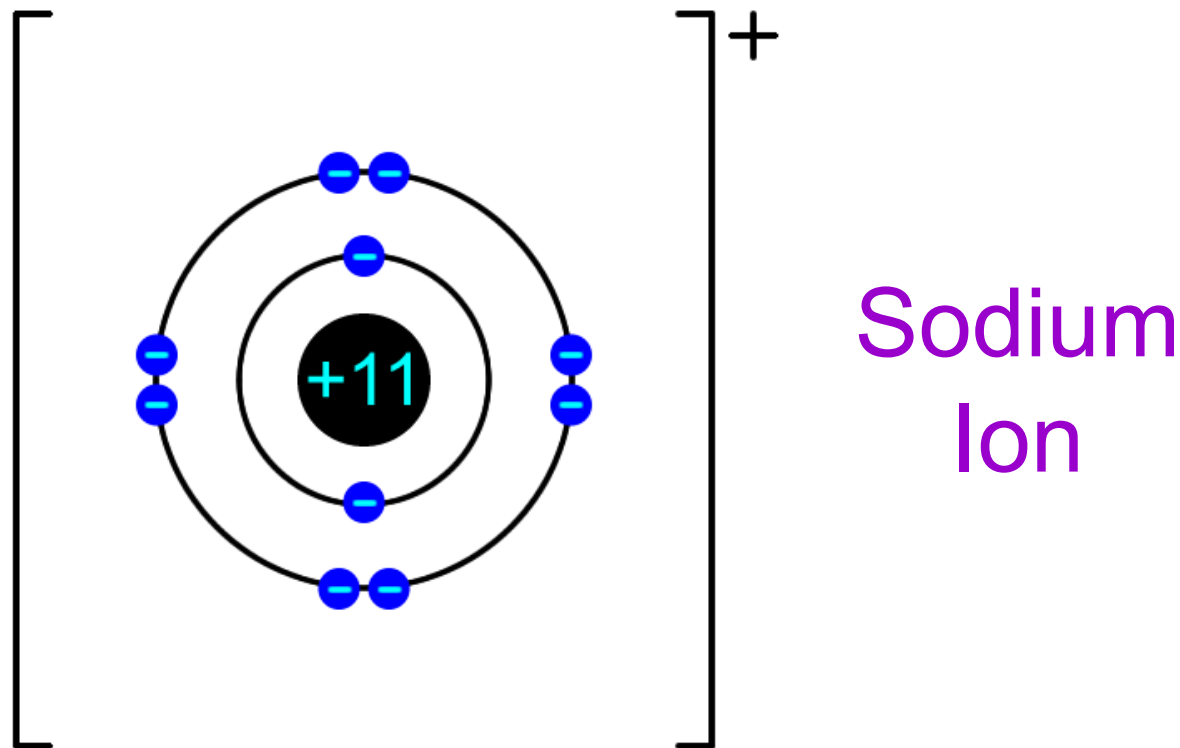
Ions – Anions and Cations



Sodium
Atom

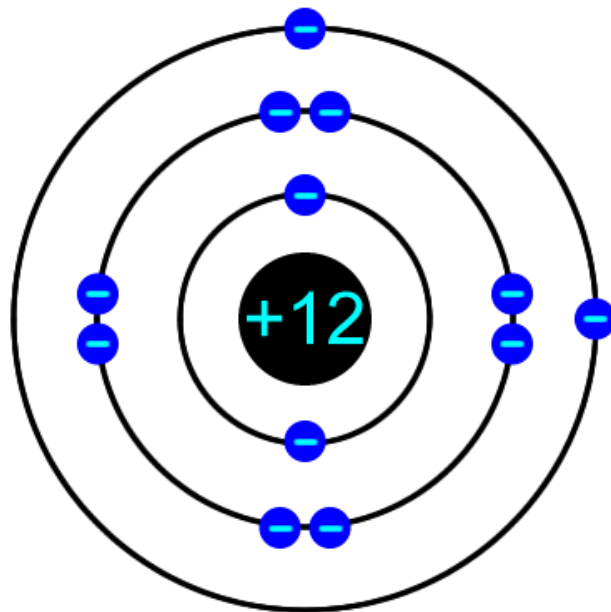
- Atoms of metallic elements *lose* their valence electron(s) in order to obtain the stable electronic configuration of a noble gas.
- If a neutral sodium atom with *11 protons* and *11 electrons* loses its *single* valence electron...

Ions – Anions and Cations



...the resulting particle will have *11 protons* and *10 electrons*. Adding up the charges on all of the protons and all of the electrons $[(+11) + (-10)]$ results in an overall charge on the particle of *+1*. The particle is described as a *cation* or *positive ion*.

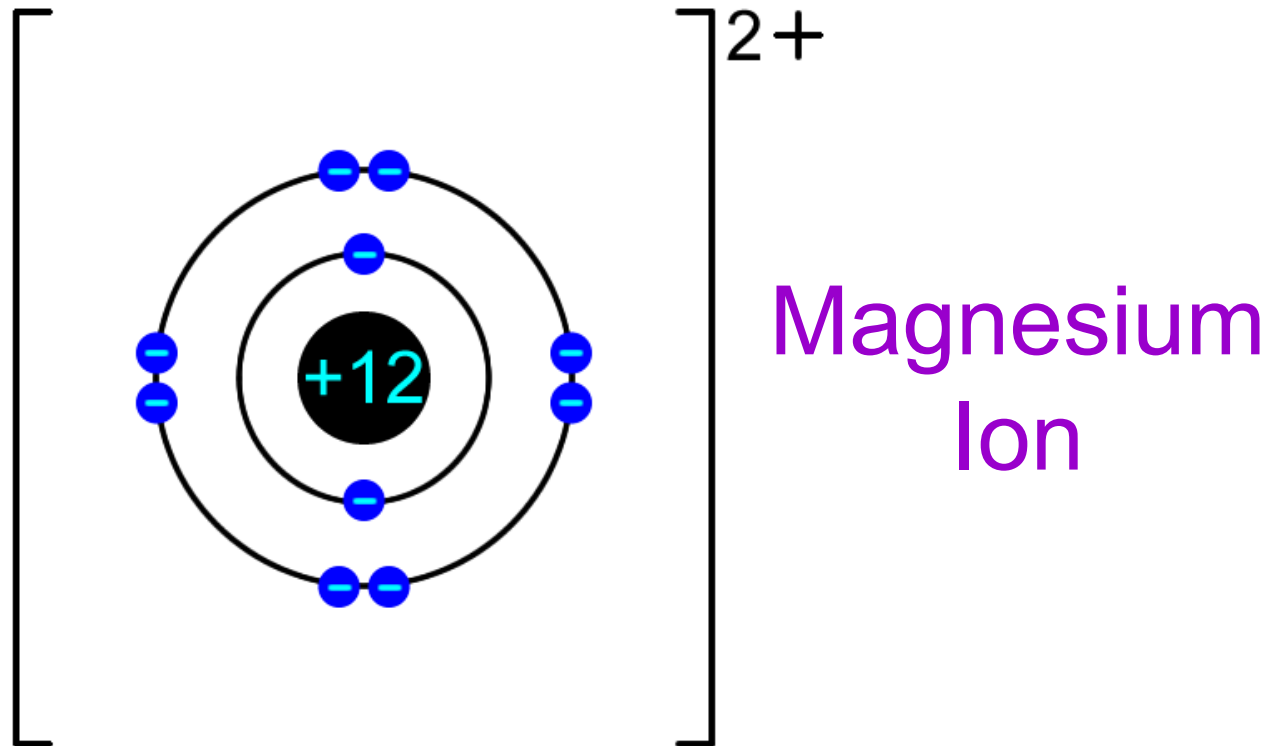
Ions – Anions and Cations



Magnesium
Atom

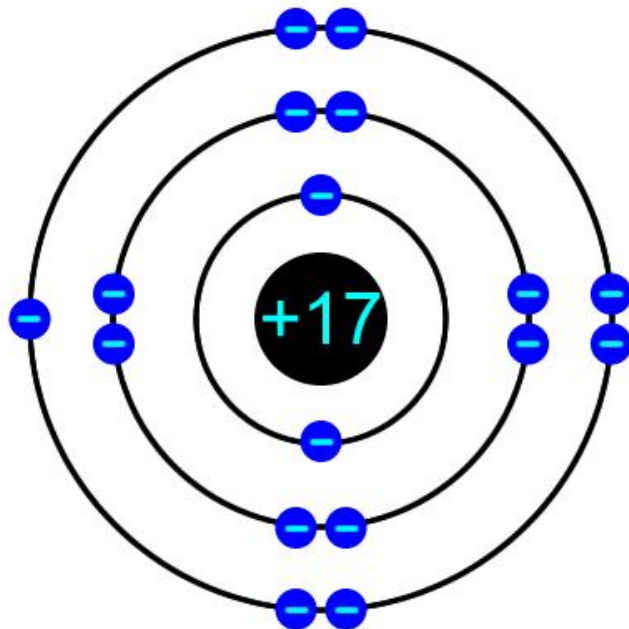
- Atoms of metallic elements *lose* their valence electron(s) in order to obtain the stable electronic configuration of a noble gas.
- If a neutral magnesium atom with *12 protons* and *12 electrons* loses its *two* valence electrons...

Ions – Anions and Cations



...the resulting particle will have *12 protons* and *10 electrons*. Adding up the charges on all of the protons and all of the electrons $[(+12) + (-10)]$ results in an overall charge on the particle of *+2*. The particle is described as a *cation* or *positive ion*.

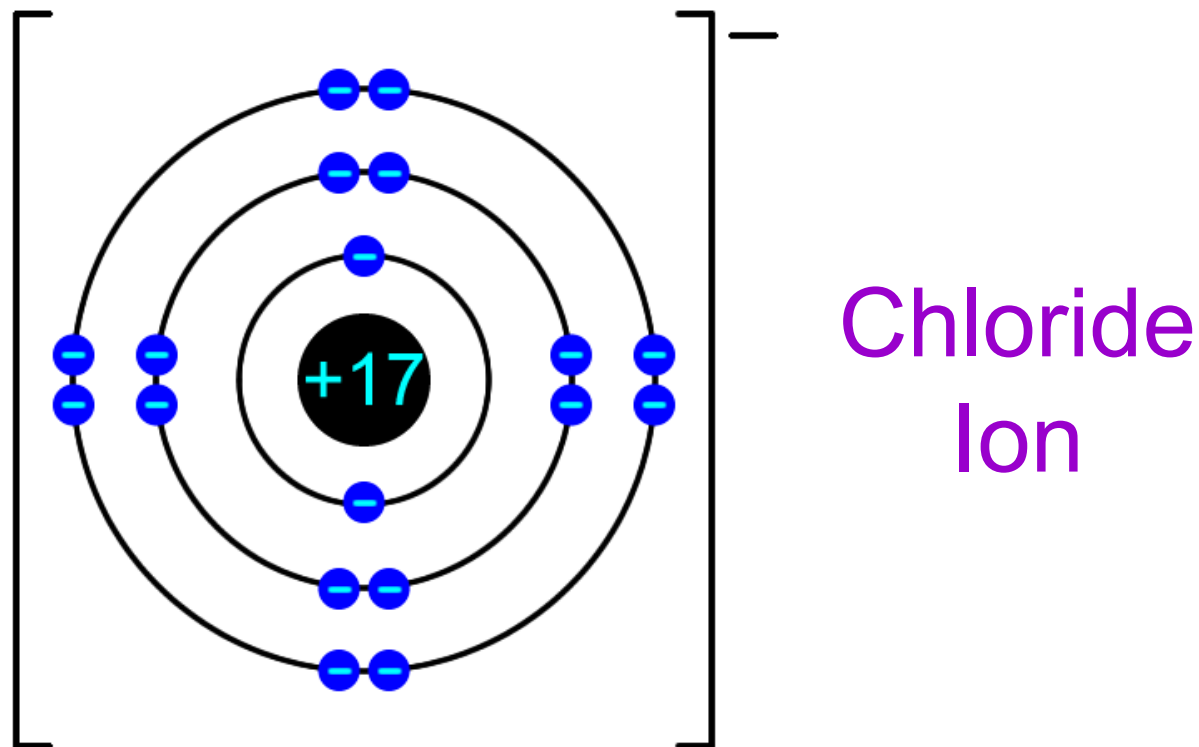
Ions – Anions and Cations



Chlorine
Atom

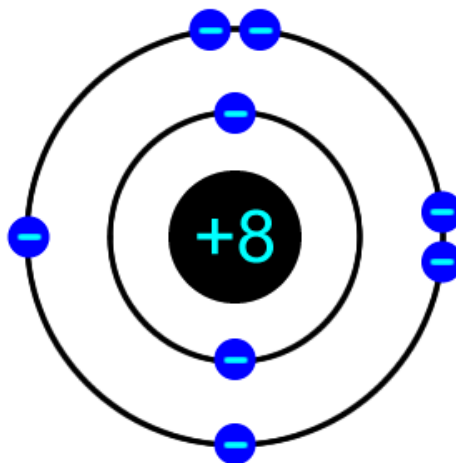
- Atoms of non-metallic elements *gain* electrons in order to obtain the stable electronic configuration of a noble gas.
 - If a neutral chlorine atom with *17 protons* and *17 electrons* were to gain a *single* electron...

Ions – Anions and Cations



...the resulting particle will have *17 protons* and *18 electrons*. Adding up the charges on all of the protons and all of the electrons $[(+17) + (-18)]$ results in an overall charge on the particle of *-1*. The particle is described as an *anion* or *negative ion*.

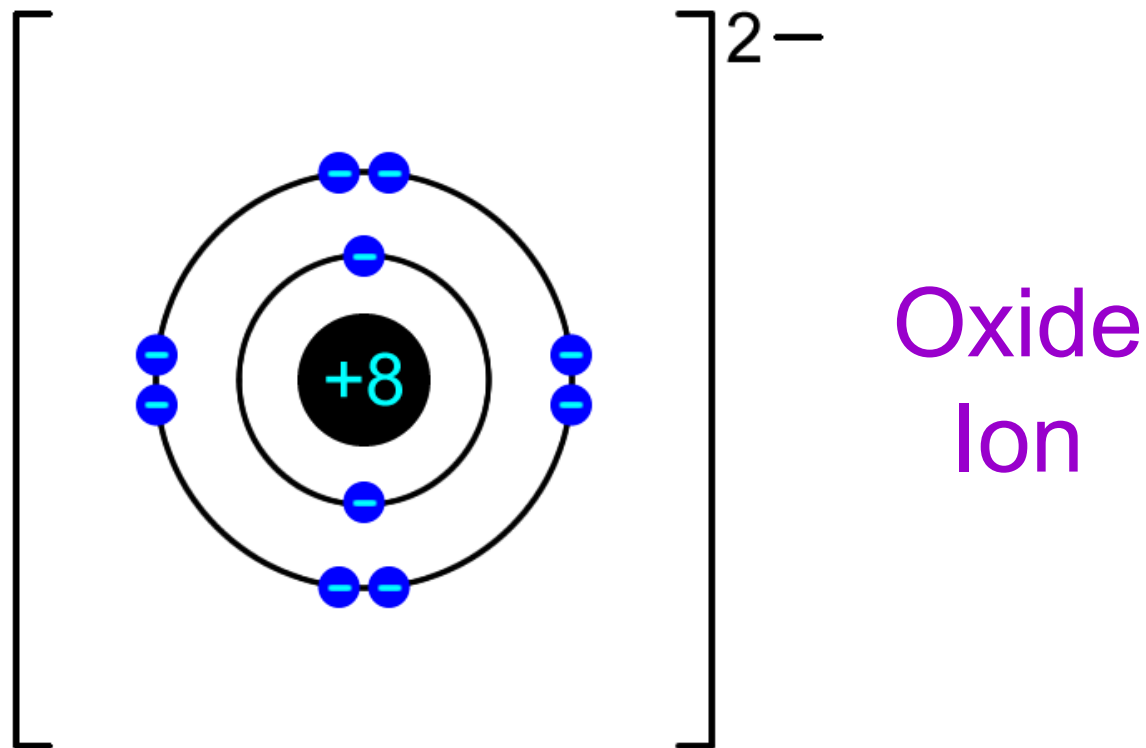
Ions – Anions and Cations



Oxygen
Atom

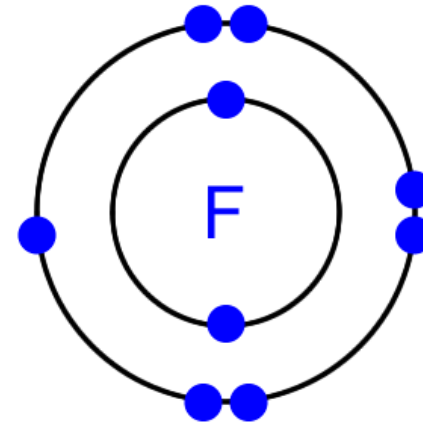
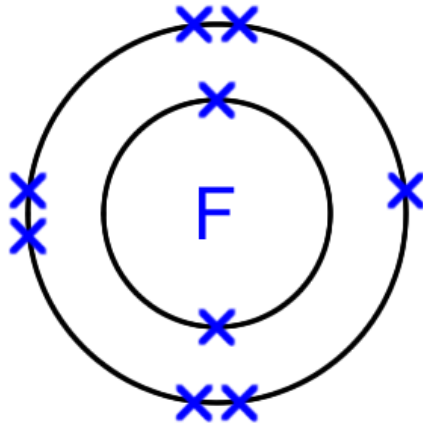
- Atoms of non-metallic elements *gain* electrons in order to obtain the stable electronic configuration of a noble gas.
 - If a neutral oxygen atom with *8 protons* and *8 electrons* were to gain *two* electrons...

Ions – Anions and Cations



...the resulting particle will have *8 protons* and *10 electrons*. Adding up the charges on all of the protons and all of the electrons $[(+8) + (-10)]$ results in an overall charge on the particle of *-2*. The particle is described as an *anion* or *negative ion*.

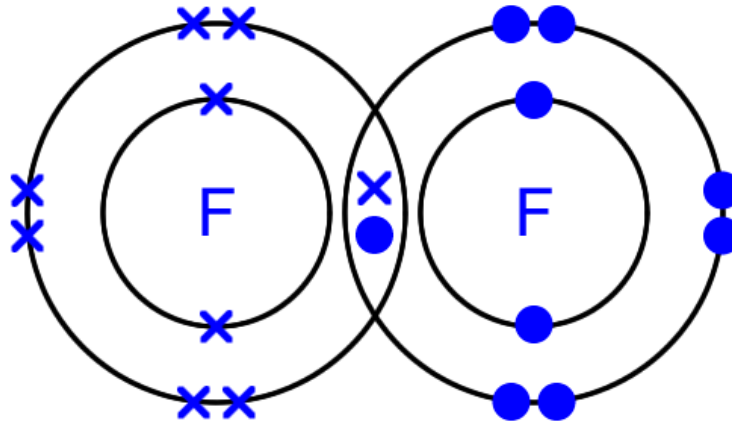
Simple Covalent Molecules



Two fluorine atoms – $2 \times \text{F}$

- Atoms of two or more *non-metallic* elements can also obtain the stable electronic configuration of a noble gas by *sharing* electrons.

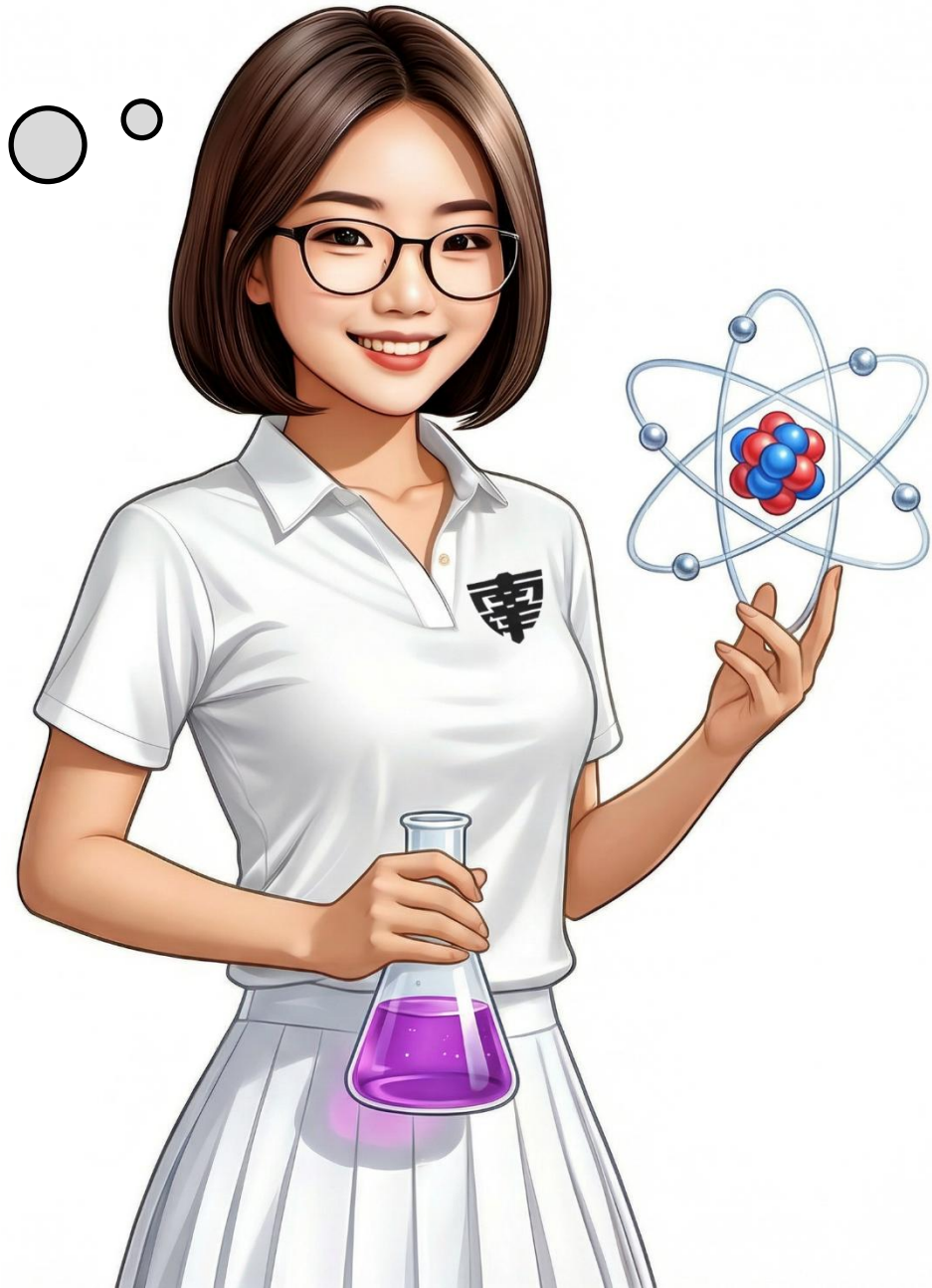
Simple Covalent Molecules



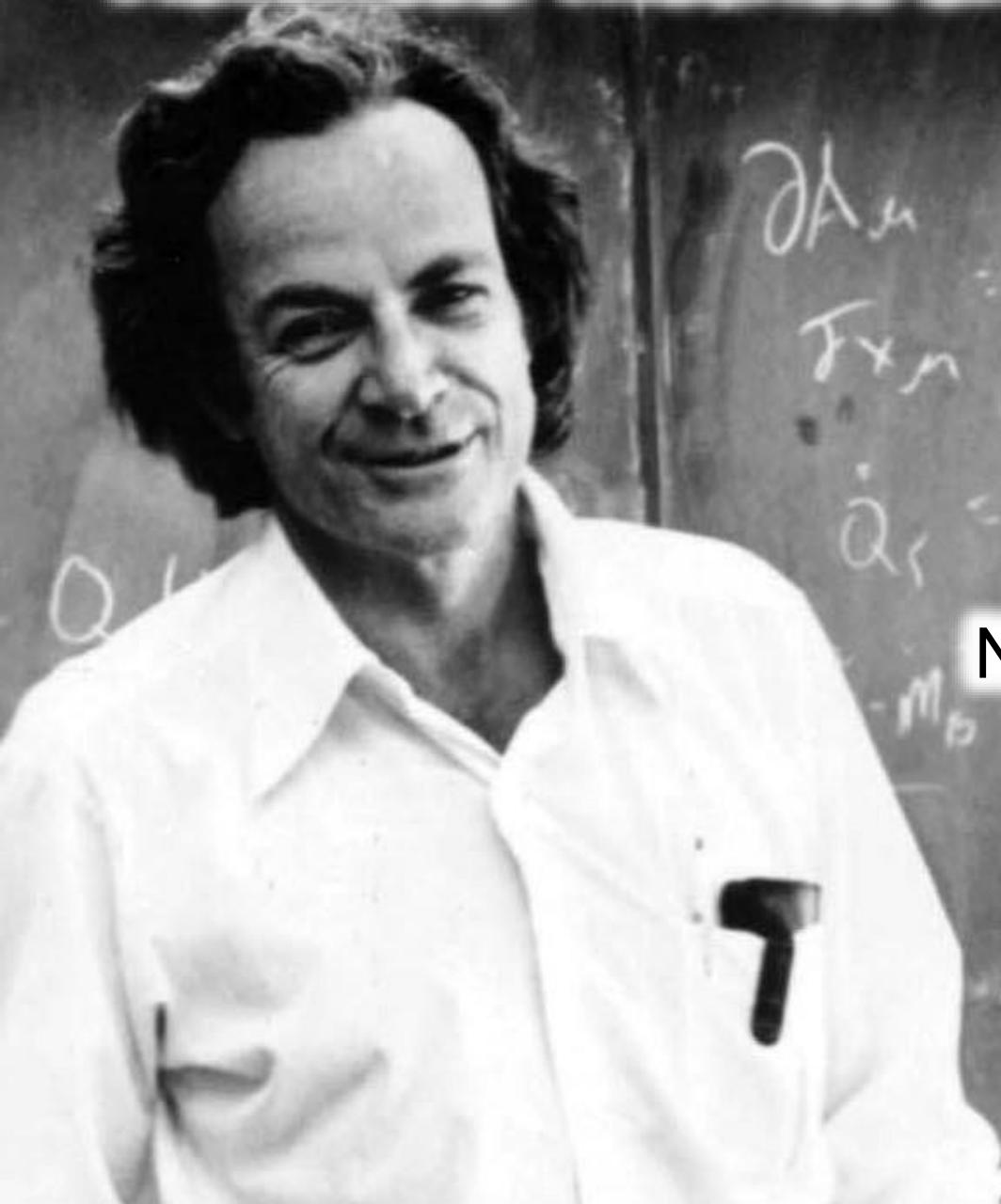
A single molecule of fluorine – F_2

- A pair of electrons shared between two atoms is referred to as a *covalent bond*.
- A small number of atoms of non-metallic elements that are held together by covalent bonds is referred to as a *simple covalent molecule*.

What are the
more advanced
theories of
atomic
structure?



Advanced Theories of Atomic Structure



- Richard Feynman, 1918 – 1988. Winner of the 1965 Nobel Prize in Physics.

Advanced Theories of Atomic Structure

- The modern scientific understanding of atomic structure – in which electrons are assumed to behave like *waves*, and orbit the nucleus of the atom in *atomic orbitals* – is based on *quantum mechanics*.
- Quantum mechanics is the branch of physics that deals with mathematical descriptions of how subatomic particles behave and interact.
 - Richard Feynman introduced volume III of *The Feynman Lectures on Physics* with the words: “*I think I can safely say that nobody understands quantum mechanics*”.

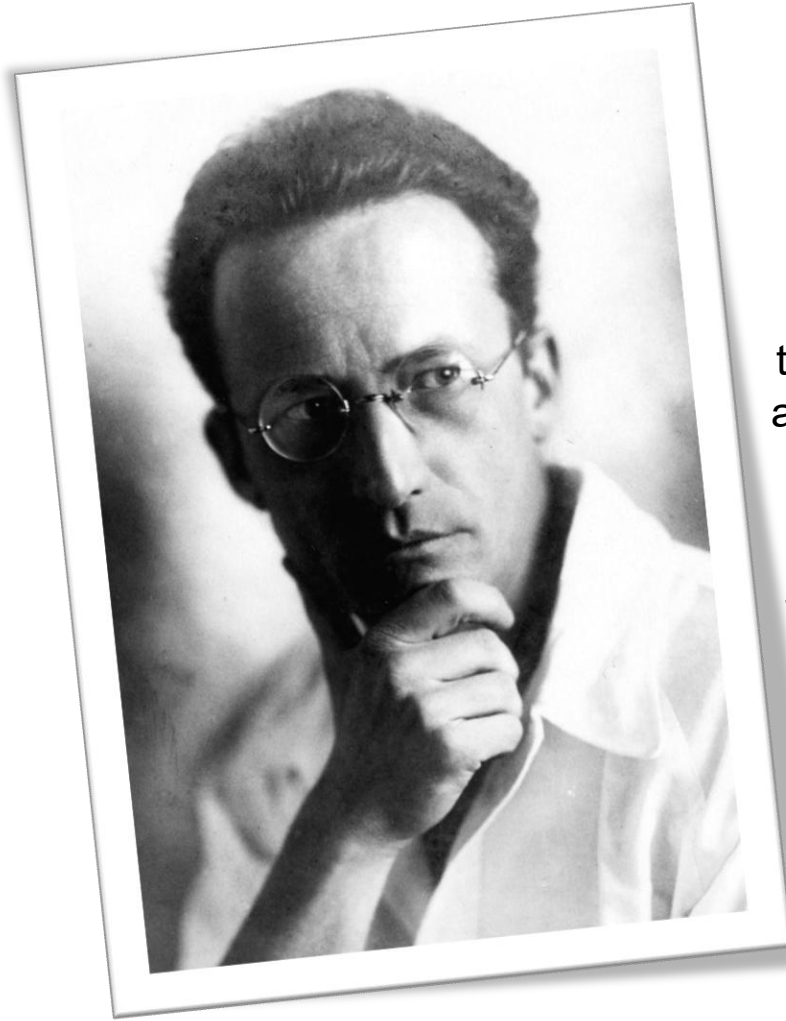
Advanced Theories of Atomic Structure

- Fortunately, following the rules of quantum mechanics is far simpler than trying to visualise what they actually mean. The ability to follow through the consequences of a particular set of assumptions carefully, without getting too hung up on the philosophical implications, is one of the most important skills a scientist can learn.
- When deriving theories related to quantum mechanics, scientists set out their initial assumptions and compute their consequences. If they arrive at a set of predictions that agree with their observations of the natural world around them, then they accept the theory as good.

Advanced Theories of Atomic Structure

- Many problems in quantum mechanics are far too difficult to solve in a single mental leap, and deep understanding rarely emerges in *eureka* moments.
 - The trick is for scientists to make sure that they understand each little step and, after a sufficient number of steps, the bigger picture starts to emerge. If this is not the case, then the scientists need to go back to the drawing board and start to derive a new theory.
- This is true for a scientific understanding of the atomic orbital structure of the atom. It should be attempted one step at-a-time until the big picture of how electrons orbit the nucleus of the atom starts to emerge.

Advanced Theories of Atomic Structure

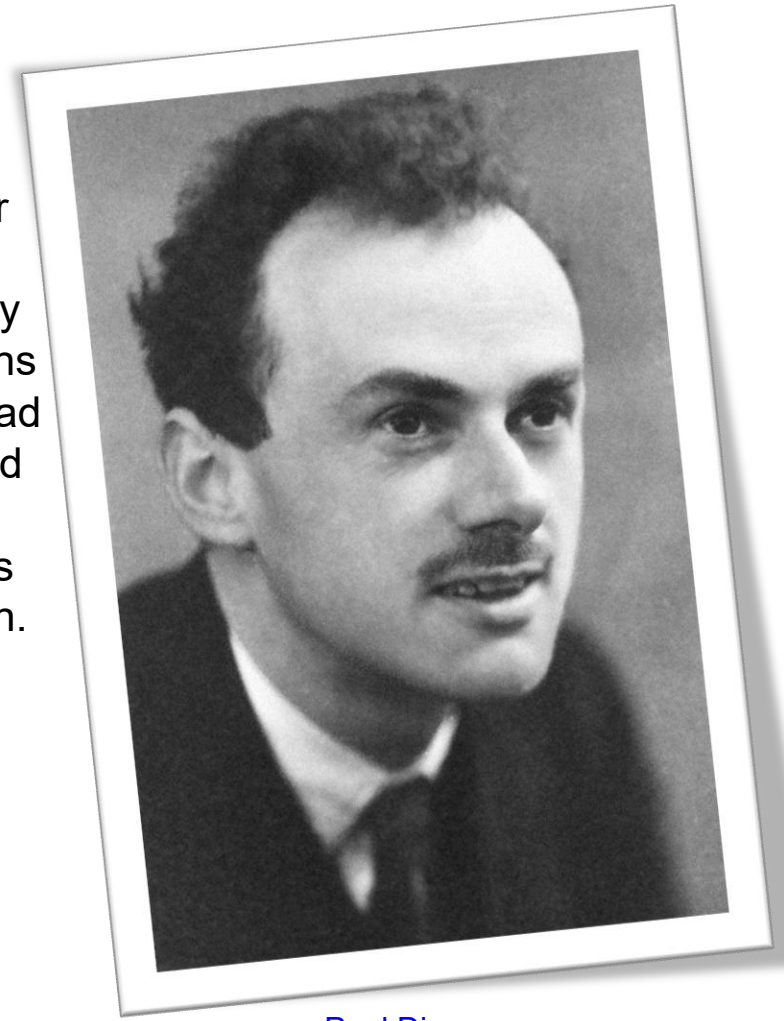


Erwin Schrödinger

1887 – 1961

Awarded the Nobel Prize for
Physics in 1933.

- Schrödinger and Dirac mathematically treated electrons as *waves* instead of *particles* and formulated Schrödinger's Wave Equation.



Paul Dirac

1902 – 1984

Awarded the Nobel Prize for
Physics in 1933.

Advanced Theories of Atomic Structure

Schrödinger's Wave Equation

- In 1924, Louis de Broglie made the bold suggestion that *electrons* may have the properties of *waves* as well as the properties of *particles*.
- Schrödinger and Dirac mathematically treated *electrons* as *waves* instead of *particles* and formulated Schrödinger's Wave Equation.

$$\frac{\delta^2 \Psi}{\delta x^2} + \frac{\delta^2 \Psi}{\delta y^2} + \frac{\delta^2 \Psi}{\delta z^2} + \frac{8m\pi^2}{h^2} (E - V)\Psi = 0$$

- Graphical solutions for this complex equation give rise to *atomic orbitals*.

Advanced Theories of Atomic Structure

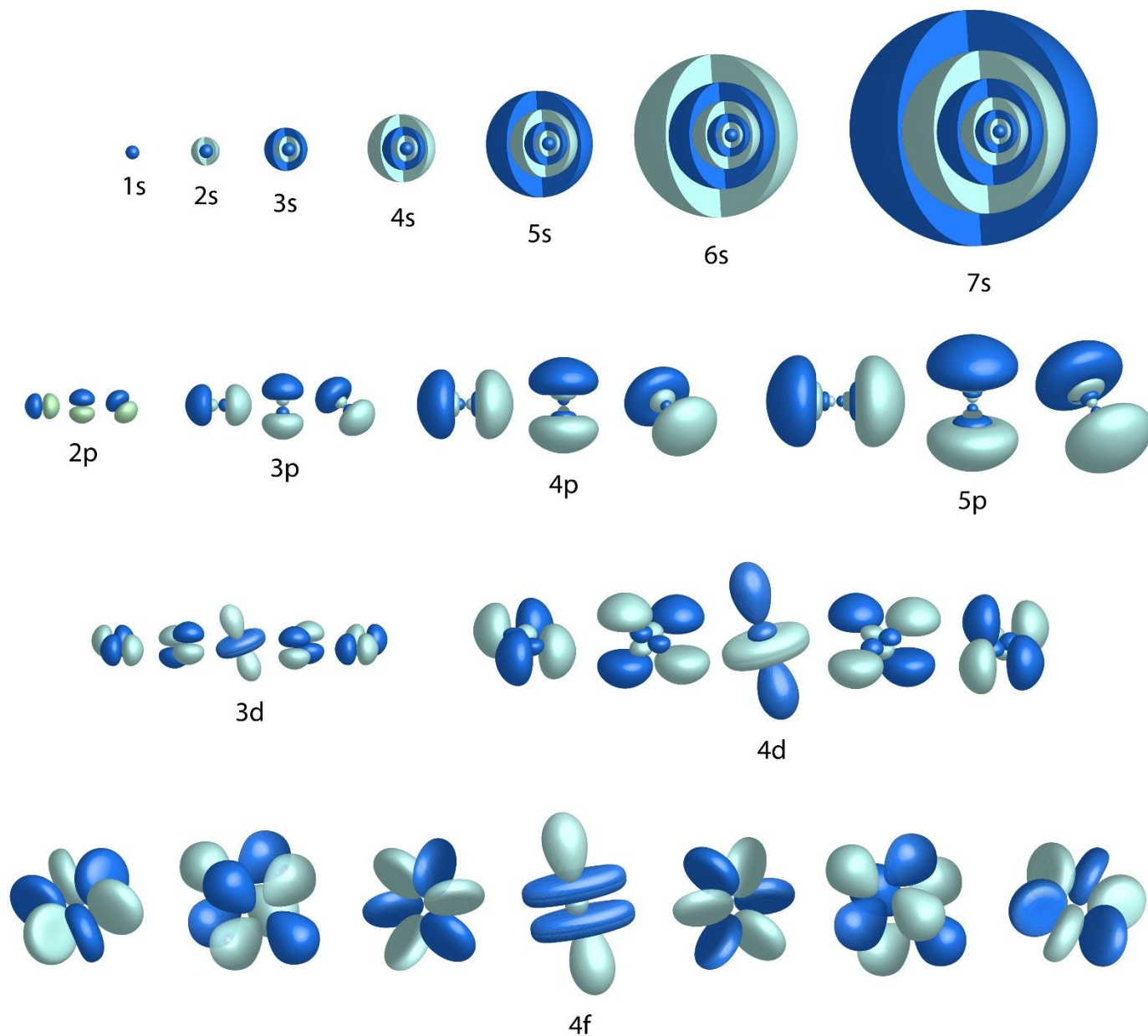
Atomic Orbitals

- An atomic orbital is the volume of space around the nucleus of an atom in which there is a high probability (*95%*) of finding an electron.
- An atomic orbital can hold a maximum number of *two* electrons.
 - The location of an electron in an atom (*i.e.* which atomic orbital it belongs to) is given by *four electron quantum numbers*.
- *Pauli's Exclusion Principle* states that no two electrons in the same atom can have the same set of electron quantum numbers.

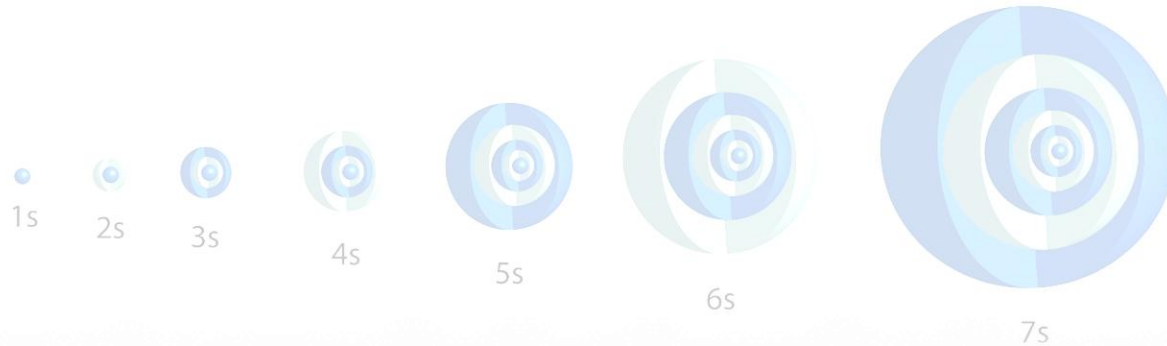
Advanced Theories of Atomic Structure

- In simple terms, the four electron quantum numbers are:
 - **First (n)**: Principle quantum number – the principle quantum shell that the electron occupies.
 - **Second (l)**: The sub-shell within the quantum shell that the electron occupies, e.g. *s-orbitals*, *p-orbitals* or *d-orbitals*.
 - **Third (m)**: The orbital within the sub-shell that the electron occupies, e.g. *p-orbitals* are always arranged in groups of three, so if an electron occupies a *p-orbital*, this electron quantum number states exactly which one, the *p_x-orbital*, the *p_y-orbital* or the *p_z-orbital*.
 - **Fourth (s)**: For two electrons to occupy exactly the same orbital, they must have *opposite spin*. This electron quantum number states whether the electron has a spin of $+1/2$ or $-1/2$.

Advanced Theories of Atomic Structure



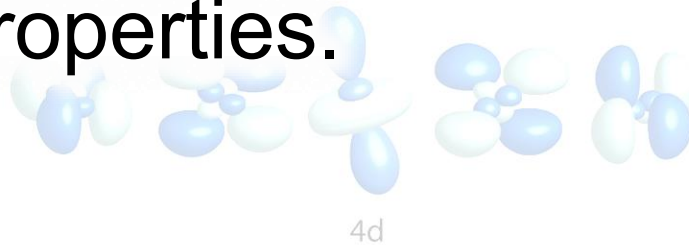
Advanced Theories of Atomic Structure



- Different numerical values for the various electron quantum numbers give rise to orbitals with different shapes and different properties.



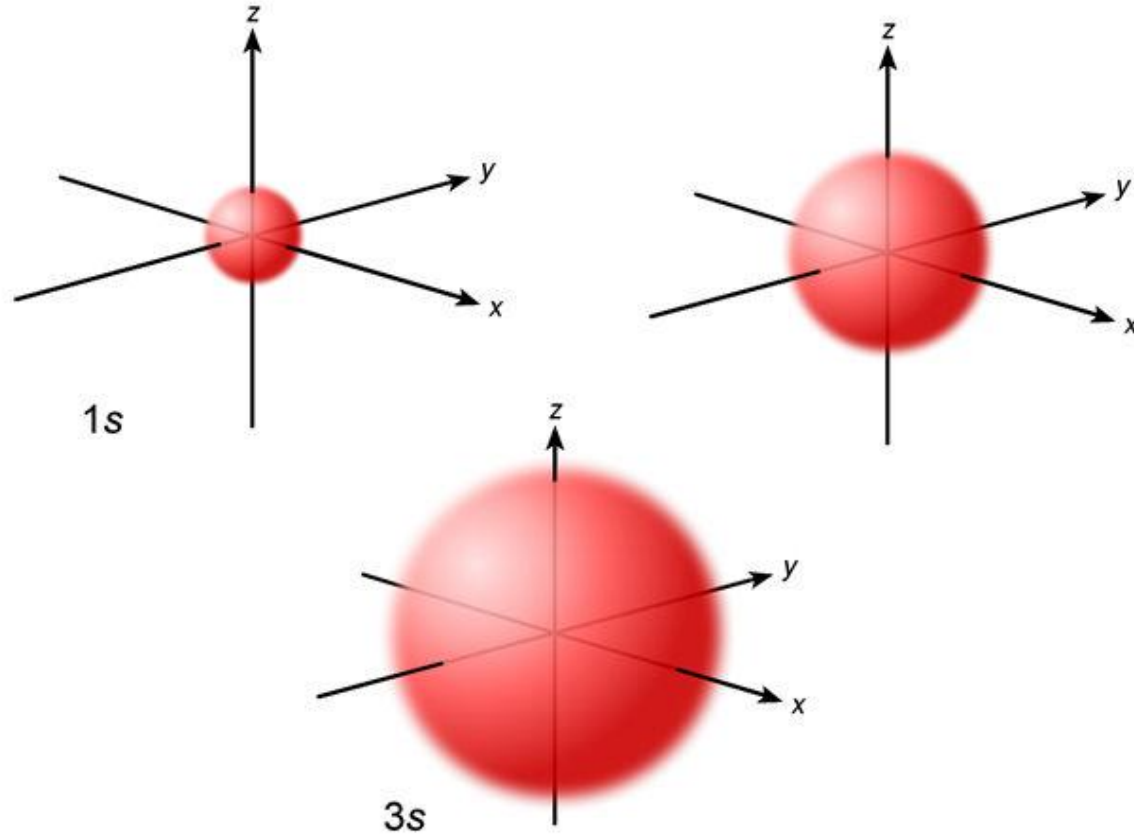
3d



4d

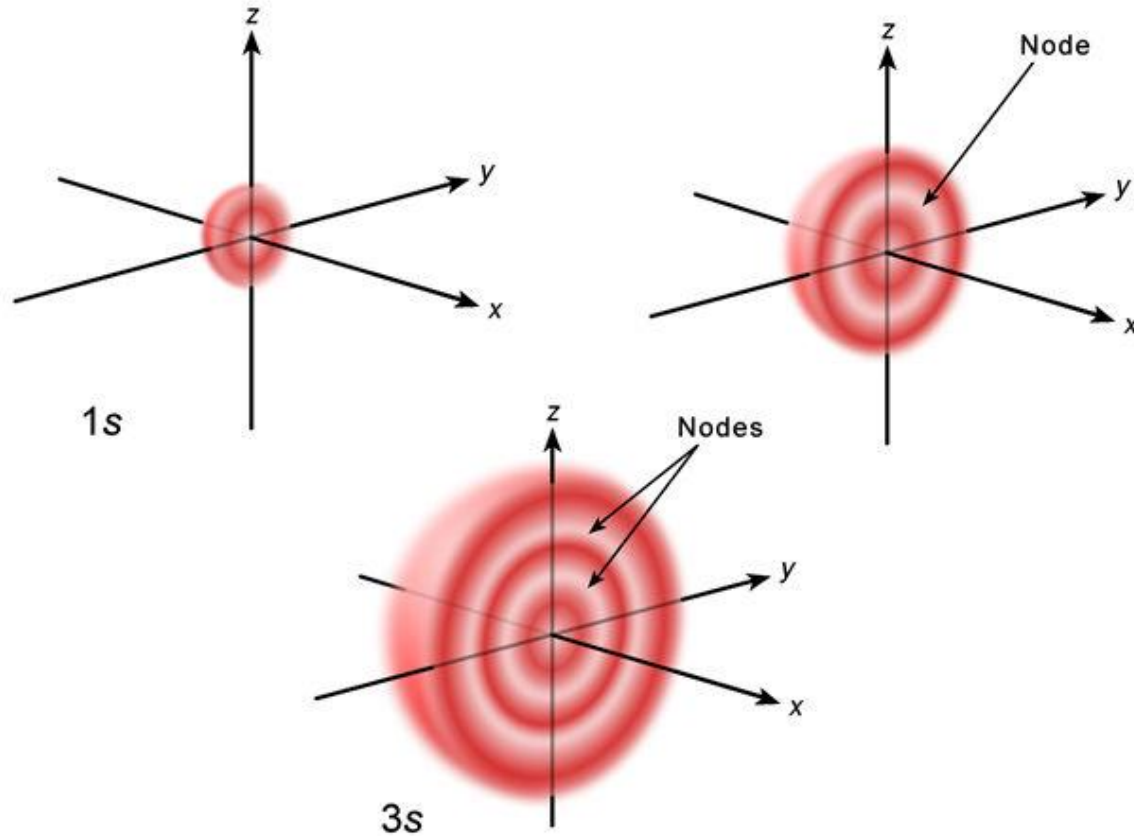


Advanced Theories of Atomic Structure



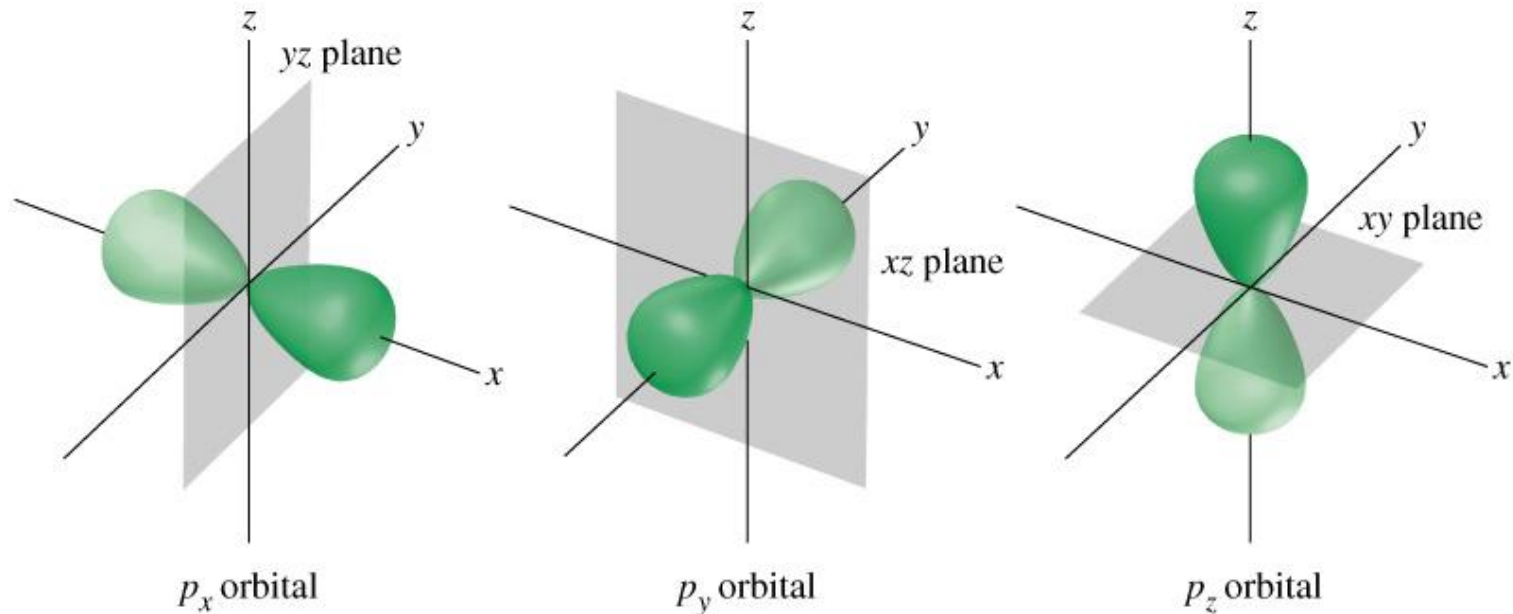
- A graphical solution to Schrödinger's Wave Equation – *s-orbitals*.

Advanced Theories of Atomic Structure



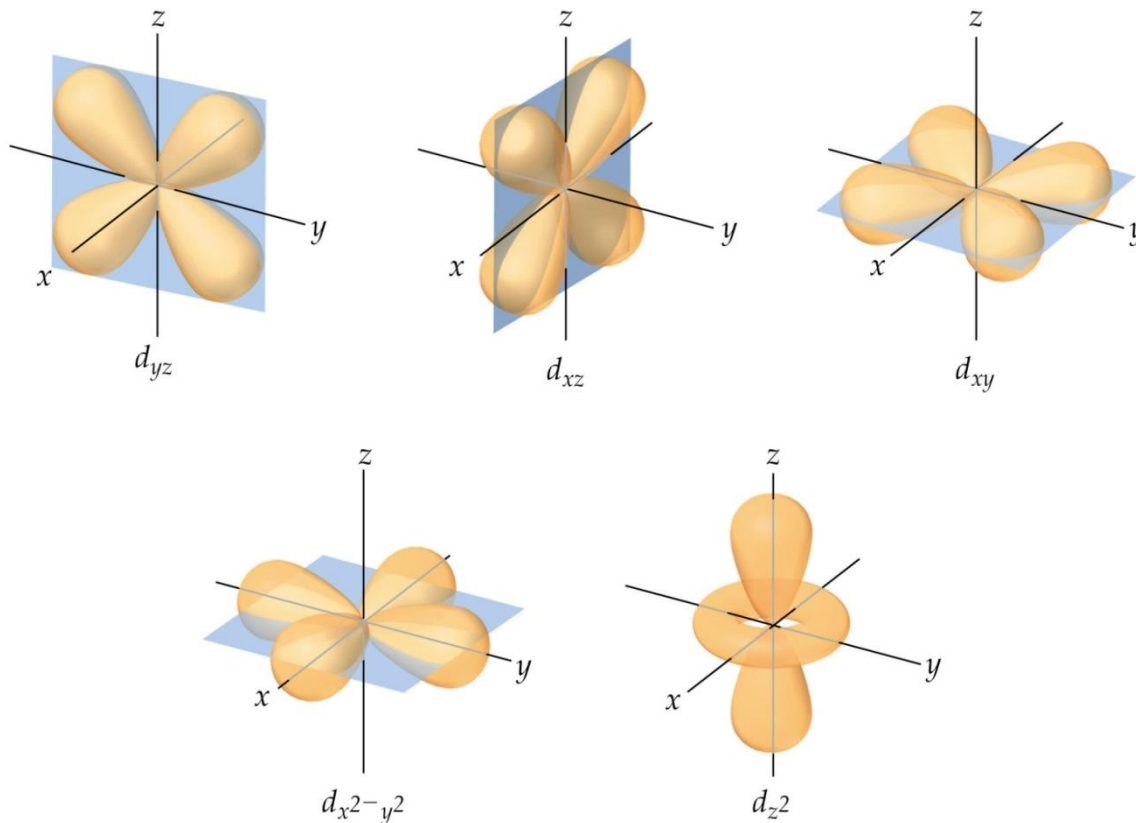
- A graphical solution to Schrödinger's Wave Equation – *s-orbitals*.

Advanced Theories of Atomic Structure



- A graphical solution to Schrödinger's Wave Equation – *p-orbitals*. The orbitals have been drawn separately for clarity. In reality, it is assumed that the three *p*-orbitals are superimposed on top of each other.













































Advanced Theories of Atomic Structure



- A graphical solution to Schrödinger's Wave Equation – *d-orbitals*. The orbitals have been drawn separately for clarity. In reality, it is assumed that the five *d*-orbitals are superimposed on top of each other.

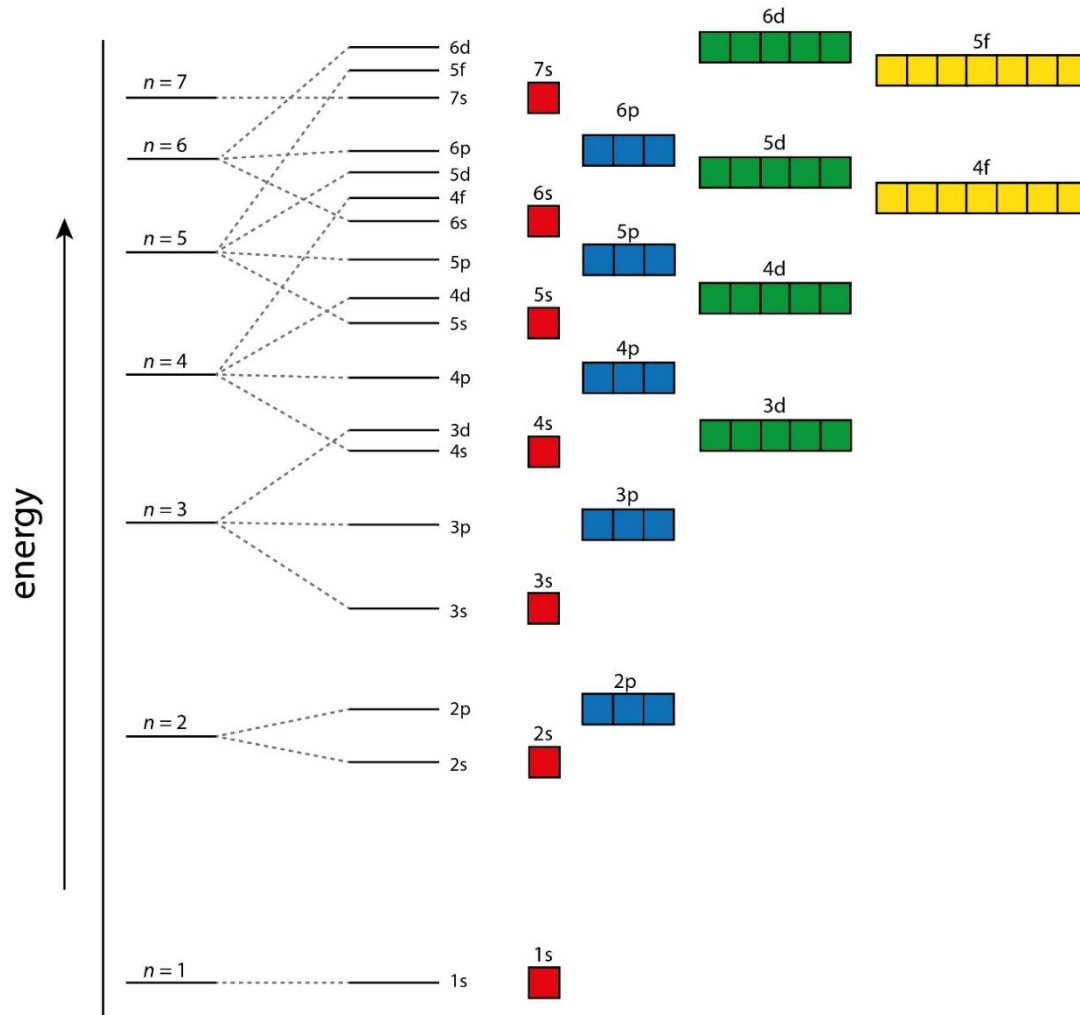
Advanced Theories of Atomic Structure

Shapes and Occurrence of Atomic Orbitals

l	0	1			2					3						
m_l	0	-1	0	1	-2	-1	0	1	2	-3	-2	-1	0	1	2	3
n	s	p_x	p_y	p_z	d_{xy}	d_{xz}	d_{z^2}	d_{yz}	$d_{x^2-y^2}$	$f_{x(x^2-3y^2)}$	f_{xz^2}	f_{yz^2}	f_{z^2}	f_{yz^2}	f_{z^2}	$f_{y(3x^2-y^2)}$
1																
2																
3																
4																
5																
6																
7																

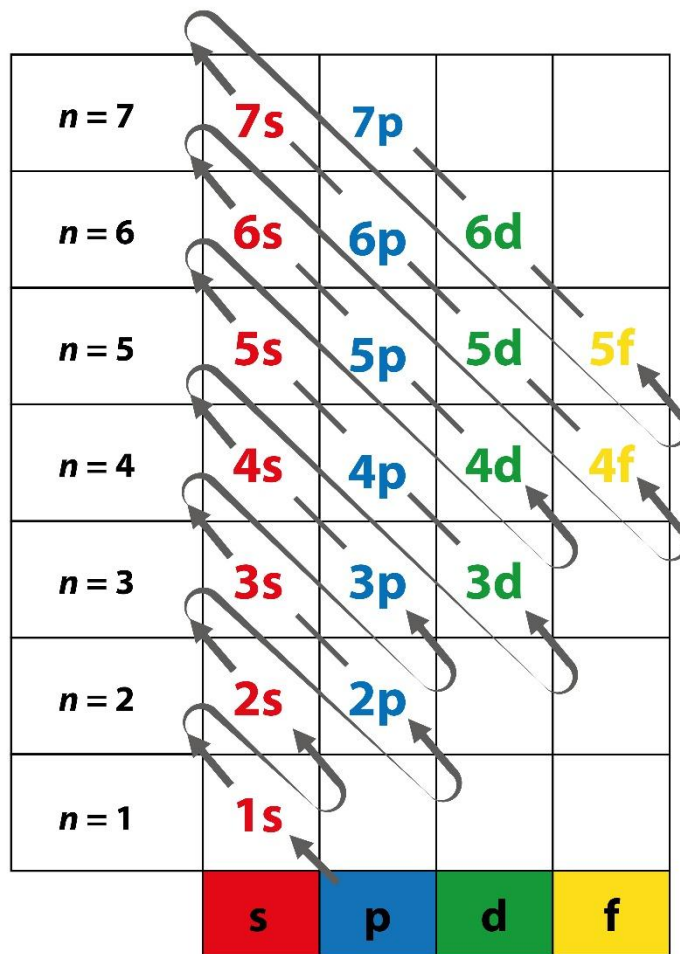
- Diagram showing the various orbitals that arise from different values of the electron quantum numbers n , l and m .

Advanced Theories of Atomic Structure



- Diagram showing the energy levels of the various orbitals.
Note: Orbitals fill from the lowest energy to the highest energy.

Advanced Theories of Atomic Structure



- Diagram showing the order in which atomic orbitals fill-up with electrons.

Advanced Theories of Atomic Structure

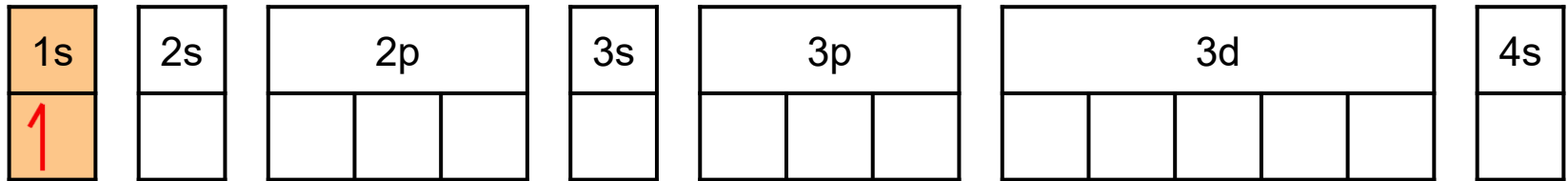
Rules for Filling Atomic Orbitals

- Heisenberg's Uncertainty Principle – It is not possible to determine both the position and the momentum of an electron at the same time. This gives rise to the idea that an electron's position in an atom is *uncertain*, and therefore scientists can only identify where there is the *highest probability* of finding an electron – which is how the atomic orbital is defined.

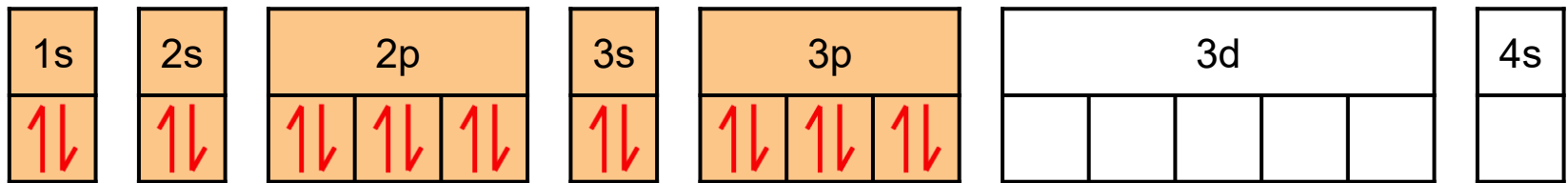
Advanced Theories of Atomic Structure

Rules for Filling Atomic Orbitals

- The Aufbau Principle – Electrons fill-up atomic orbitals from the lowest energy to the highest energy. Left undisturbed, objects will tend to their lowest possible energy.



→ Low Energy →



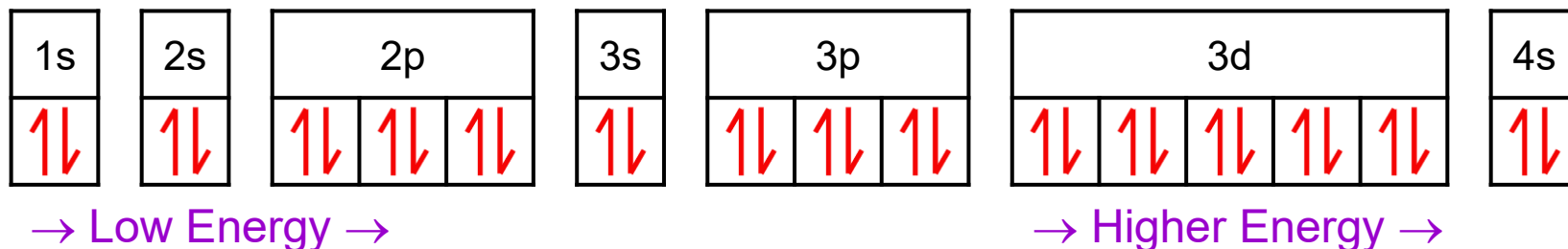
→ Low Energy →

→ Higher Energy →

Advanced Theories of Atomic Structure

Rules for Filling Atomic Orbitals

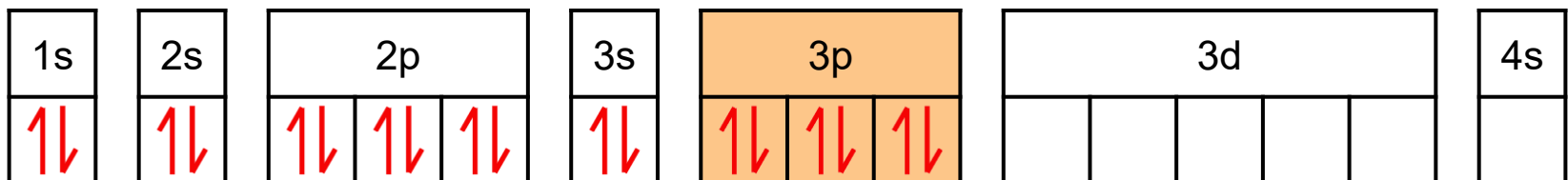
- **Pauli's Exclusion Principle** – No two electrons within the same atom can have the same values for all four quantum numbers, *i.e.* every electron in the same atom must have a unique combination of quantum numbers. As a consequence, electrons in the same orbital must spin in opposite directions. In atomic orbital diagrams, the spin quantum number is represented by an arrow (\uparrow or \downarrow). Two arrows pointing in opposite directions represent two electrons with opposite spin (\uparrow and \downarrow).



Advanced Theories of Atomic Structure

Rules for Filling Atomic Orbitals

- Hund's Rule of Maximum Multiplicity – When placed in atomic orbitals of equal energy, electrons will remain unpaired. Electrons carry a charge of -1 . There will be an electrostatic force of repulsion between electrons in the same orbital. Placing electrons in different atomic orbitals of the same energy will reduce the electrostatic force of repulsion between the electrons and make the system more stable.



→ Low Energy →

Advanced Theories of Atomic Structure

Shapes and Occurrence of Atomic Orbitals

- Each principle quantum shell is divided into one or more sub-shells.

Principle Quantum Shell (n)	Sub-shell (l)	Maximum Number of Electrons
1	1s	2
2	2s, 2p	8
3	3s, 3p, 3d	18
4	4s, 4p, 4d, 4f	32

Advanced Theories of Atomic Structure

Shapes and Occurrence of Atomic Orbitals

- There are four sub-shells, arranged in increasing energy $s \rightarrow p \rightarrow d \rightarrow f$. Each sub-shell holds a different number of electrons

Principle Quantum Shell (n)	Sub-shell (l)	Maximum Number of Electrons
1	1s	2
2	2s, 2p	8
3	3s, 3p, 3d	18
4	4s, 4p, 4d, 4f	32

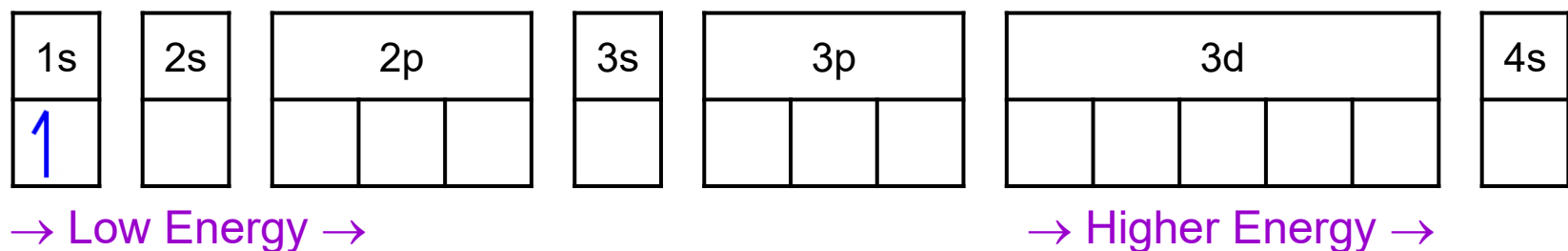
Advanced Theories of Atomic Structure

Shapes and Occurrence of Atomic Orbitals

Orbital	Shape	Occurrence
s (<i>sharp</i>)	spherical	1 in every principle level
p (<i>principle</i>)	dumb-bell or hour glass	3 in every level from 2 onwards
d (<i>diffuse</i>)	complex and various	5 in every level from 3 onwards
f (<i>fundamental</i>)	complex and various	7 in every level from 4 onwards

Advanced Theories of Atomic Structure

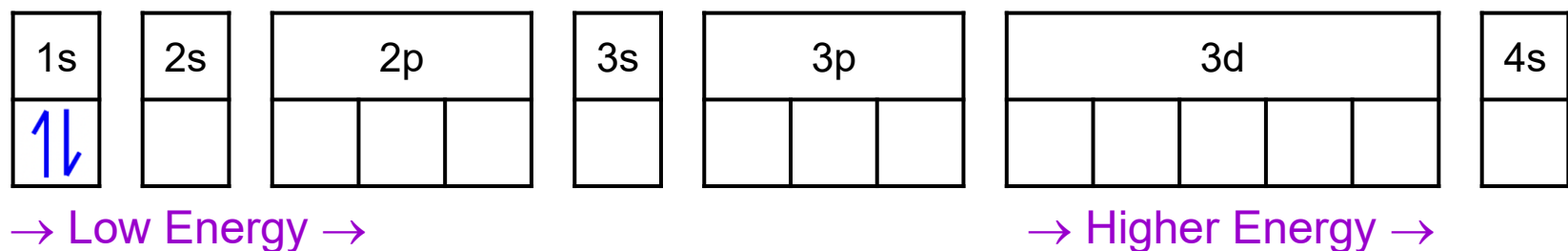
- Atomic Number: 1
- Name: Hydrogen
- Symbol: H
- Electronic Configuration: $1s^1$



- Electrons are represented by *arrows* (\uparrow and \downarrow) which fill atomic orbitals that are represented by *boxes*.

Advanced Theories of Atomic Structure

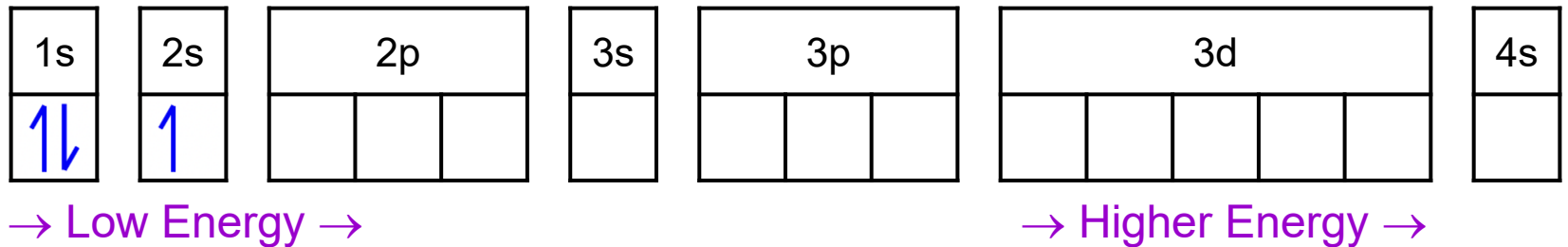
- Atomic Number: 2
- Name: Helium
- Symbol: He
- Electronic Configuration: $1s^2$



- Electrons are represented by *arrows* (\uparrow and \downarrow) which fill atomic orbitals that are represented by *boxes*.

Advanced Theories of Atomic Structure

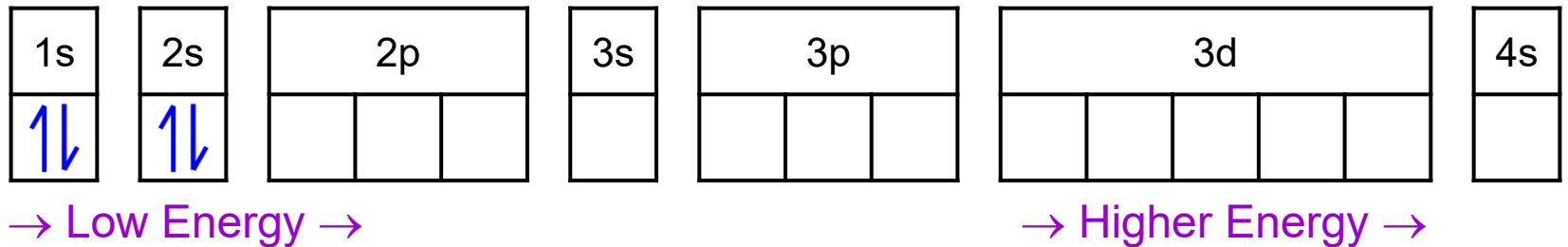
- Atomic Number: 3
- Name: Lithium
- Symbol: Li
- Electronic Configuration: $1s^2 2s^1$



- Electrons are represented by *arrows* (\uparrow and \downarrow) which fill atomic orbitals that are represented by *boxes*.
- Electrons occupy atomic orbitals from the lowest energy to the highest energy. This is known as the *Aufbau Principle*.

Advanced Theories of Atomic Structure

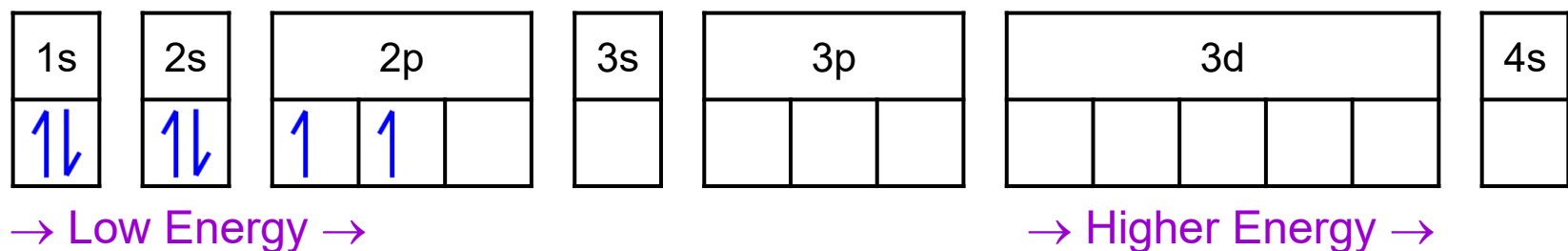
- Atomic Number: 4
- Name: Beryllium
- Symbol: Be
- Electronic Configuration: $1s^2 2s^2$



- Electrons are represented by *arrows* (↑ and ↓) which fill atomic orbitals that are represented by *boxes*.
- Electrons occupy atomic orbitals from the lowest energy to the highest energy. This is known as the *Aufbau Principle*.

Advanced Theories of Atomic Structure

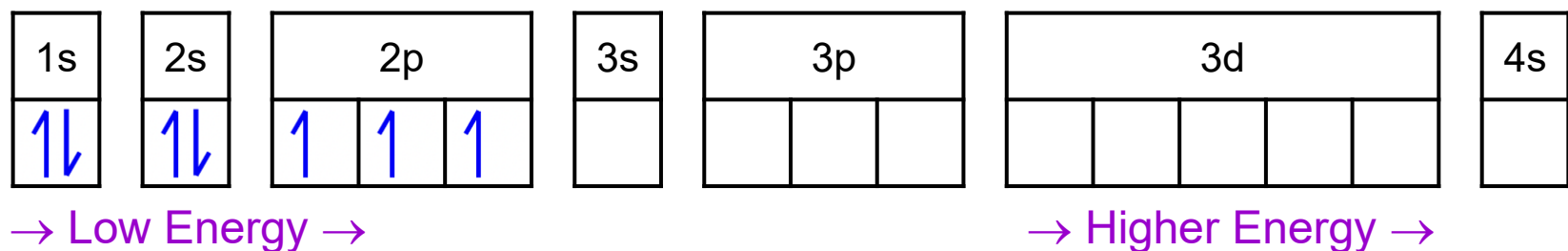
- Atomic Number: 6
- Name: Carbon
- Symbol: C
- Electronic Configuration: $1s^2 2s^2 2p^2$



- Electrons are represented by *arrows* (↑ and ↓) which fill atomic orbitals that are represented by *boxes*.
- Electrons occupy atomic orbitals from the lowest energy to the highest energy. This is known as the *Aufbau Principle*.
- A single electron will occupy a single atomic orbital before two electrons are forced to *spin pair-up* with each other in the same orbital (*Hund's Rule*).

Advanced Theories of Atomic Structure

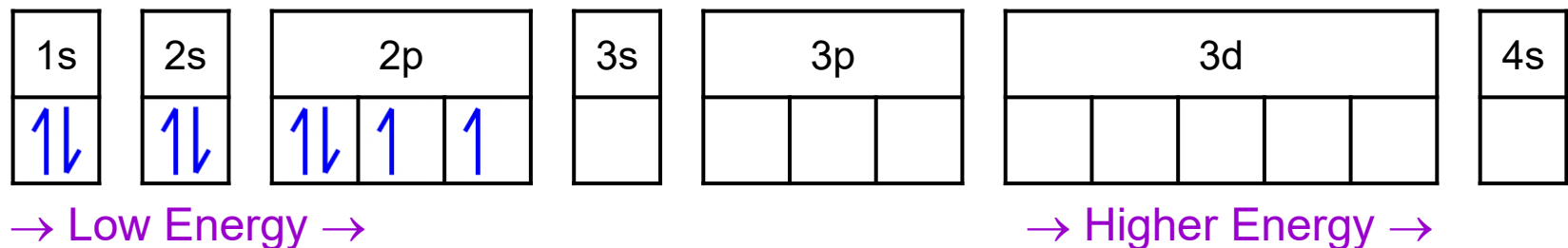
- Atomic Number: 7
- Name: Nitrogen
- Symbol: N
- Electronic Configuration: $1s^2 2s^2 2p^3$



- Electrons are represented by *arrows* (\uparrow and \downarrow) which fill atomic orbitals that are represented by *boxes*.
- Electrons occupy atomic orbitals from the lowest energy to the highest energy. This is known as the *Aufbau Principle*.
- A single electron will occupy a single atomic orbital before two electrons are forced to *spin pair-up* with each other in the same orbital (*Hund's Rule*).

Advanced Theories of Atomic Structure

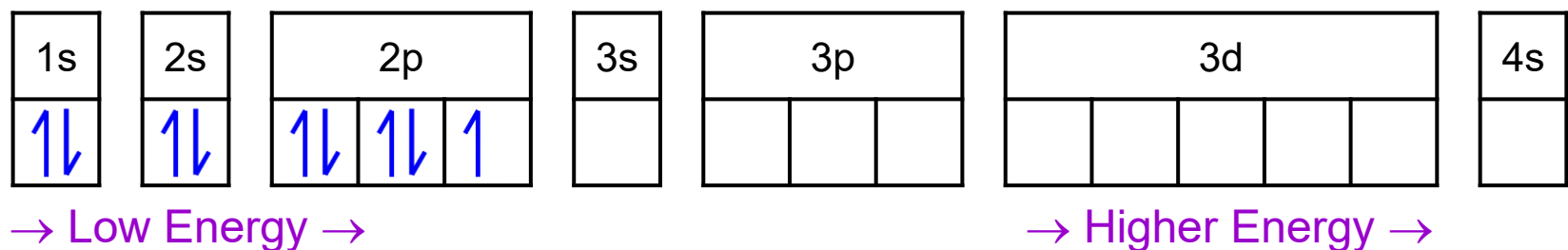
- Atomic Number: 8
- Name: Oxygen
- Symbol: O
- Electronic Configuration: $1s^2 2s^2 2p^4$



- Electrons are represented by *arrows* (\uparrow and \downarrow) which fill atomic orbitals that are represented by *boxes*.
- Electrons occupy atomic orbitals from the lowest energy to the highest energy. This is known as the *Aufbau Principle*.
- A single electron will occupy a single atomic orbital before two electrons are forced to *spin pair-up* with each other in the same orbital (*Hund's Rule*).

Advanced Theories of Atomic Structure

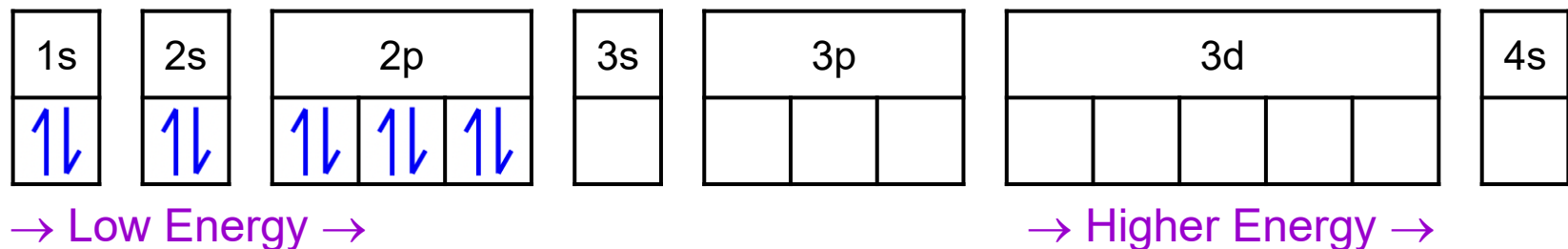
- Atomic Number: 9
- Name: Fluorine
- Symbol: F
- Electronic Configuration: $1s^2 2s^2 2p^5$



- Electrons are represented by *arrows* (\uparrow and \downarrow) which fill atomic orbitals that are represented by *boxes*.
- Electrons occupy atomic orbitals from the lowest energy to the highest energy. This is known as the *Aufbau Principle*.
- A single electron will occupy a single atomic orbital before two electrons are forced to *spin pair-up* with each other in the same orbital (*Hund's Rule*).

Advanced Theories of Atomic Structure

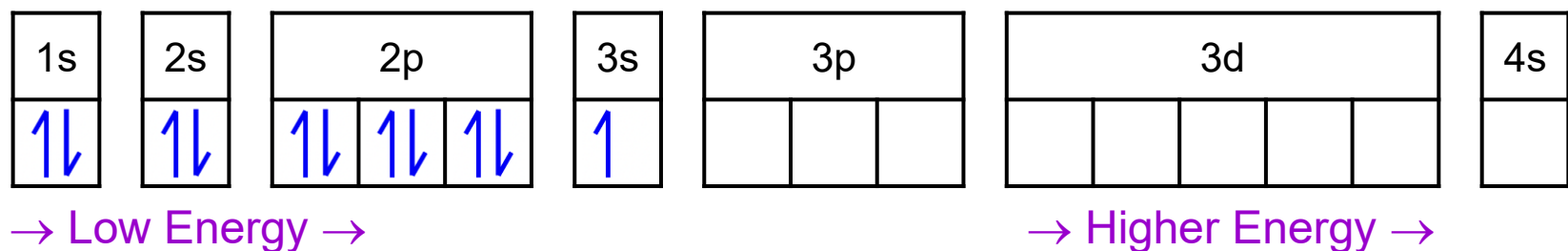
- Atomic Number: 10
- Name: Neon
- Symbol: Ne
- Electronic Configuration: $1s^2 2s^2 2p^6$



- Electrons are represented by *arrows* (\uparrow and \downarrow) which fill atomic orbitals that are represented by *boxes*.
- Electrons occupy atomic orbitals from the lowest energy to the highest energy. This is known as the *Aufbau Principle*.
- A single electron will occupy a single atomic orbital before two electrons are forced to *spin pair-up* with each other in the same orbital (*Hund's Rule*).

Advanced Theories of Atomic Structure

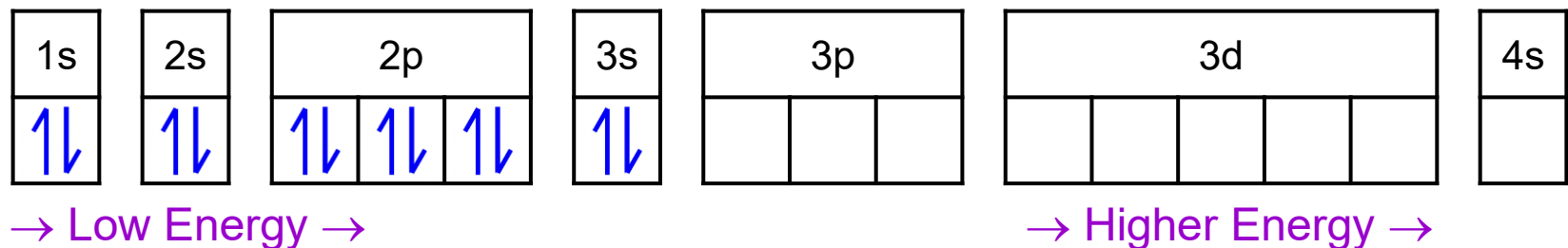
- Atomic Number: 11
- Name: Sodium
- Symbol: Na
- Electronic Configuration: $1s^2 2s^2 2p^6 3s^1$



- Electrons are represented by *arrows* (\uparrow and \downarrow) which fill atomic orbitals that are represented by *boxes*.
- Electrons occupy atomic orbitals from the lowest energy to the highest energy. This is known as the *Aufbau Principle*.
- A single electron will occupy a single atomic orbital before two electrons are forced to *spin pair-up* with each other in the same orbital (*Hund's Rule*).

Advanced Theories of Atomic Structure

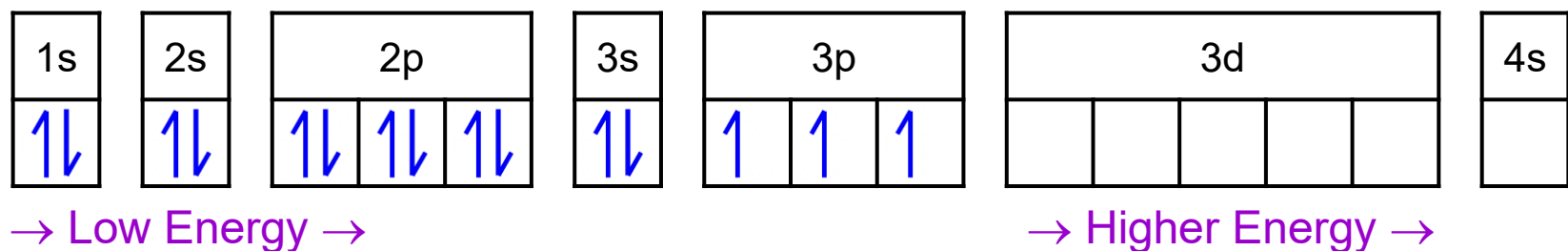
- Atomic Number: 12
- Name: Magnesium
- Symbol: Mg
- Electronic Configuration: $1s^2 2s^2 2p^6 3s^2$



- Electrons are represented by *arrows* (\uparrow and \downarrow) which fill atomic orbitals that are represented by *boxes*.
- Electrons occupy atomic orbitals from the lowest energy to the highest energy. This is known as the *Aufbau Principle*.
- A single electron will occupy a single atomic orbital before two electrons are forced to *spin pair-up* with each other in the same orbital (*Hund's Rule*).

Advanced Theories of Atomic Structure

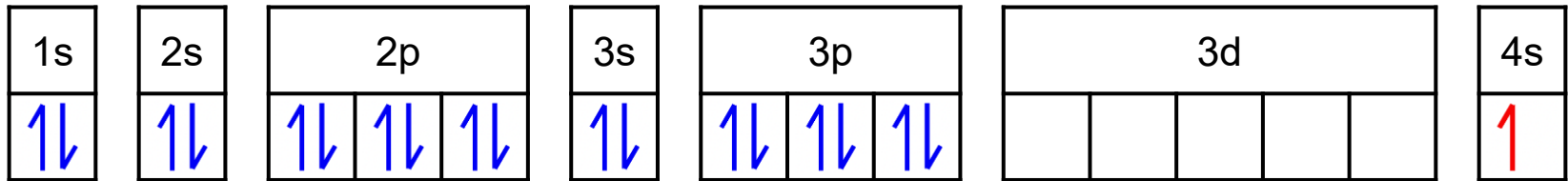
- Atomic Number: 15
- Name: Phosphorus
- Symbol: P
- Electronic Configuration: $1s^2 2s^2 2p^6 3s^2 3p^3$



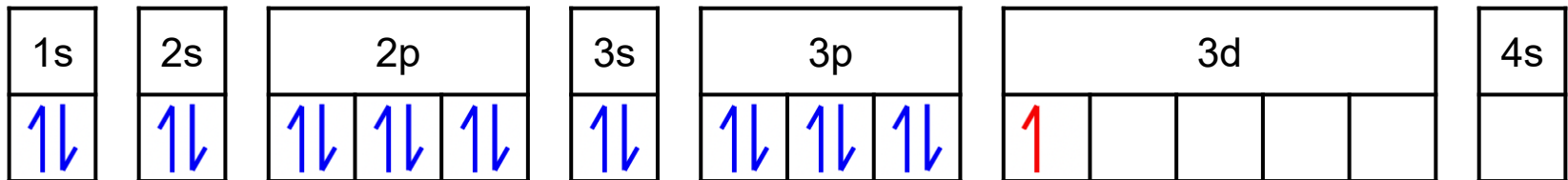
- Electrons are represented by *arrows* (\uparrow and \downarrow) which fill atomic orbitals that are represented by *boxes*.
- Electrons occupy atomic orbitals from the lowest energy to the highest energy. This is known as the *Aufbau Principle*.
- A single electron will occupy a single atomic orbital before two electrons are forced to *spin pair-up* with each other in the same orbital (*Hund's Rule*).

Advanced Theories of Atomic Structure

- Why is the electronic configuration of *potassium*...



...instead of...

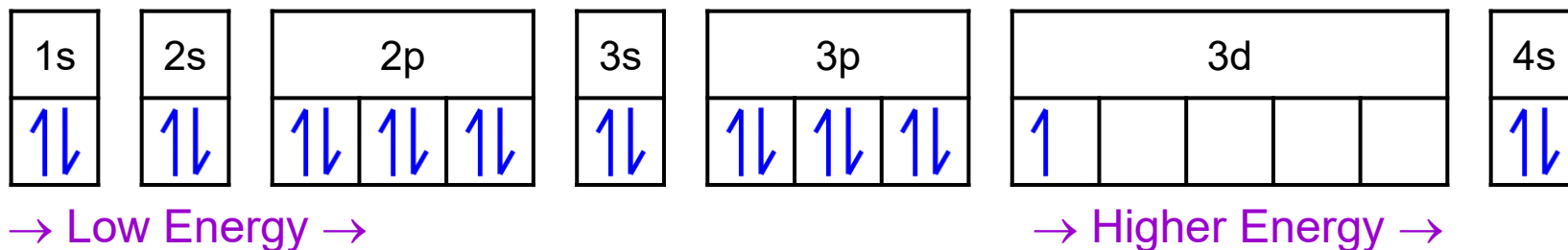


- Although the 4s sub-shell is further from the nucleus than the 3d sub-shell, the 4s sub-shell is *lower in energy* than the 3d sub-shell.
- According to the Aufbau Principle (electrons fill-up atomic orbitals from lower energy to higher energy) the *lower energy* 4s sub-shell will fill with electrons before the *higher energy* 3d sub-shell.

Advanced Theories of Atomic Structure

- Atomic Number: 21
- Name: Scandium
- Symbol: Sc
- Electronic Configuration: $1s^2 2s^2 2p^6 3s^2 3p^6 3d^1 4s^2$

• Lower energy than 3d

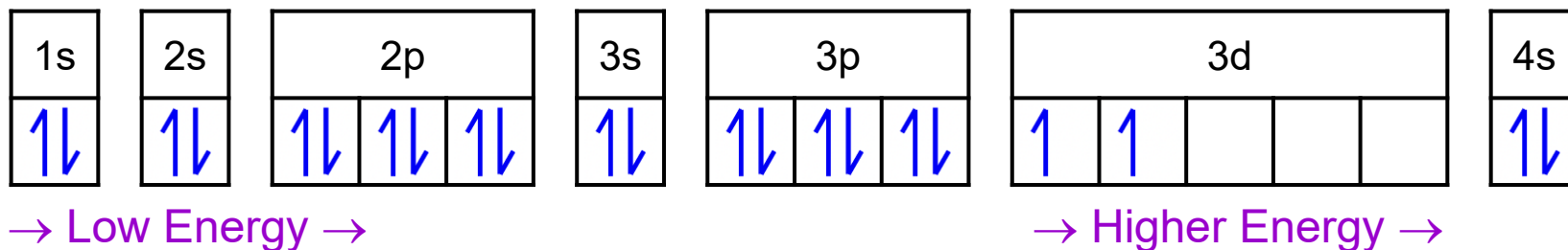


- Electrons are represented by *arrows* (\uparrow and \downarrow) which fill atomic orbitals that are represented by *boxes*.
- Electrons occupy atomic orbitals from the lowest energy to the highest energy. This is known as the *Aufbau Principle*.
- A single electron will occupy a single atomic orbital before two electrons are forced to *spin pair-up* with each other in the same orbital (*Hund's Rule*).

Advanced Theories of Atomic Structure

- Atomic Number: 22
- Name: Titanium
- Symbol: Ti
- Electronic Configuration: $1s^2 2s^2 2p^6 3s^2 3p^6 3d^2 4s^2$

• Lower energy than 3d

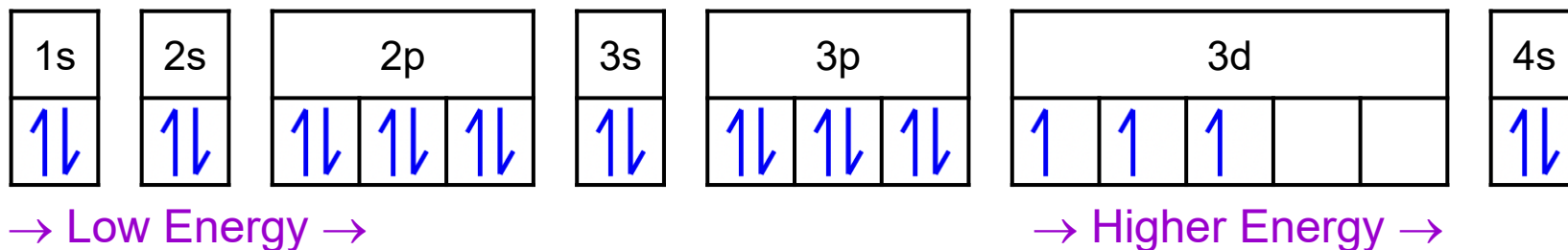


- Electrons are represented by *arrows* (↑ and ↓) which fill atomic orbitals that are represented by *boxes*.
- Electrons occupy atomic orbitals from the lowest energy to the highest energy. This is known as the *Aufbau Principle*.
- A single electron will occupy a single atomic orbital before two electrons are forced to *spin pair-up* with each other in the same orbital (*Hund's Rule*).

Advanced Theories of Atomic Structure

- Atomic Number: 23
- Name: Vanadium
- Symbol: V
- Electronic Configuration: $1s^2 2s^2 2p^6 3s^2 3p^6 3d^3 4s^2$

• Lower energy than 3d

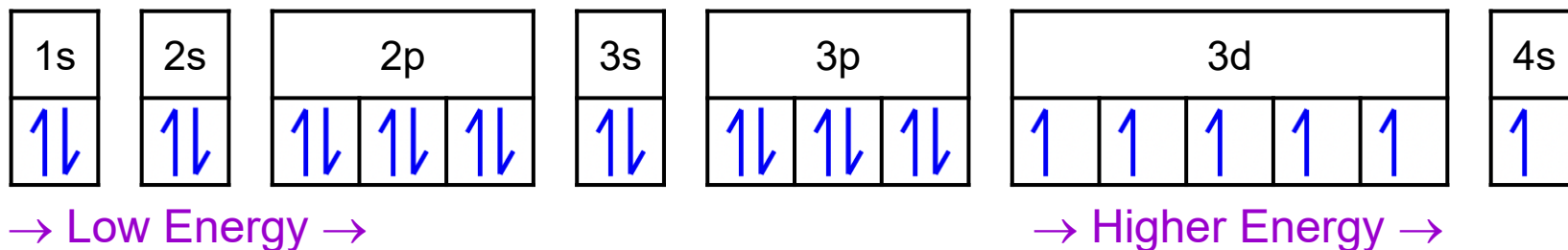


- Electrons are represented by *arrows* (↑ and ↓) which fill atomic orbitals that are represented by *boxes*.
- Electrons occupy atomic orbitals from the lowest energy to the highest energy. This is known as the *Aufbau Principle*.
- A single electron will occupy a single atomic orbital before two electrons are forced to *spin pair-up* with each other in the same orbital (*Hund's Rule*).

Advanced Theories of Atomic Structure

- Atomic Number: 24
- Name: Chromium
- Symbol: Cr
- Electronic Configuration: $1s^2 2s^2 2p^6 3s^2 3p^6 3d^5 4s^1$

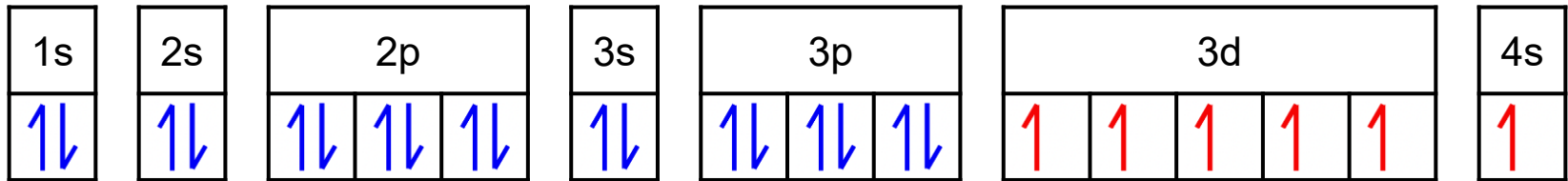
• Lower energy than 3d



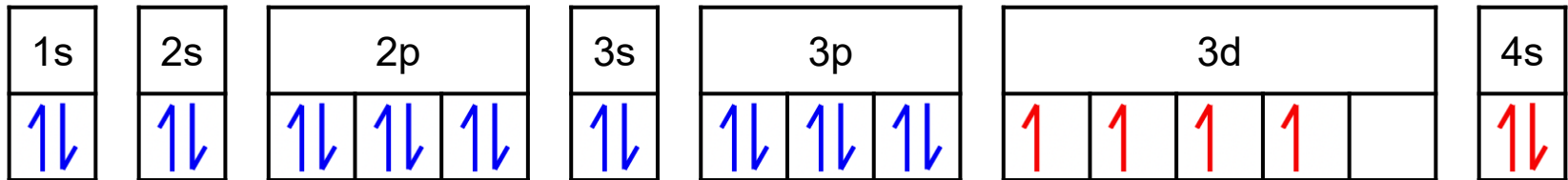
- Electrons are represented by *arrows* (↑ and ↓) which fill atomic orbitals that are represented by *boxes*.
- Electrons occupy atomic orbitals from the lowest energy to the highest energy. This is known as the *Aufbau Principle*.
- A single electron will occupy a single atomic orbital before two electrons are forced to *spin pair-up* with each other in the same orbital (*Hund's Rule*).

Advanced Theories of Atomic Structure

- Why is the electronic configuration of *chromium*...



...instead of...

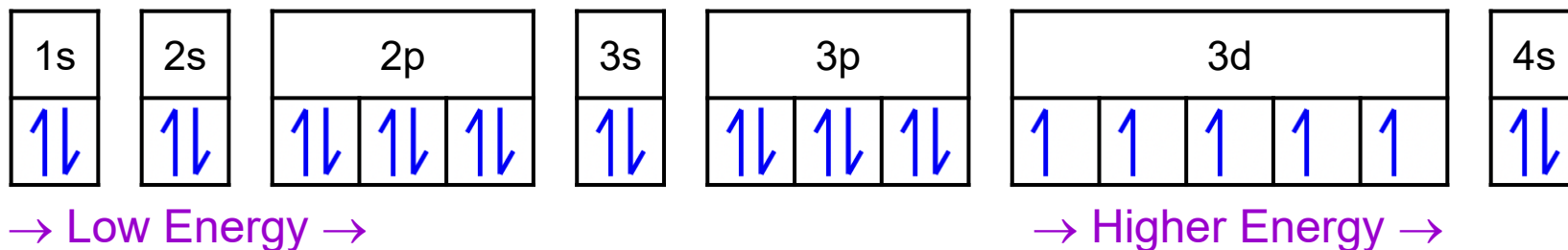


- Completely filled sub-shells are *more stable* than partially filled sub-shells.
- A sub-shell that is exactly half-filled is *more stable* than a sub-shell that is not exactly half-filled.
- An electron in the 4s orbital is transferred to an empty 3d orbital so as to obtain two stable half-filled sub-shells ($3d^5$ and $4s^1$) instead of one incomplete sub-shell ($3d^4$) and one complete sub-shell ($4s^2$).

Advanced Theories of Atomic Structure

- Atomic Number: 25
- Name: Manganese
- Symbol: Mn
- Electronic Configuration: $1s^2 2s^2 2p^6 3s^2 3p^6 3d^5 4s^2$

• Lower energy than 3d

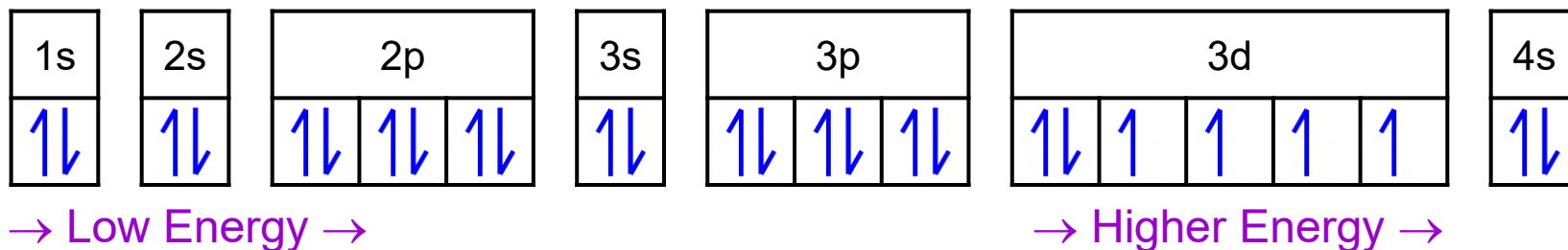


- Electrons are represented by *arrows* (\uparrow and \downarrow) which fill atomic orbitals that are represented by *boxes*.
- Electrons occupy atomic orbitals from the lowest energy to the highest energy. This is known as the *Aufbau Principle*.
- A single electron will occupy a single atomic orbital before two electrons are forced to *spin pair-up* with each other in the same orbital (*Hund's Rule*).

Advanced Theories of Atomic Structure

- Atomic Number: 26
- Name: Iron
- Symbol: Fe
- Electronic Configuration: $1s^2 2s^2 2p^6 3s^2 3p^6 3d^6 4s^2$

• Lower energy than 3d

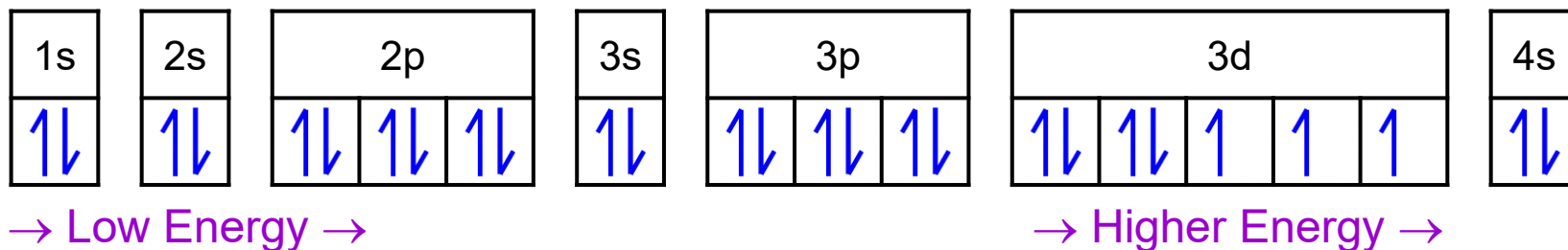


- Electrons are represented by *arrows* (\uparrow and \downarrow) which fill atomic orbitals that are represented by *boxes*.
- Electrons occupy atomic orbitals from the lowest energy to the highest energy. This is known as the *Aufbau Principle*.
- A single electron will occupy a single atomic orbital before two electrons are forced to *spin pair-up* with each other in the same orbital (*Hund's Rule*).

Advanced Theories of Atomic Structure

- Atomic Number: 27
- Name: Cobalt
- Symbol: Co
- Electronic Configuration: $1s^2 2s^2 2p^6 3s^2 3p^6 3d^7 4s^2$

• Lower energy than 3d

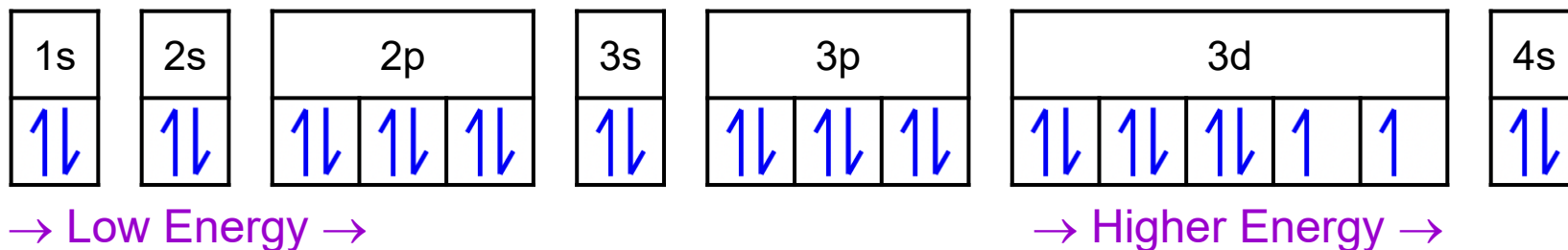


- Electrons are represented by *arrows* (↑ and ↓) which fill atomic orbitals that are represented by *boxes*.
- Electrons occupy atomic orbitals from the lowest energy to the highest energy. This is known as the *Aufbau Principle*.
- A single electron will occupy a single atomic orbital before two electrons are forced to *spin pair-up* with each other in the same orbital (*Hund's Rule*).

Advanced Theories of Atomic Structure

- Atomic Number: 28
- Name: Nickel
- Symbol: Ni
- Electronic Configuration: $1s^2 2s^2 2p^6 3s^2 3p^6 3d^8 4s^2$

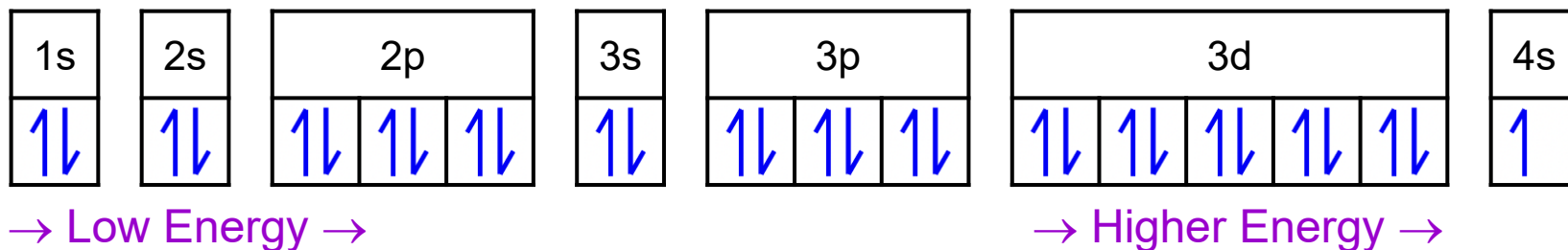
• Lower energy than 3d



- Electrons are represented by *arrows* (↑ and ↓) which fill atomic orbitals that are represented by *boxes*.
- Electrons occupy atomic orbitals from the lowest energy to the highest energy. This is known as the *Aufbau Principle*.
- A single electron will occupy a single atomic orbital before two electrons are forced to *spin pair-up* with each other in the same orbital (*Hund's Rule*).

Advanced Theories of Atomic Structure

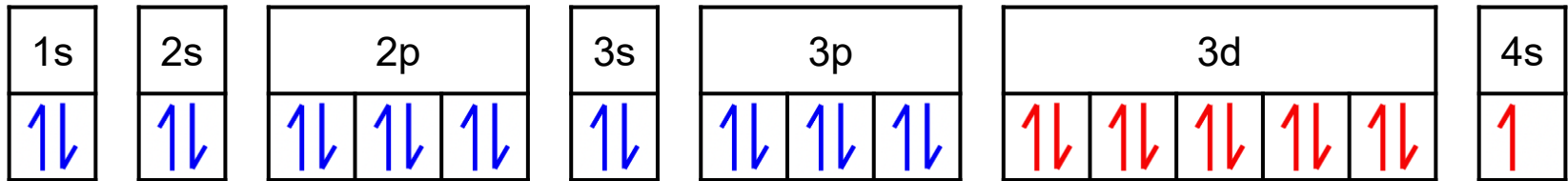
- Atomic Number: 29
- Name: Copper
- Symbol: Cu
- Electronic Configuration: $1s^2 2s^2 2p^6 3s^2 3p^6 3d^{10} 4s^1$ • Lower energy than 3d



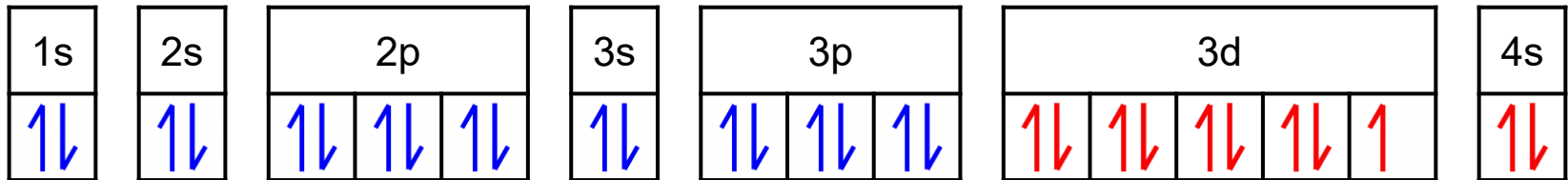
- Electrons are represented by *arrows* (\uparrow and \downarrow) which fill atomic orbitals that are represented by *boxes*.
- Electrons occupy atomic orbitals from the lowest energy to the highest energy. This is known as the *Aufbau Principle*.
- A single electron will occupy a single atomic orbital before two electrons are forced to *spin pair-up* with each other in the same orbital (*Hund's Rule*).

Advanced Theories of Atomic Structure

- Why is the electronic configuration of *copper*...



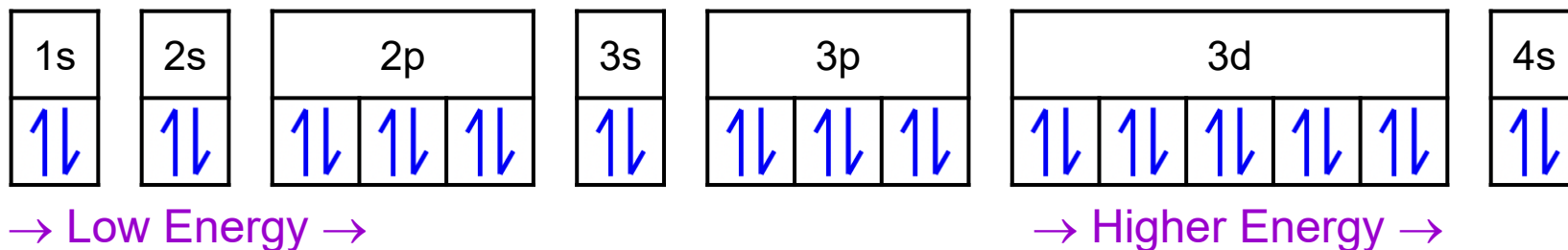
...instead of...



- Completely filled sub-shells are *more stable* than partially filled sub-shells.
- A sub-shell that is exactly half-filled is *more stable* than a sub-shell that is not exactly half-filled.
- An electron in the 4s orbital is transferred to a 3d orbital so as to obtain one stable complete sub-shell ($3d^{10}$) and one stable half-filled sub-shell ($4s^1$) instead of one incomplete sub-shell ($3d^9$) and one complete sub-shell ($4s^2$).

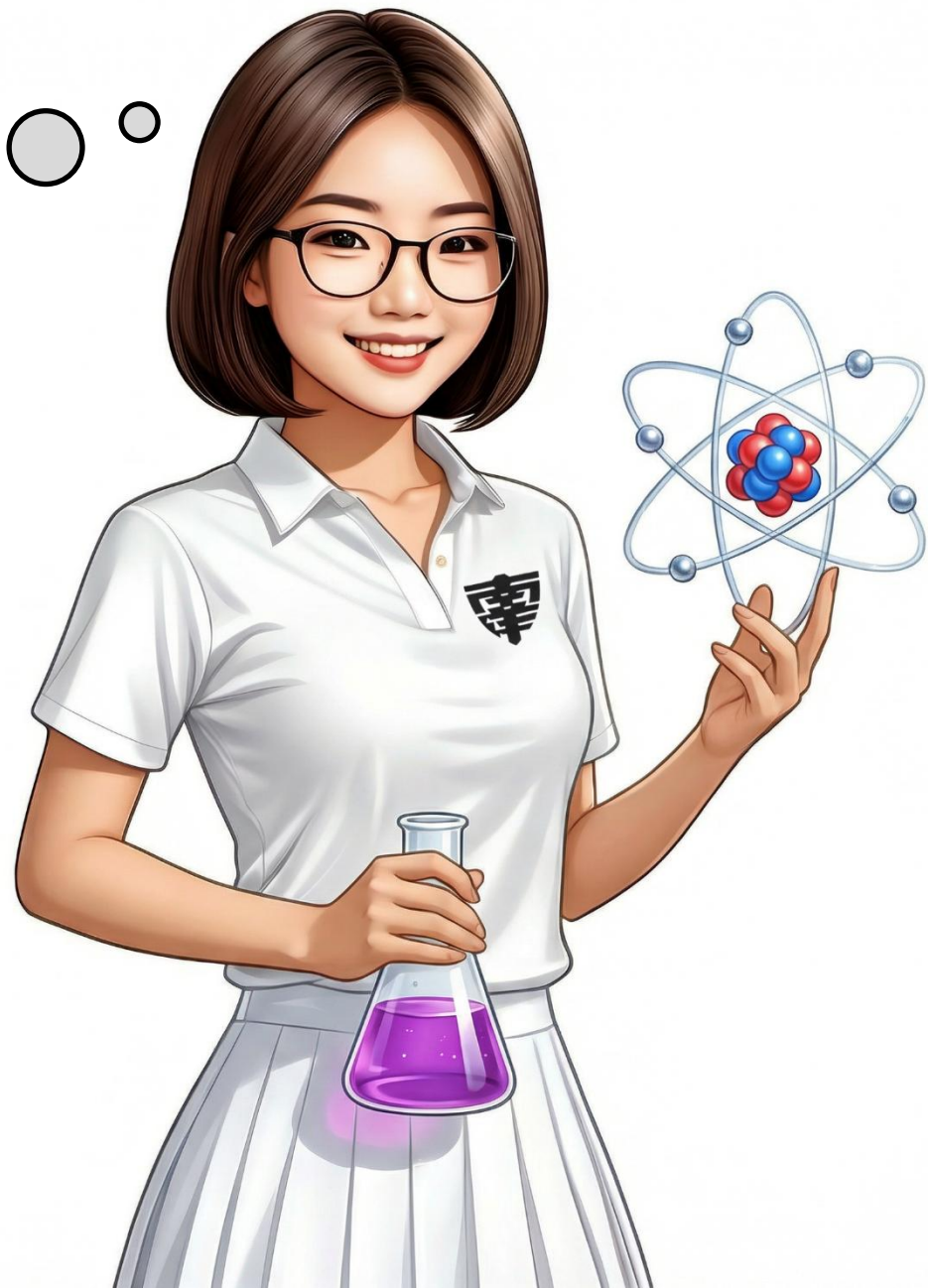
Advanced Theories of Atomic Structure

- Atomic Number: 30
- Name: Zinc
- Symbol: Zn
- Electronic Configuration: $1s^2 2s^2 2p^6 3s^2 3p^6 3d^{10} 4s^2$. Lower energy than 3d



- Electrons are represented by *arrows* (\uparrow and \downarrow) which fill atomic orbitals that are represented by *boxes*.
- Electrons occupy atomic orbitals from the lowest energy to the highest energy. This is known as the *Aufbau Principle*.
- A single electron will occupy a single atomic orbital before two electrons are forced to *spin pair-up* with each other in the same orbital (*Hund's Rule*).

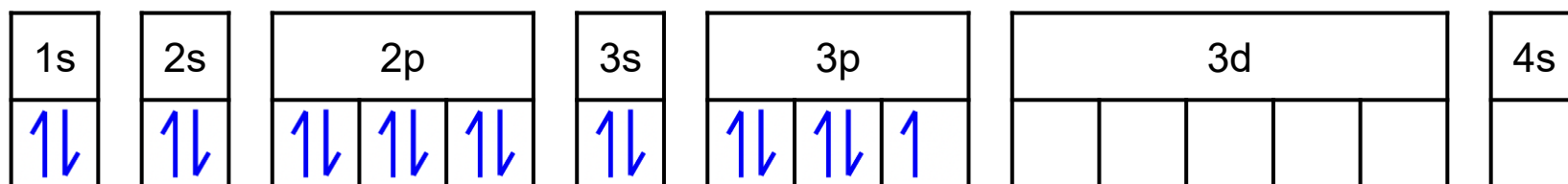
How are the
electronic
configurations
of ions drawn
using orbital
notation?



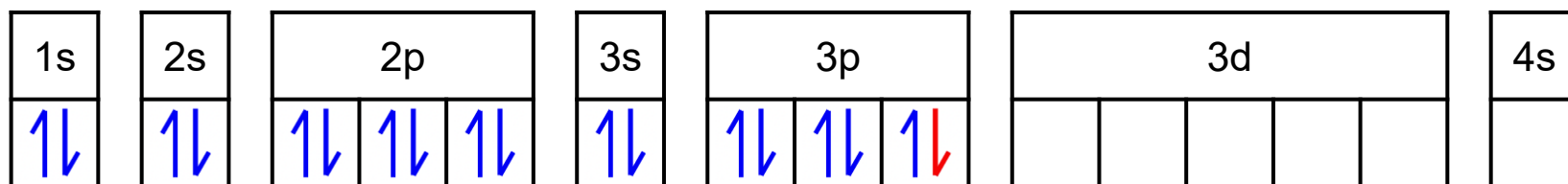
Advanced Theories of Atomic Structure

Electronic Configurations of Ions

Electronic Configuration of a Chlorine Atom



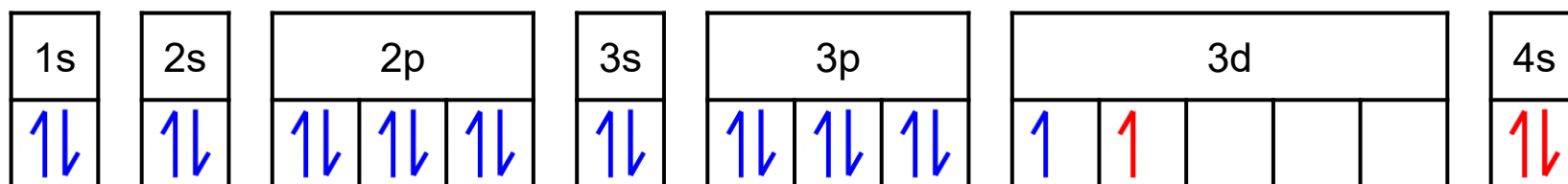
Electronic Configuration of a Chloride Ion – Cl^-



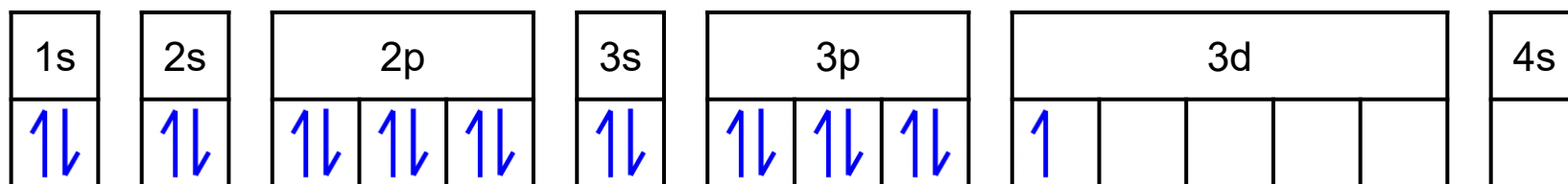
Advanced Theories of Atomic Structure

Electronic Configurations of Ions

Electronic Configuration of a Titanium Atom



Electronic Configuration of a Titanium(III) Ion – Ti^{3+}

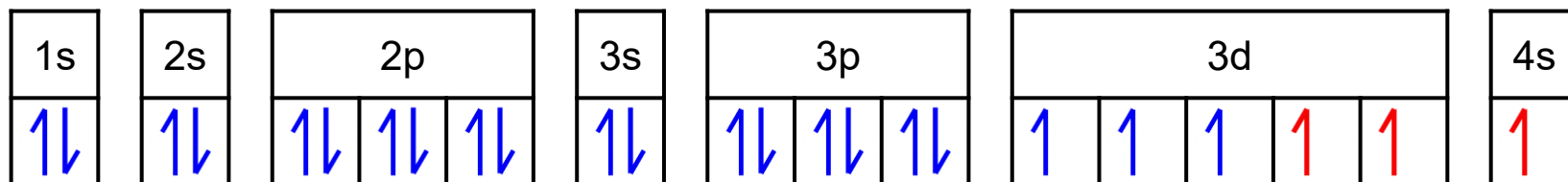


Note: When atoms of the transition metals react to form ions, electrons from the 4s orbital are the first to be removed.

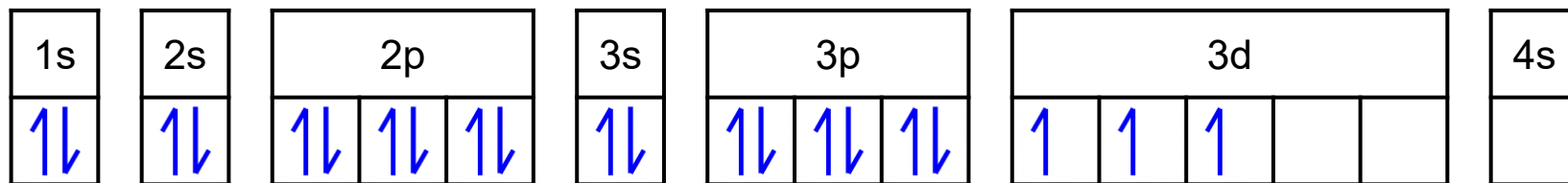
Advanced Theories of Atomic Structure

Electronic Configurations of Ions

Electronic Configuration of a Chromium Atom



Electronic Configuration of a Chromium(III) Ion – Cr^{3+}

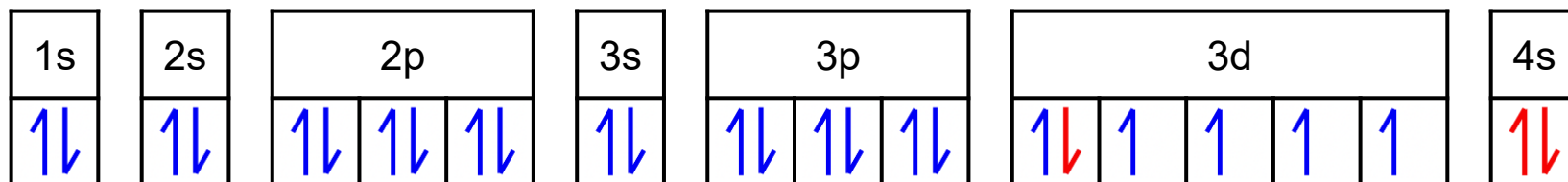


Note: When atoms of the transition metals react to form ions, electrons from the 4s orbital are the first to be removed.

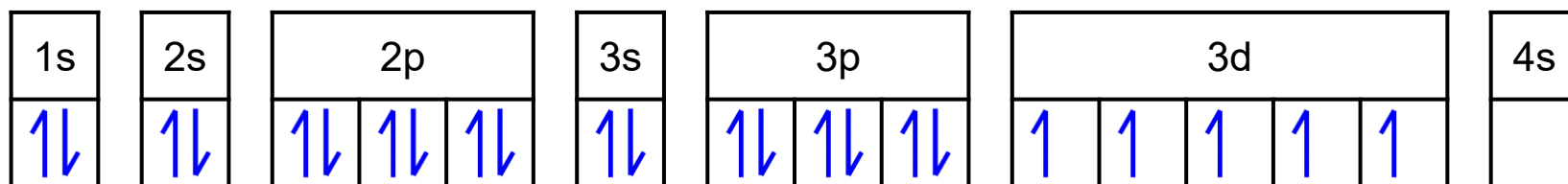
Advanced Theories of Atomic Structure

Electronic Configurations of Ions

Electronic Configuration of an Iron Atom



Electronic Configuration of a Iron(III) Ion – Fe³⁺

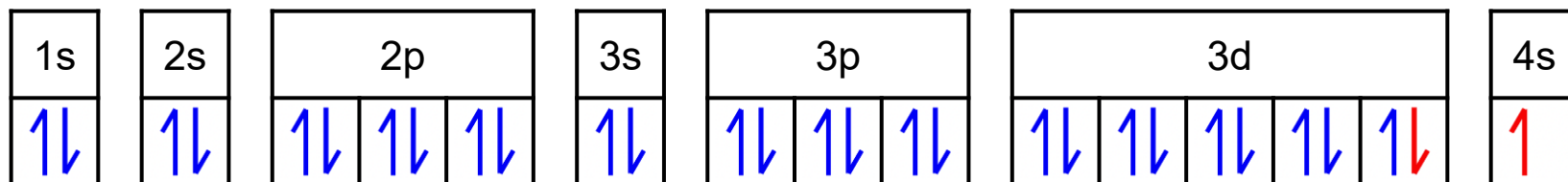


Note: When atoms of the transition metals react to form ions, electrons from the 4s orbital are the first to be removed.

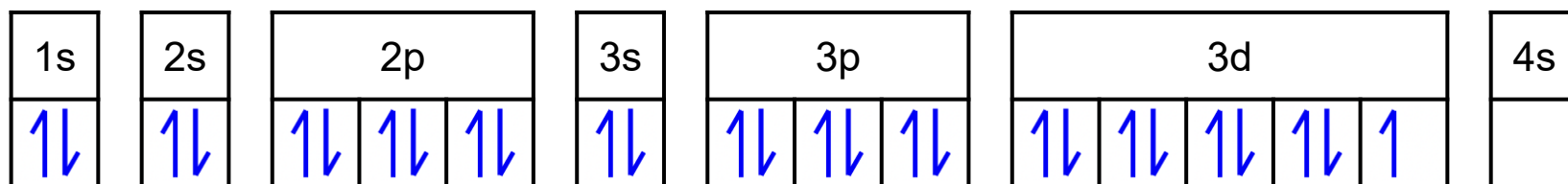
Advanced Theories of Atomic Structure

Electronic Configurations of Ions

Electronic Configuration of a Copper Atom



Electronic Configuration of a Copper(II) Ion – Cu^{2+}

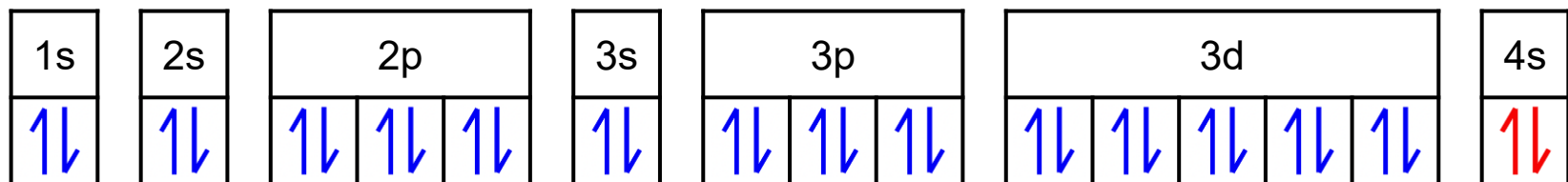


Note: When atoms of the transition metals react to form ions, electrons from the 4s orbital are the first to be removed.

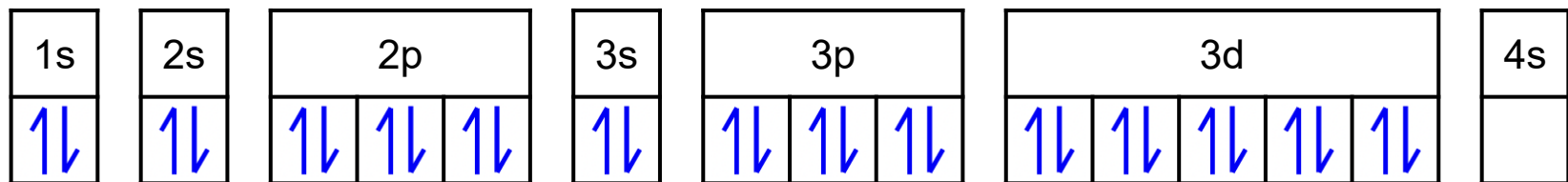
Advanced Theories of Atomic Structure

Electronic Configurations of Ions

Electronic Configuration of a Zinc Atom

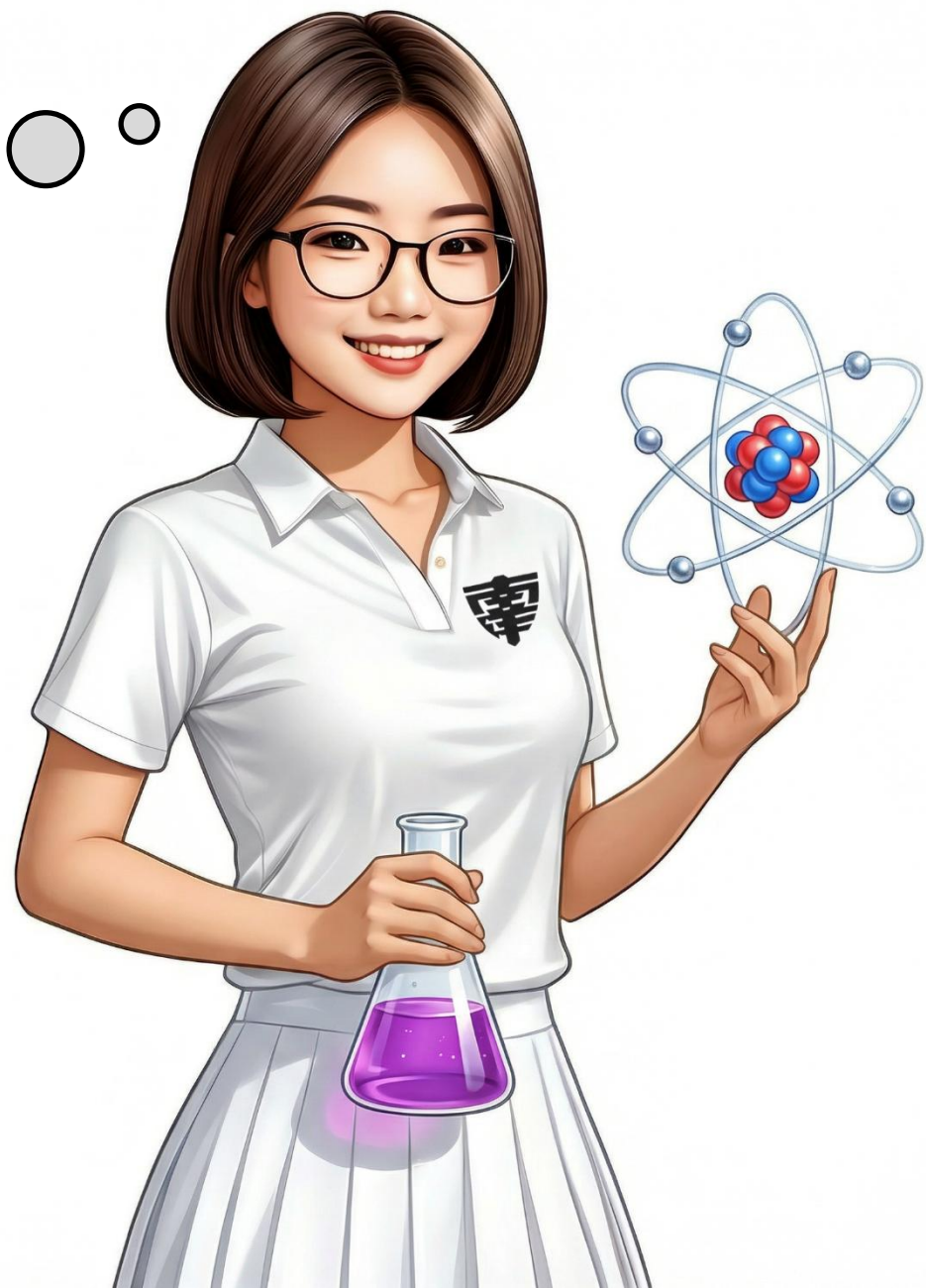


Electronic Configuration of a Zinc Ion – Zn^{2+}



Note: When atoms of the transition metals react to form ions, electrons from the 4s orbital are the first to be removed.

How is bonding
between atoms
represented
using orbital
notation?

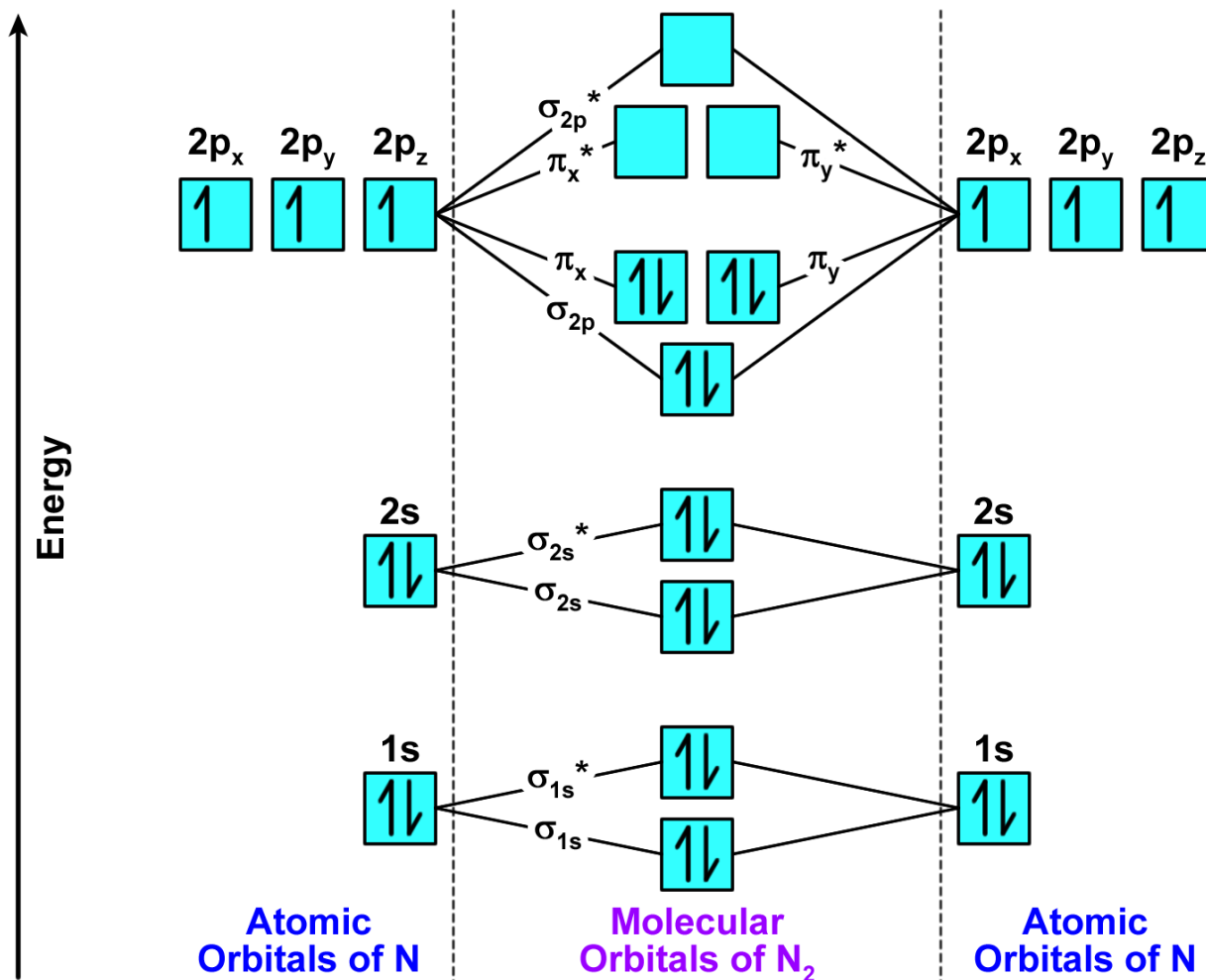


Advanced Theories of Atomic Structure

- The *atomic orbitals* of two or more atoms can combine together to form *molecular orbitals*.
 - The following diagrams show how the atomic orbitals of two atoms combine to form covalent bonds known as a σ -bonds (sigma-bonds) and π -bonds (pi-bonds).
- **Note:** To pair-up in a molecular orbital, electrons must have *opposite spin*.

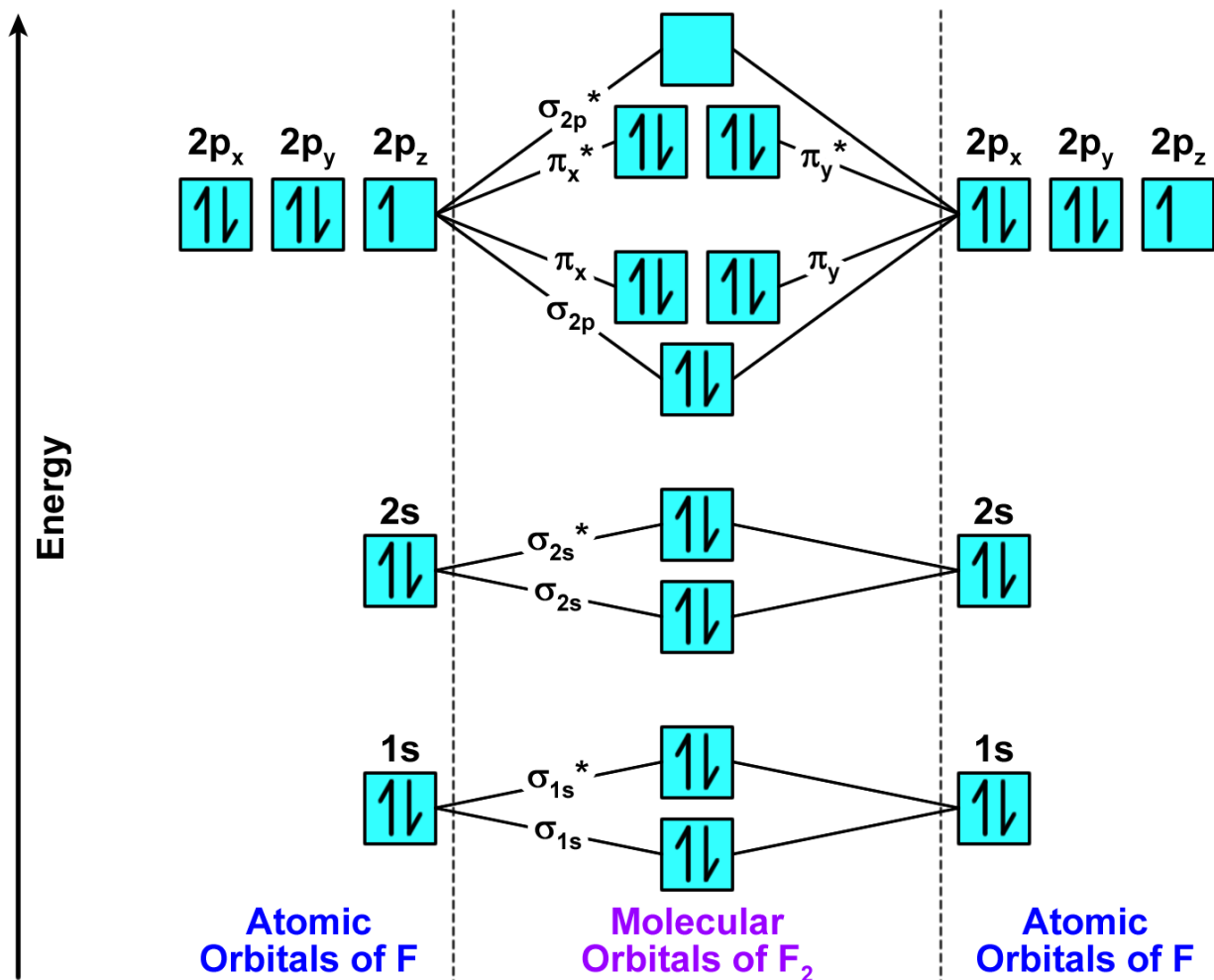
Advanced Theories of Atomic Structure

The Molecular Orbitals in Diatomic Nitrogen – N_2



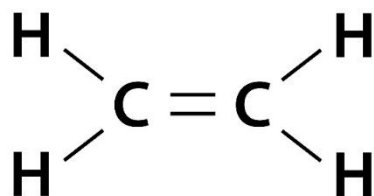
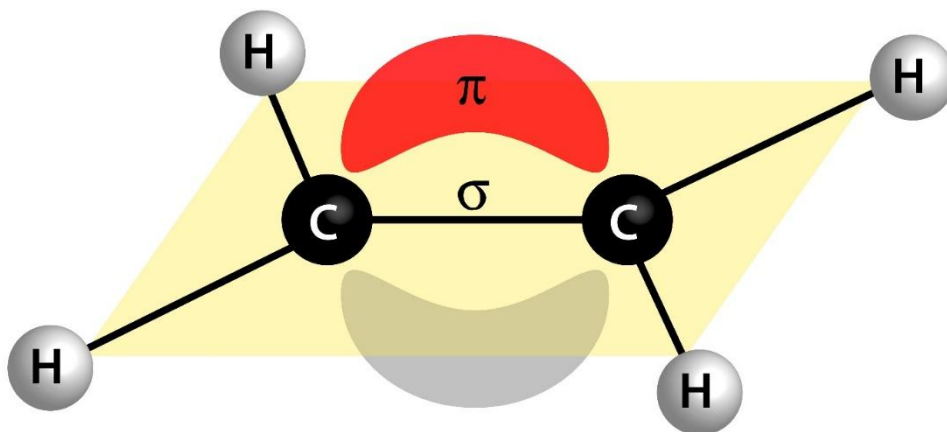
Advanced Theories of Atomic Structure

The Molecular Orbitals in Diatomic Fluorine – F_2



Advanced Theories of Atomic Structure

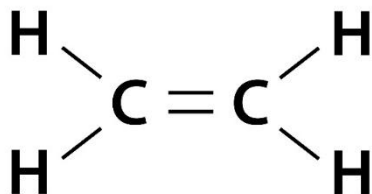
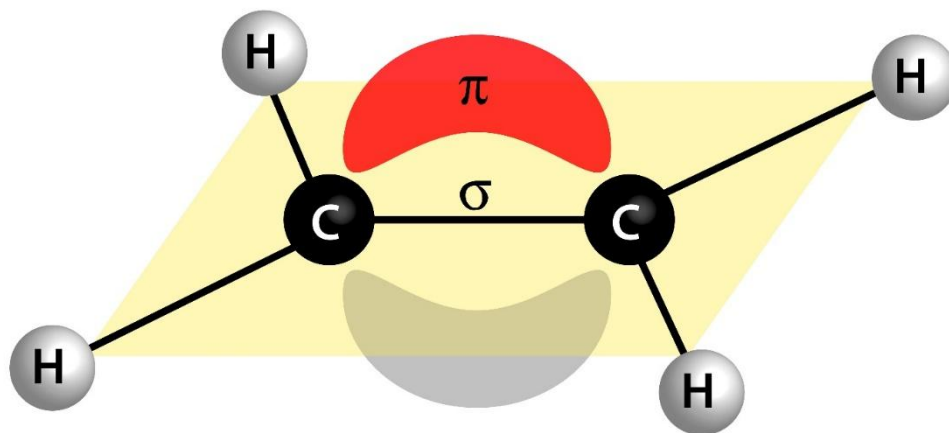
Pi-Bonds and Sigma-Bonds



- Examples of σ -bonds and π -bonds in a molecule of ethene, C₂H₄.

Advanced Theories of Atomic Structure

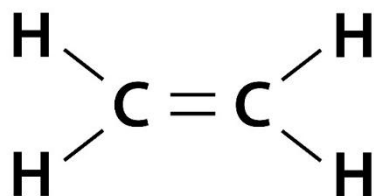
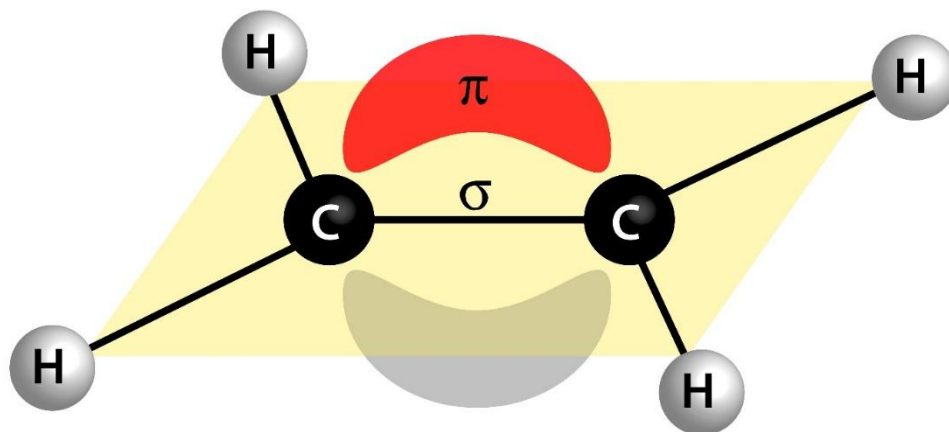
Pi-Bonds and Sigma-Bonds



- A σ -bond is formed when two atomic orbitals overlap, and the region of overlap lies on an imaginary straight line that connects the nuclei of the two bonding atoms.

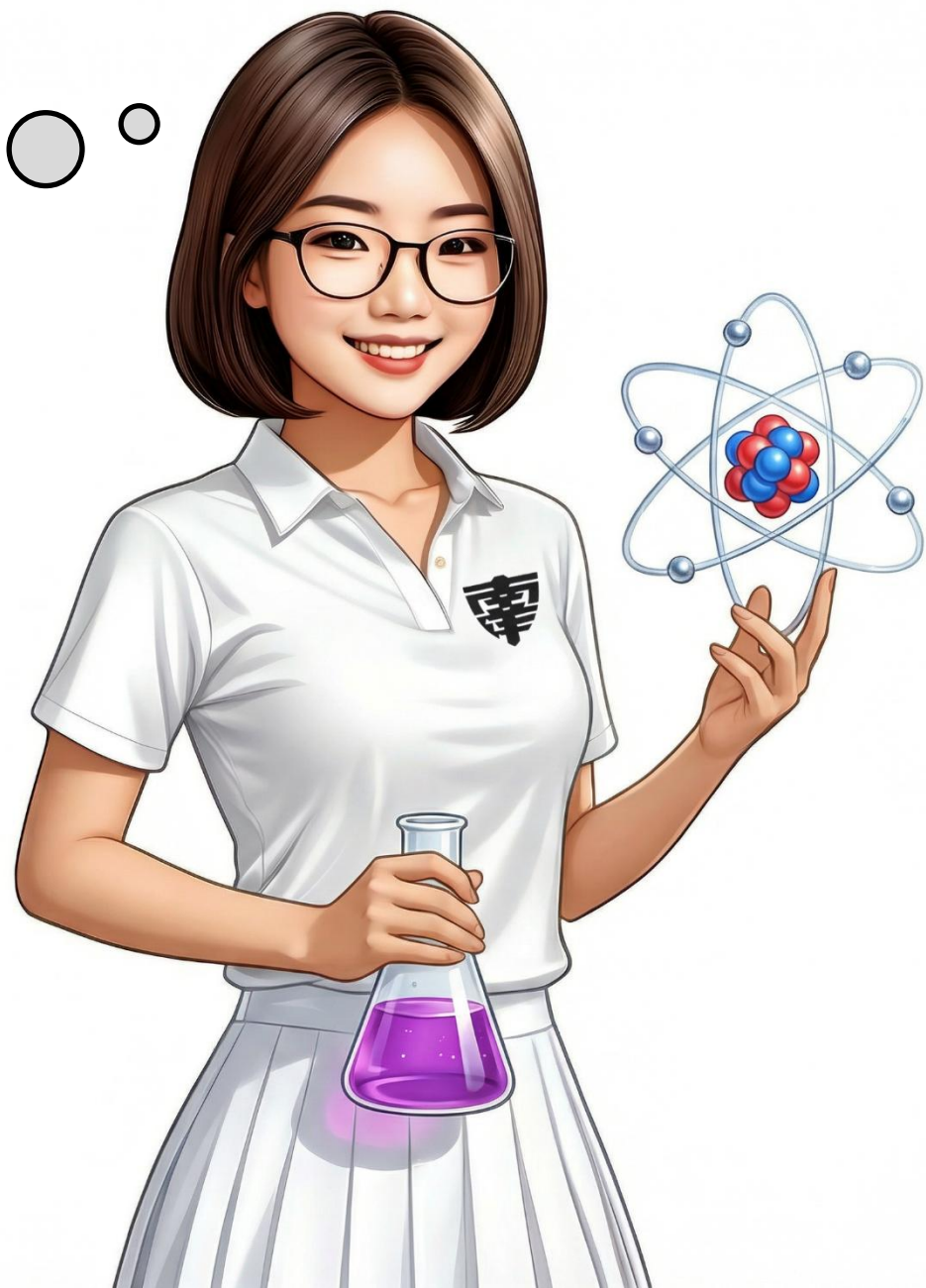
Advanced Theories of Atomic Structure

Pi-Bonds and Sigma-Bonds



- A π -bond is formed when two atomic orbitals (usually p -orbitals) overlap, and the region of overlap lies above and below an imaginary straight line that connects the nuclei of the two bonding atoms.

What are some periodic trends that can be explained using orbital theory?



Periodic Trends

- The force of attraction between oppositely charged particles is given by Coulomb's Law:

$$F = \frac{1}{4 \times \pi \times \epsilon_0} \times \frac{q_1 \times q_2}{r^2}$$

F = force of attraction between oppositely charged particles, N

ϵ_0 = permittivity of free space, $C^2 m^{-2} N^{-1}$

q_1 = charge on particle one, C

q_2 = charge on particle two, C

r = distance between particle one and particle two, m

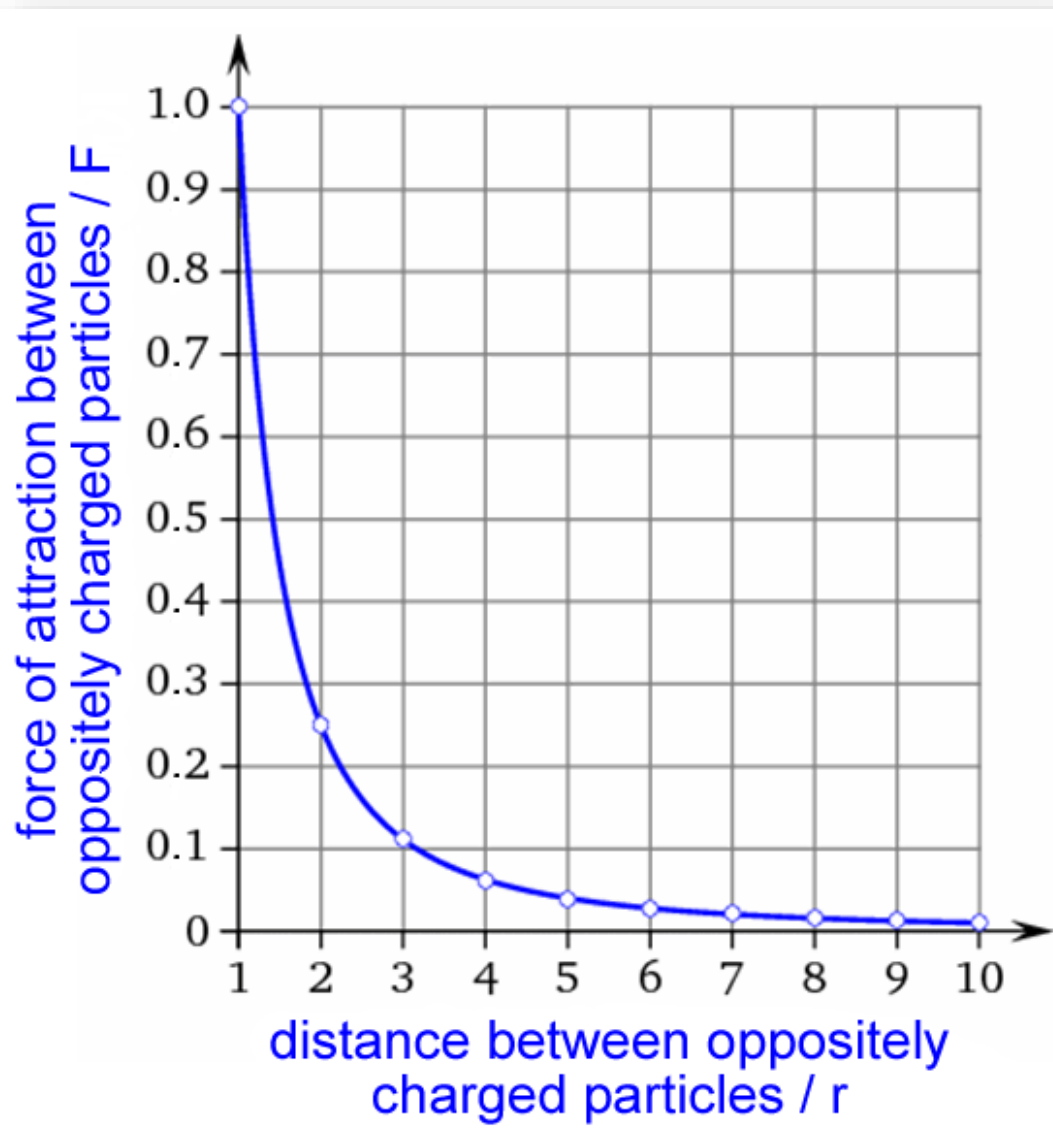
Periodic Trends

- The force of attraction between oppositely charged particles is given by Coulomb's Law:

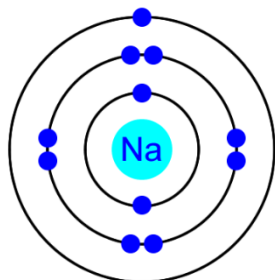
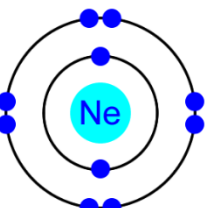
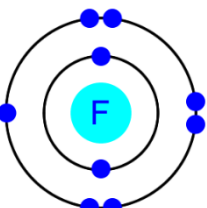
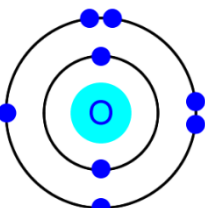
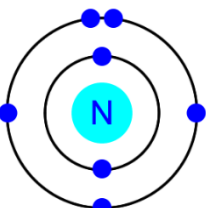
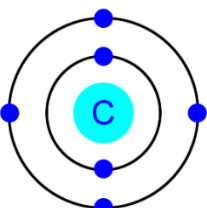
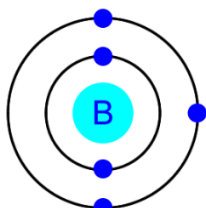
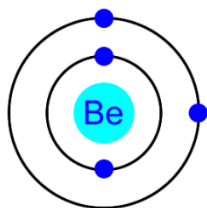
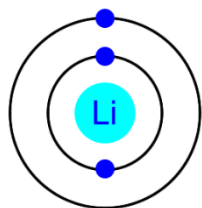
$$F = \frac{1}{4 \times \pi \times \epsilon_0} \times \frac{q_1 \times q_2}{r^2}$$

- The *force of attraction* (F) between a proton and an electron in an atom is related to their *charge* (q_1 and q_2), and it *decreases rapidly* as the *distance* between the particles (r) *increases* (inverse square law).

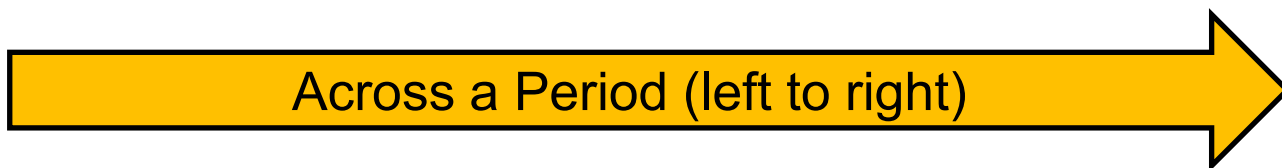
Periodic Trends



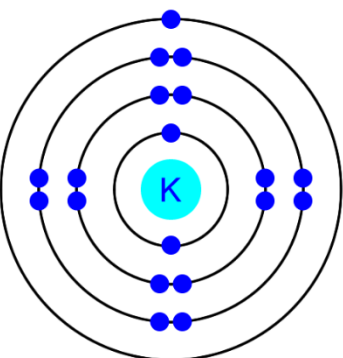
Periodic Trends



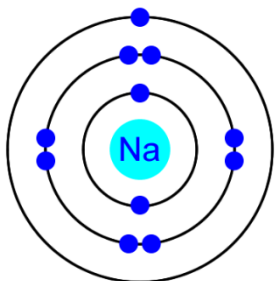
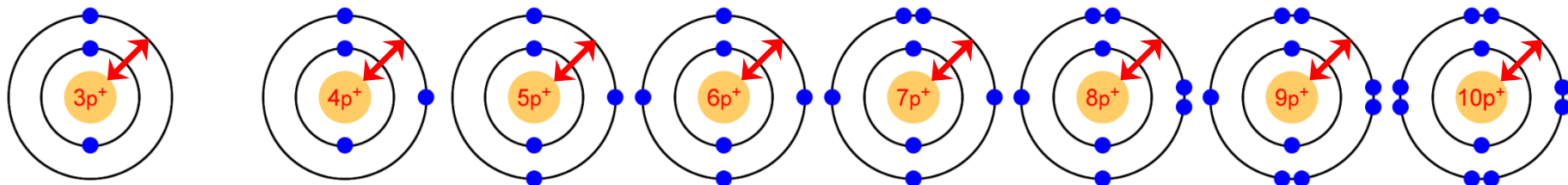
Down a Group



- Moving across a Period, there is an increase in the number of protons in the nucleus (increase in nuclear charge) but the number of electron shells remains constant.
- Due to the increasing nuclear charge, the *electrostatic force of attraction* between the nucleus and electrons in the valence electron shell *increases across a Period* from left to right.



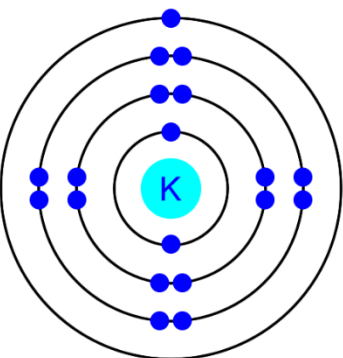
Periodic Trends



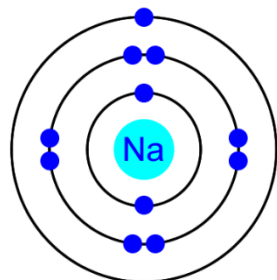
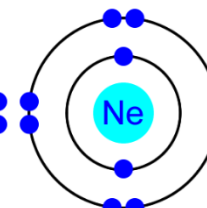
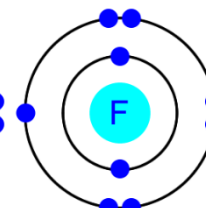
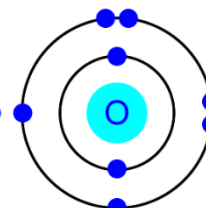
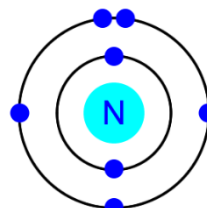
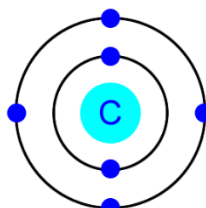
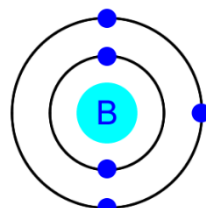
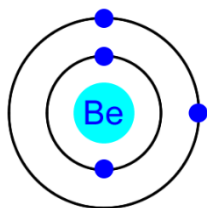
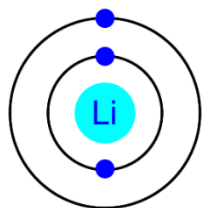
Down a Group

Across a Period (left to right)

- Moving across a Period, there is an increase in the number of protons in the nucleus (increase in nuclear charge) but the number of electron shells remains constant.
- Due to the increasing nuclear charge, the *electrostatic force of attraction* between the nucleus and electrons in the valence electron shell *increases across a Period* from left to right.



Periodic Trends

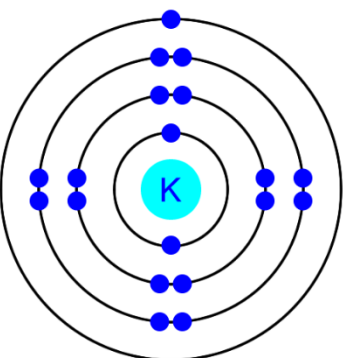


Down a Group

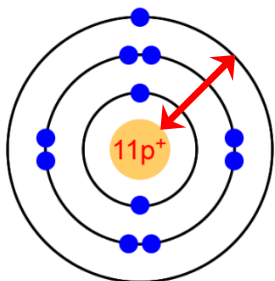
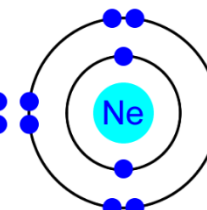
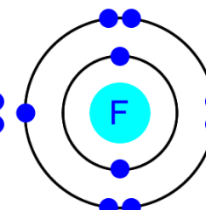
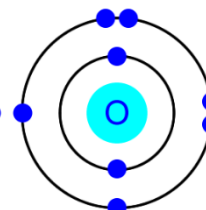
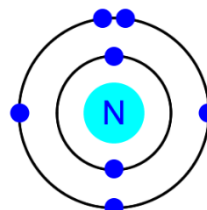
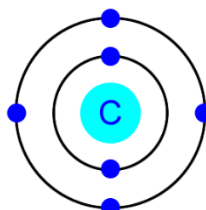
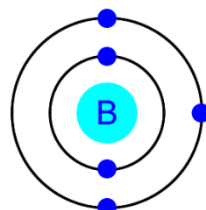
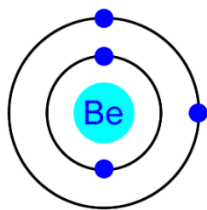
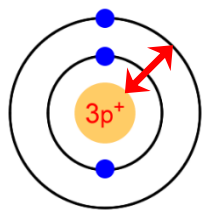


- Moving down a Group, there is an increase in the number of protons in the nucleus (increase in nuclear charge) and also an increase in the number of electron shells.

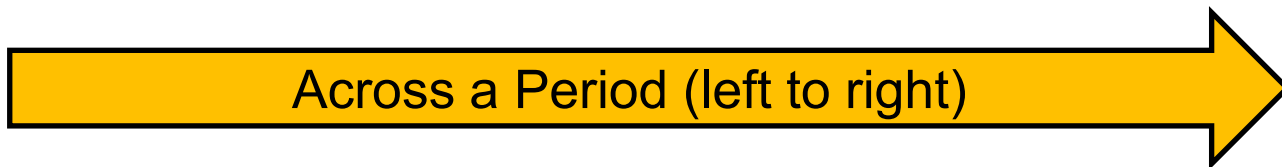
- Although there is an increase in nuclear charge, the increasing distance between the nucleus and valence electrons is more significant. The *electrostatic force of attraction* between the nucleus and electrons in the valence electron shell *decreases down a Group*.



Periodic Trends

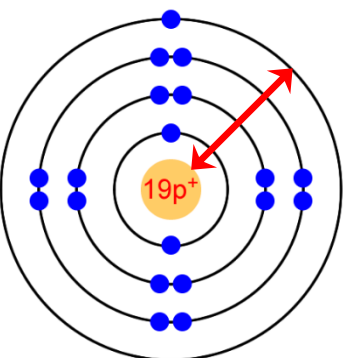


Down a Group

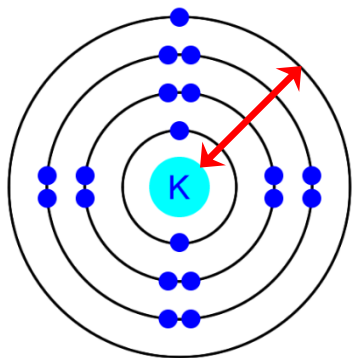
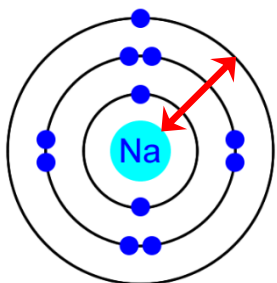
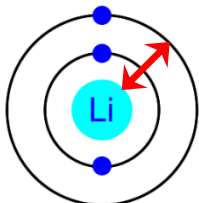


- Moving down a Group, there is an increase in the number of protons in the nucleus (increase in nuclear charge) and also an increase in the number of electron shells.

- Although there is an increase in nuclear charge, the increasing distance between the nucleus and valence electrons is more significant. The *electrostatic force of attraction* between the nucleus and electrons in the valence electron shell *decreases down a Group*.

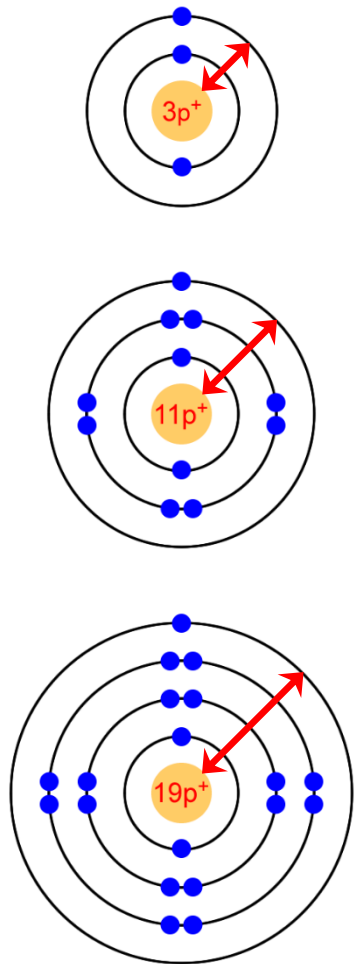


Periodic Trends



- Atomic radii *increase* from the top to the bottom of any Group of the Periodic Table.
- The number of protons in the nucleus of an atom (nuclear charge) and the number of electron shells around the nucleus of the atom *both increase* down a Group.

Periodic Trends

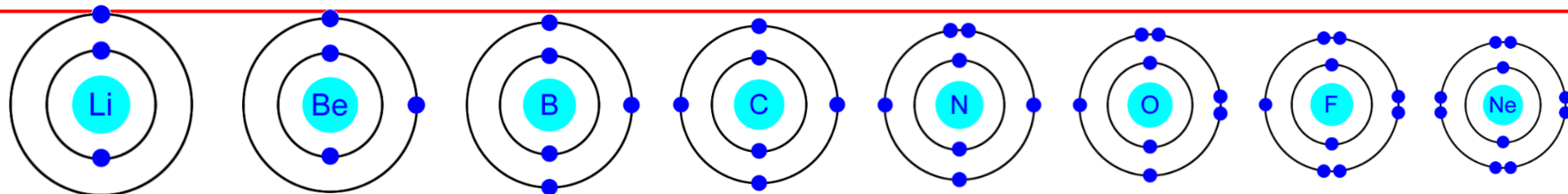


- Atomic radii *increase* from the top to the bottom of any Group of the Periodic Table.
- The number of protons in the nucleus of an atom (nuclear charge) and the number of electron shells around the nucleus of the atom *both increase* down a Group.

Periodic Trends

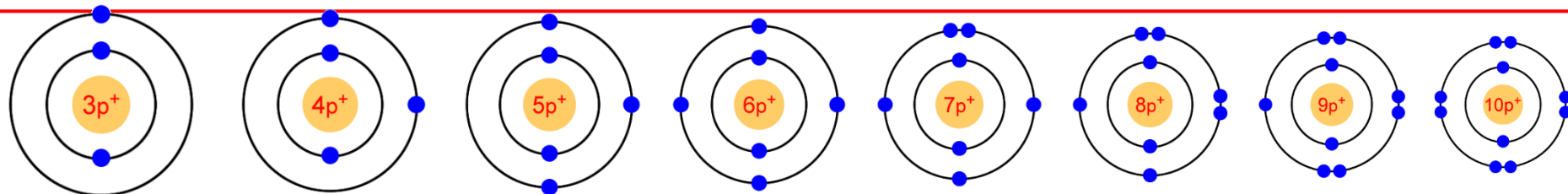
- An increase in nuclear charge means that there is an *increase in the electrostatic force of attraction* between the protons in the nucleus and electrons orbiting the nucleus. On its own, this variable would cause atomic radius to decrease down a Group.
- However, the addition of a new electron shell to the atoms is *more significant*, causing the atomic radius to increase. Also, with the addition of a new electron shell, electrons in the valence electron shell are more *shielded* from the electrostatic force of attraction of the nucleus by the electrons of the inner electron shells. Consequently, *atomic radius increases down a Group*.

Periodic Trends



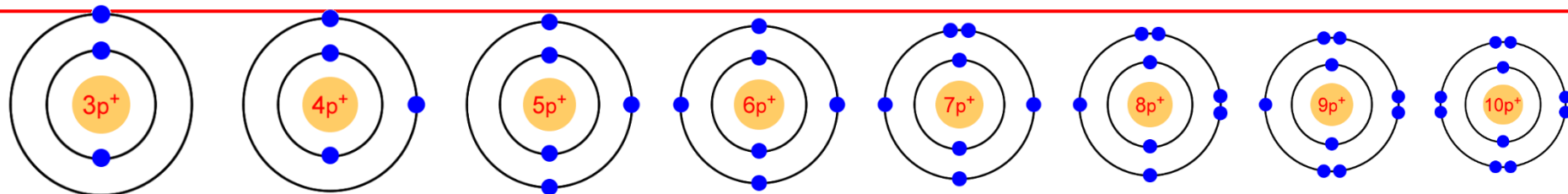
- Atomic radii *decrease* from the left-hand-side to the right-hand-side of any Period of the Periodic Table.
- The number of protons in the nucleus of an atom (nuclear charge) *increases* across a Period, while the number of electron shells is *constant* (remains the same).

Periodic Trends



- Atomic radii *decrease* from the left-hand-side to the right-hand-side of any Period of the Periodic Table.
- The number of protons in the nucleus of an atom (nuclear charge) *increases* across a Period, while the number of electron shells is *constant* (remains the same).

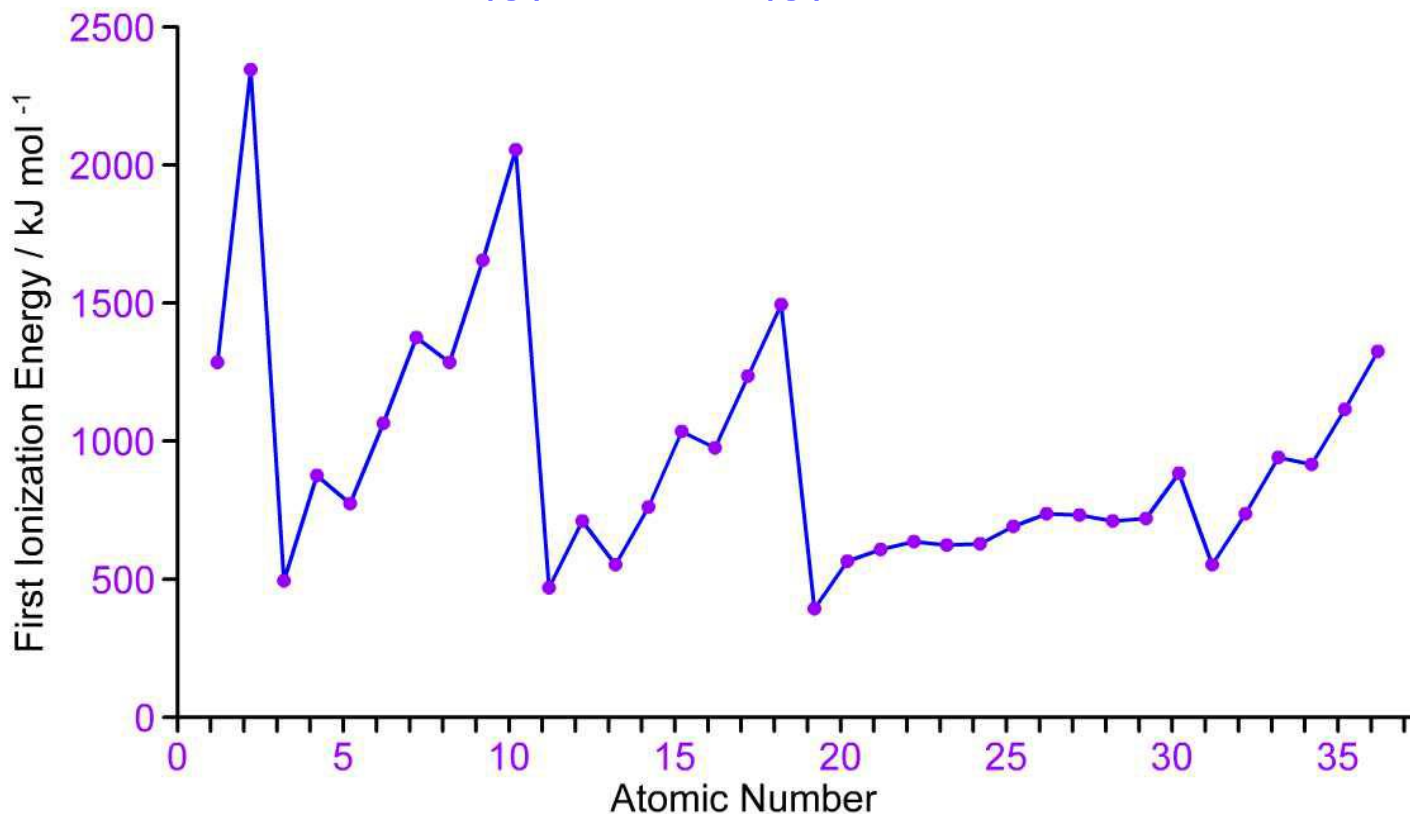
Periodic Trends



- An increase in nuclear charge means that there is an *increase in the electrostatic force of attraction* between the protons in the nucleus and electrons orbiting the nucleus.
- This increasing nuclear charge exerts a stronger electrostatic force of attraction on the orbiting electrons and causes a steady *decrease in atomic radius across a Period*.

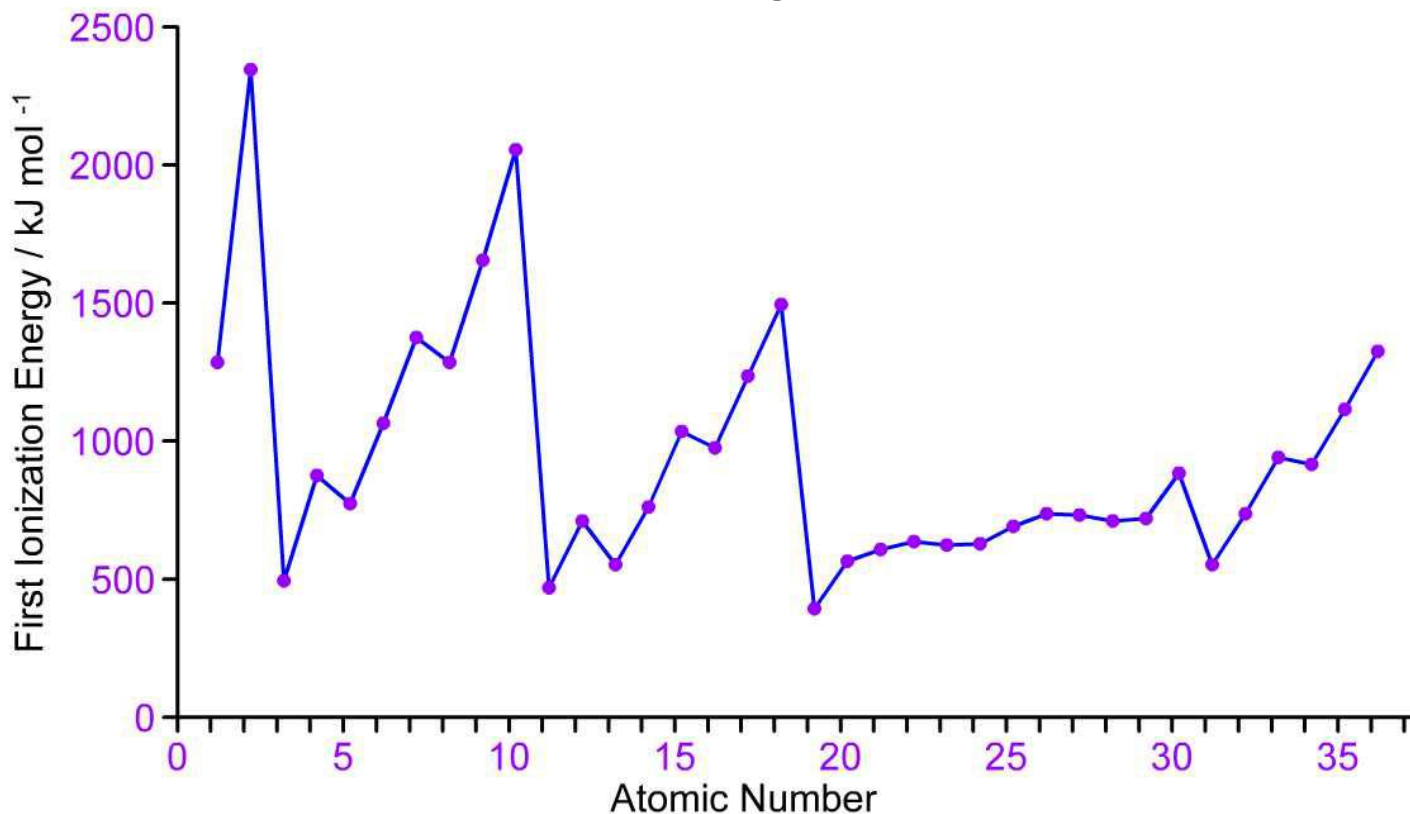
Periodic Trends

- First ionization energy is the energy required to convert 1 mole (6×10^{23}) of gaseous atoms into one mole (6×10^{23}) of unipositive (1+) gaseous ions.



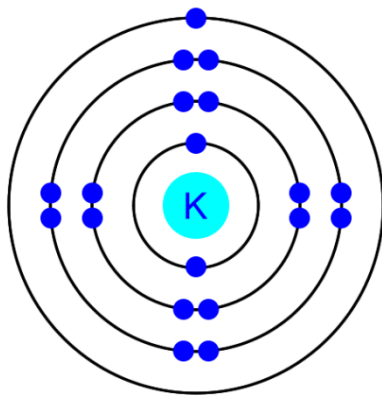
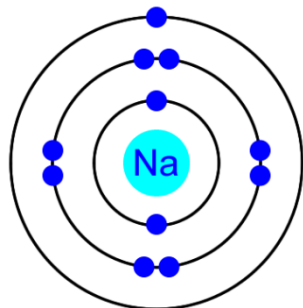
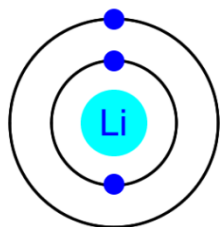
Periodic Trends

- Essentially, first ionization energy gives an indication of the amount of energy that is required to remove a single electron from the valence electron shell of a single atom.



Periodic Trends

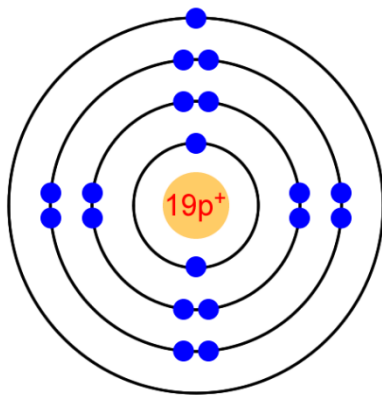
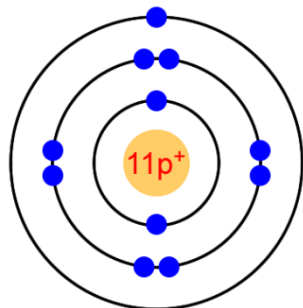
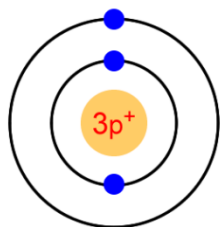
- First ionization energy *decreases down a Group*.



- The number of protons in the nucleus of an atom (nuclear charge) and the number of electron shells around the nucleus of the atom both *increase* down a Group.
- An increase in nuclear charge means that there is an *increase in the electrostatic force of attraction* between the protons in the nucleus and electrons orbiting the nucleus. On its own, this variable would cause first ionization energy to increase down a Group.

Periodic Trends

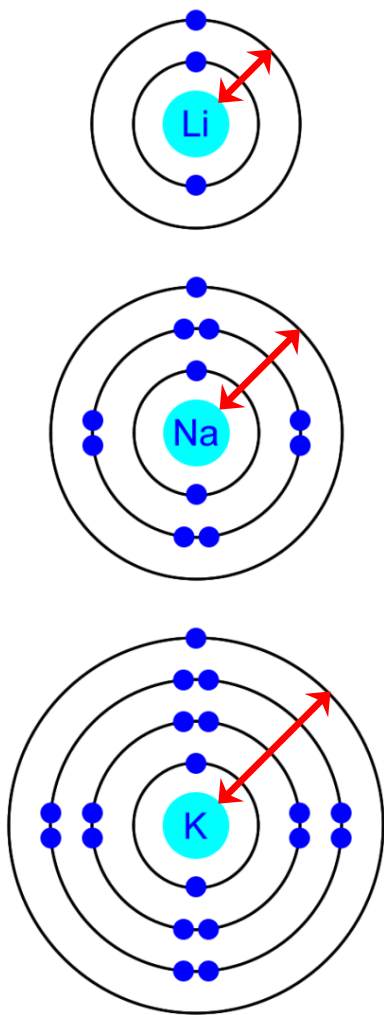
- First ionization energy *decreases down a Group*.



- The number of protons in the nucleus of an atom (nuclear charge) and the number of electron shells around the nucleus of the atom both *increase* down a Group.
- An increase in nuclear charge means that there is an *increase in the electrostatic force of attraction* between the protons in the nucleus and electrons orbiting the nucleus. On its own, this variable would cause first ionization energy to increase down a Group.

Periodic Trends

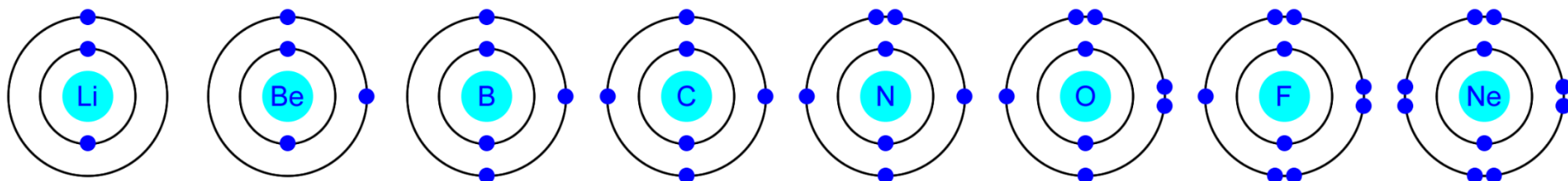
- First ionization energy *decreases down a Group*.



- However, the addition of a new electron shell to the atoms is *more significant*. The addition of a new electron shell means that electrons in the valence electron shell (lost during ionization) are **i) further** from the nucleus and are **ii) more shielded** from the attractive force of the nucleus by the electrons of the inner electron shells. These two effects combine to *reduce the electrostatic force of attraction* between the positively charged nucleus and negatively charged electrons in the valence electron shell. *Less energy* is required to remove an electron from the valence electron shell, therefore *first ionization energy decreases down a Group*.

Periodic Trends

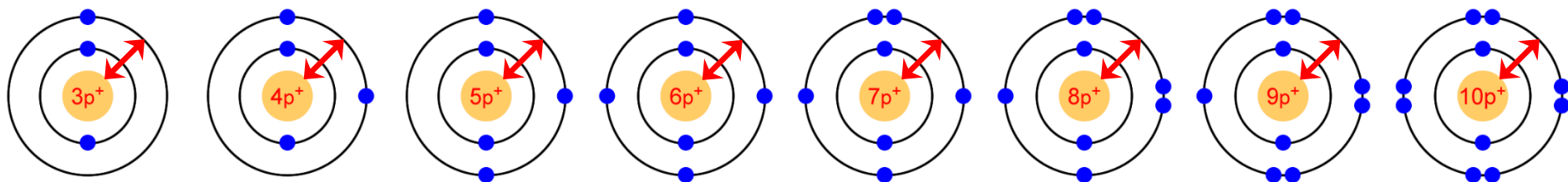
- First ionization energy *increases across a Period*.



- The number of protons in the nucleus of an atom (nuclear charge) *increases* across a Period, while the number of electron shells is *constant* (remains the same).
- An increase in nuclear charge means that there is an *increase in the electrostatic force of attraction* between the protons in the nucleus and electrons orbiting the nucleus.

Periodic Trends

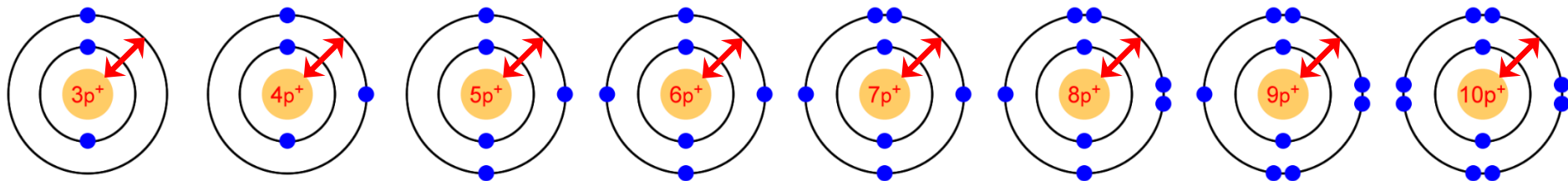
- First ionization energy *increases across a Period*.



- The number of protons in the nucleus of an atom (nuclear charge) *increases* across a Period, while the number of electron shells is *constant* (remains the same).
- An increase in nuclear charge means that there is an *increase in the electrostatic force of attraction* between the protons in the nucleus and electrons orbiting the nucleus.

Periodic Trends

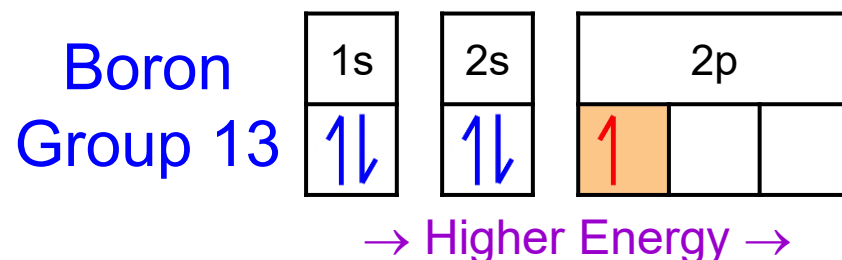
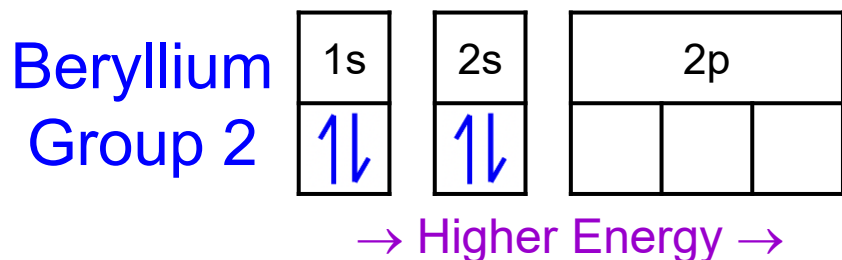
- First ionization energy *increases across a Period*.



- This increasing nuclear charge exerts a *stronger electrostatic force of attraction* on electrons in the valence electron shell of the atom (the electrons that are lost during ionization). Therefore, *more energy* is required to remove an electron from the valence electron shell of the atom, resulting in an *increase in first ionization energy across a Period*.

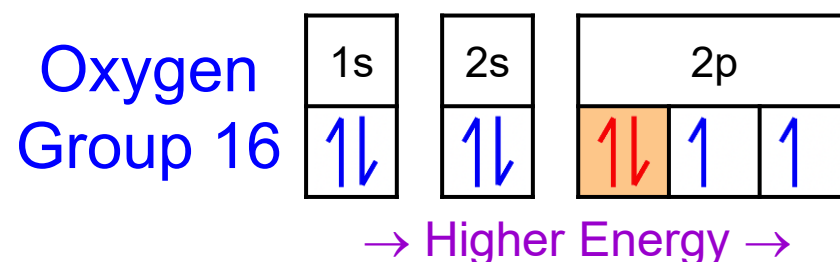
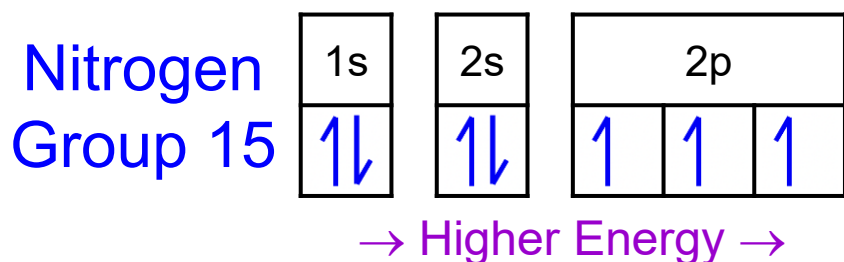
Periodic Trends

- First ionization energy *decreases* slightly between *Group 2 and Group 13* elements.
- Moving from Group 2 to Group 13, the additional electron enters a *p*-orbital of the same principle quantum shell.
- An electron in a *p*-orbital is *higher in energy* than an electron in the *s*-orbital of the same principle quantum shell. Consequently, *less energy* is required to remove the *p*-orbital electron (ionization) compared to an electron in the corresponding *s*-orbital, and first ionization energy *decreases* slightly between Group 2 and Group 13.

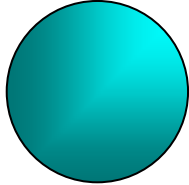
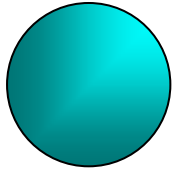
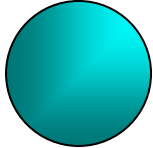
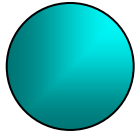
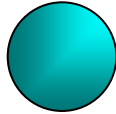




Periodic Trends

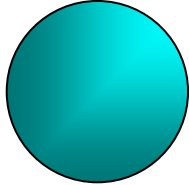
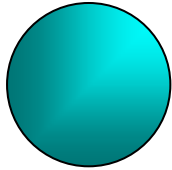
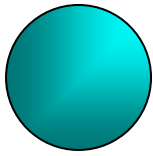
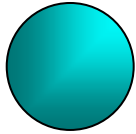
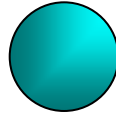


- First ionization energy *decreases* slightly between *Group 15* and *Group 16* elements.
- Moving from Group 15 to Group 16, the additional electron must spin pair with an existing electron in one of the atom's *p*-orbitals.
- An *electrostatic force of repulsion* between the two spin paired electrons that share the same *p*-orbital means that *less energy* is required to remove (ionization) an electron from the *p*-orbital, and first ionization energy *decreases* slightly between Group 15 and Group 16.



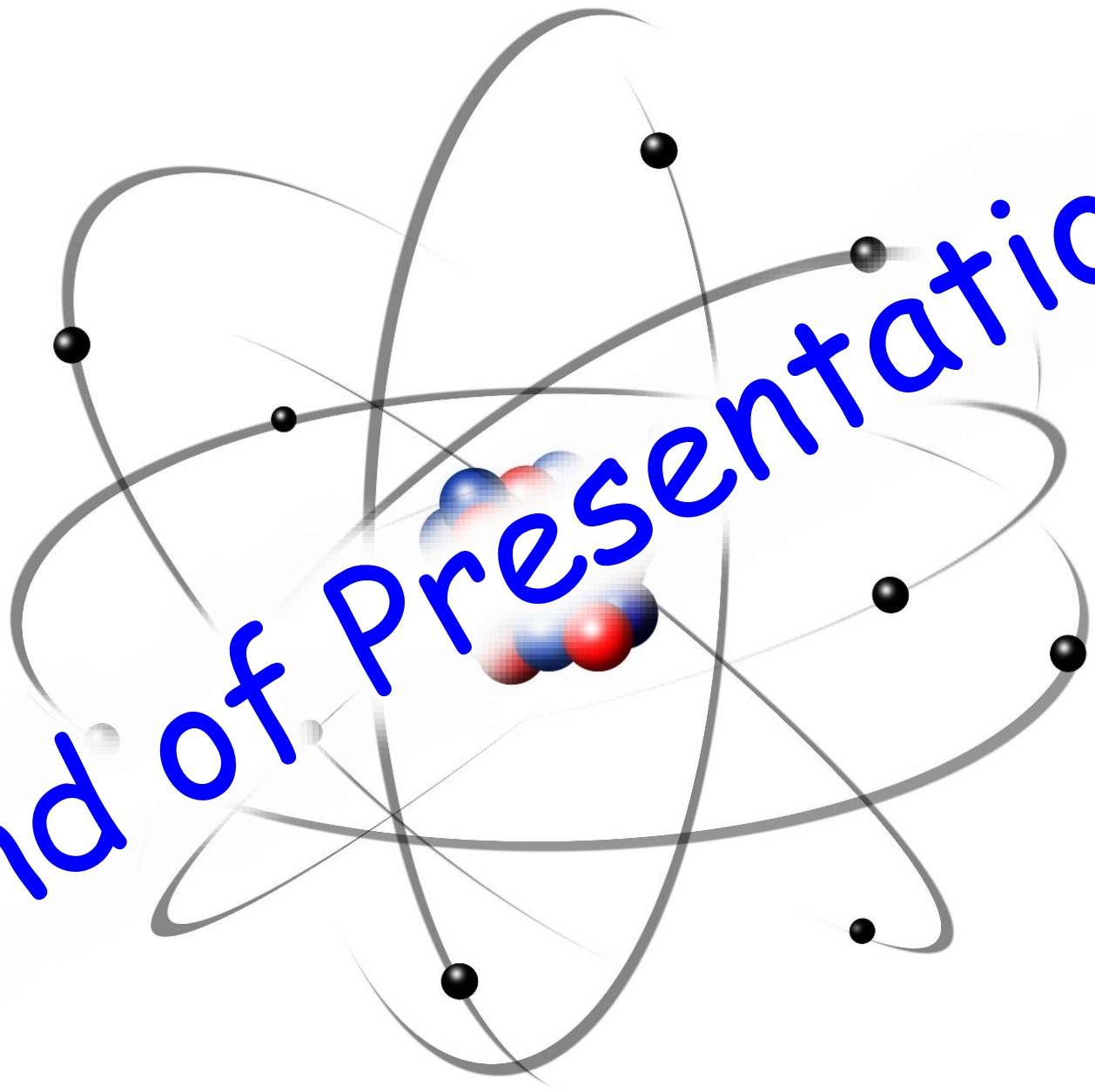
Periodic Trends – Summary

Elements of Third Period	Na Group 1	Mg Group 2	Al Group 13	Si Group 14	P Group 15	S Group 16	Cl Group 17
Atomic Number	11	12	13	14	15	16	17
Atomic Radius / nm	0.157 	0.136 	0.125 	0.117 	0.110 	0.104 	0.099 
First Ionization Energy / kJ mol ⁻¹	494	736	577	786	1060	1000	1260

Periodic Trends – Summary

Elements of Third Period	Na Group 1	Mg Group 2	Al Group 13	Si Group 14	P Group 15	S Group 16	Cl Group 17
Atomic Number	<ul style="list-style-type: none"> • Number of protons within the nucleus of the atom increases. • Positive charge within the nucleus of the atom increases. • Electrostatic force of attraction between the positive nucleus and negative electrons orbiting the nucleus increases. 						
Atomic Radius / nm	0.157 	0.136 	0.125 	0.117 	0.110 	0.104 	0.099 
First Ionization Energy / kJ mol ⁻¹	494	736	577	786	1060	1000	1260

End of Presentation



Atomic Structure

Presentation on
Atomic Structure

by Dr. Chris Slatter

christopher_john_slatter@nygh.edu.sg

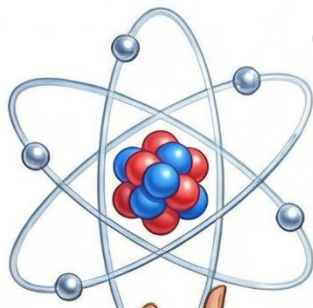
Nanyang Girls' High School

2 Linden Drive

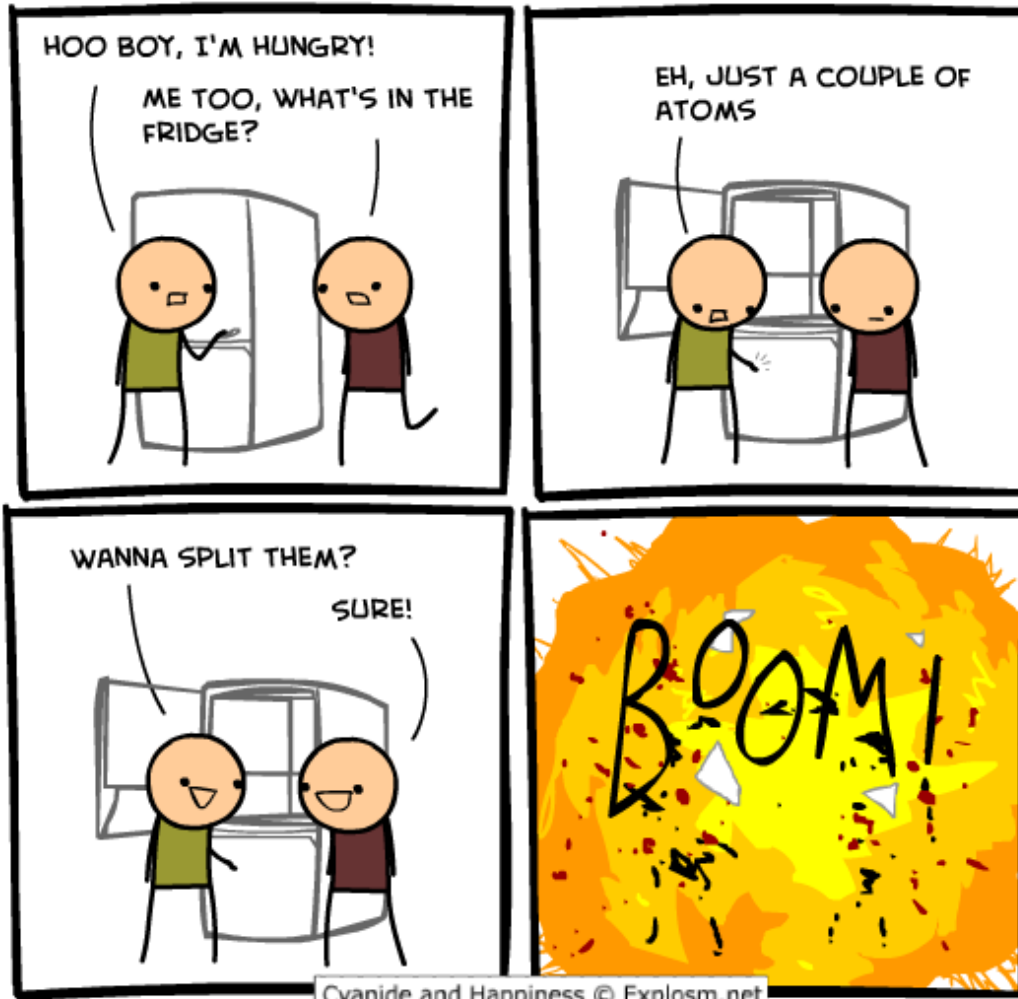
Singapore

288683

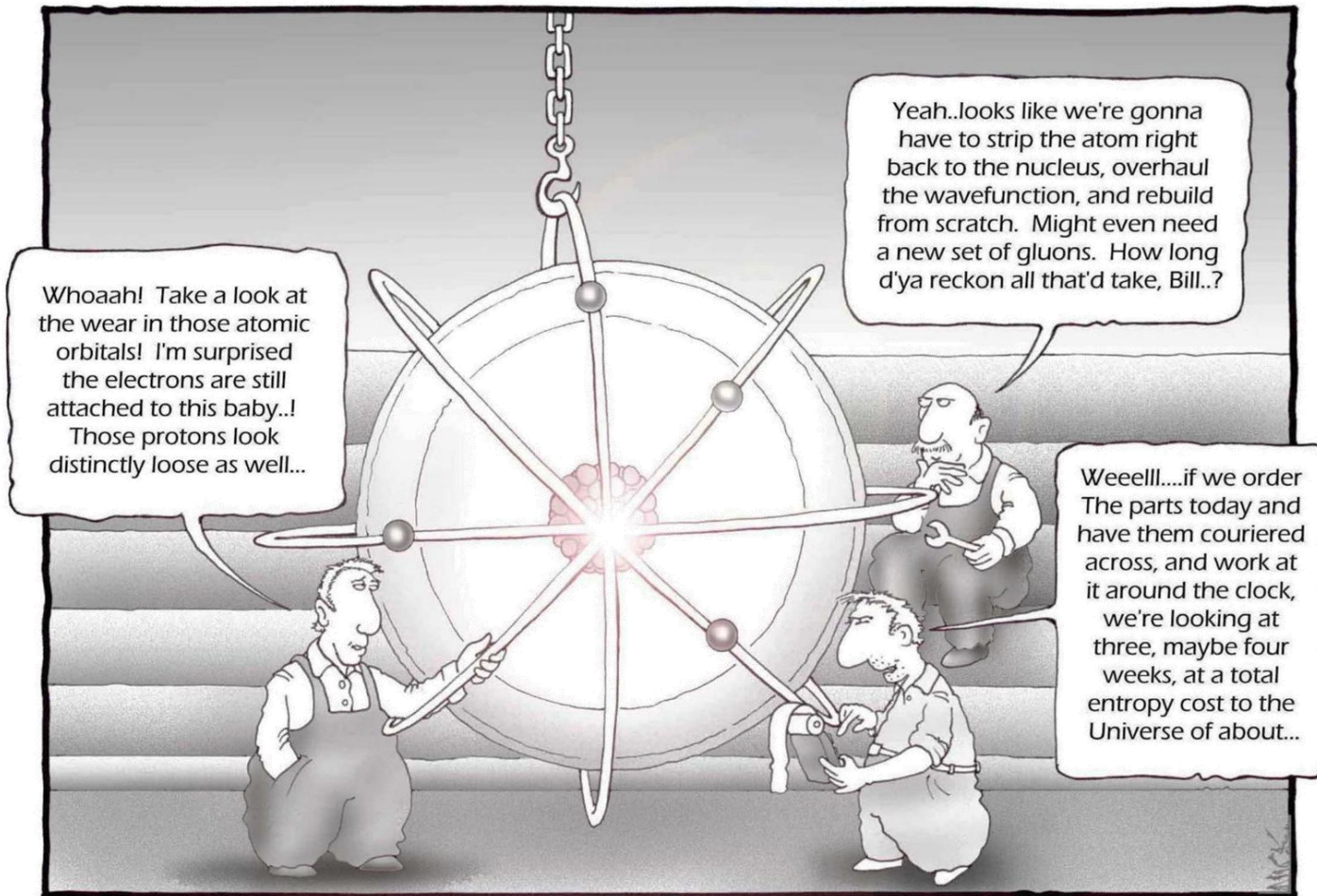
Updated for alignment with the 2017
A' Level Chemistry Periodic Table
8th January 2017



Atomic Structure



Atomic Structure



Quantum mechanics.