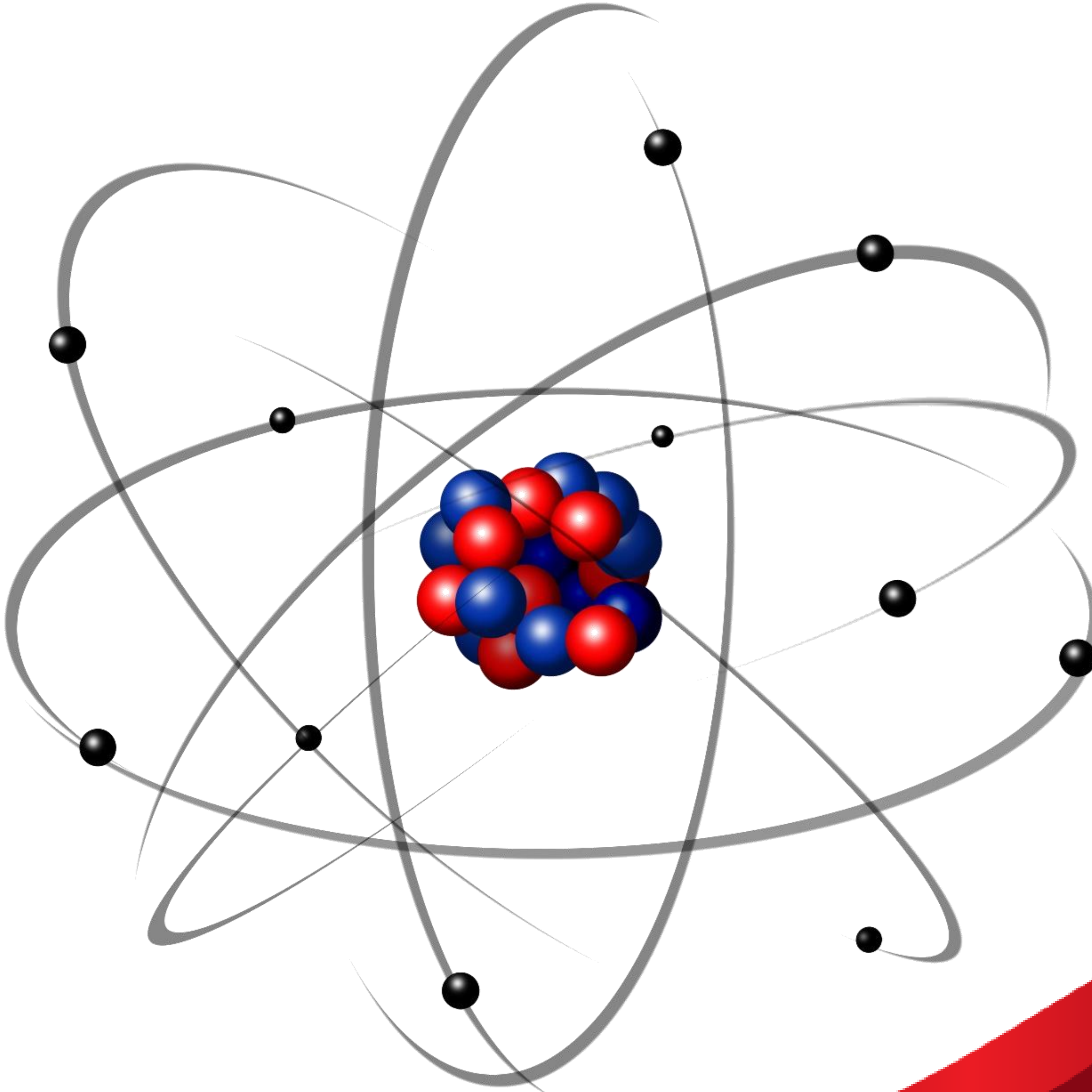
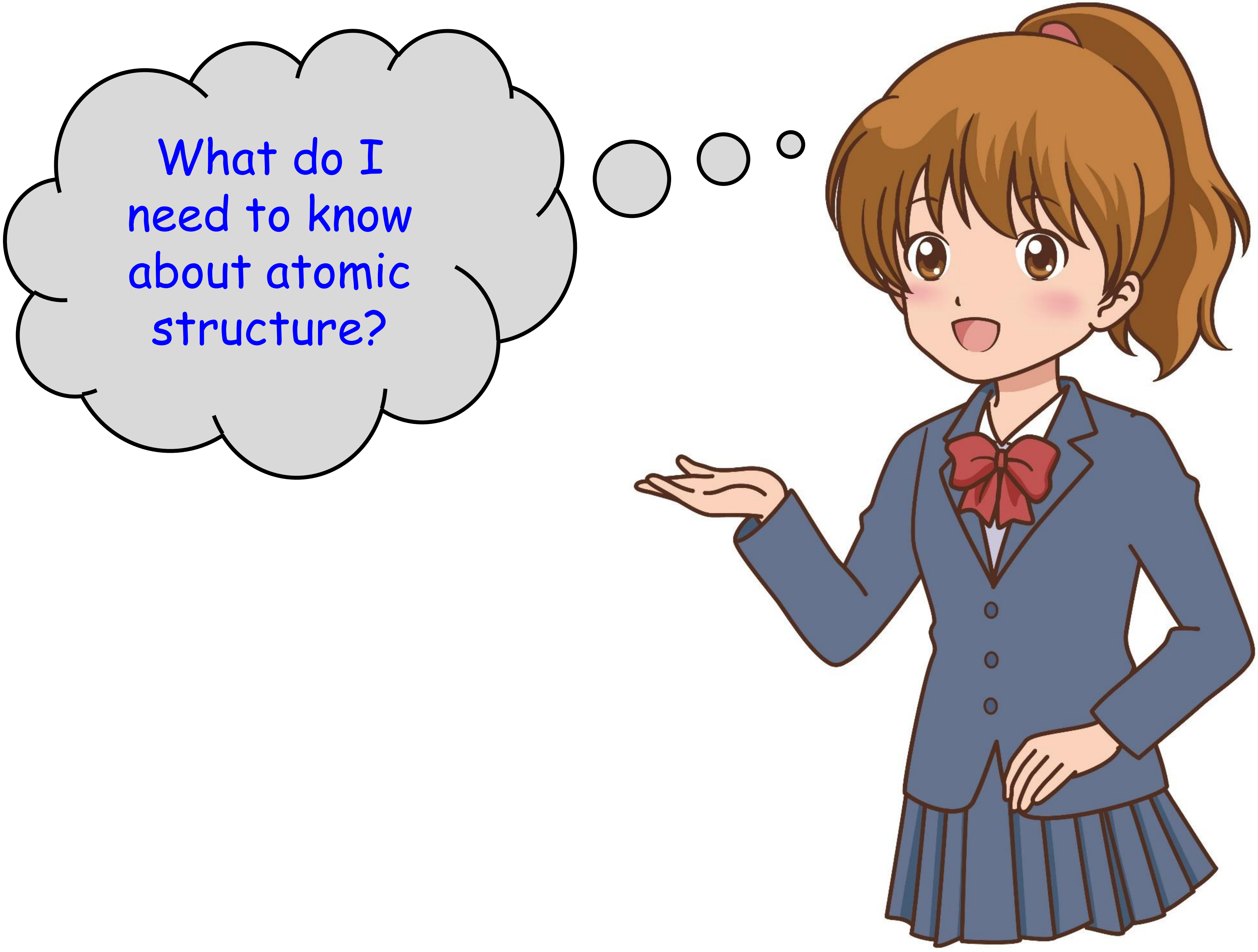


Introduction to Atomic Structure


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 $^{12}_{6}C$.

(e) Define the term isotope.

(f) Deduce the numbers of protons, neutrons and electrons in atoms and ions given proton and nucleon numbers.

Periodic Table of the Chemical Elements (2017)

Key

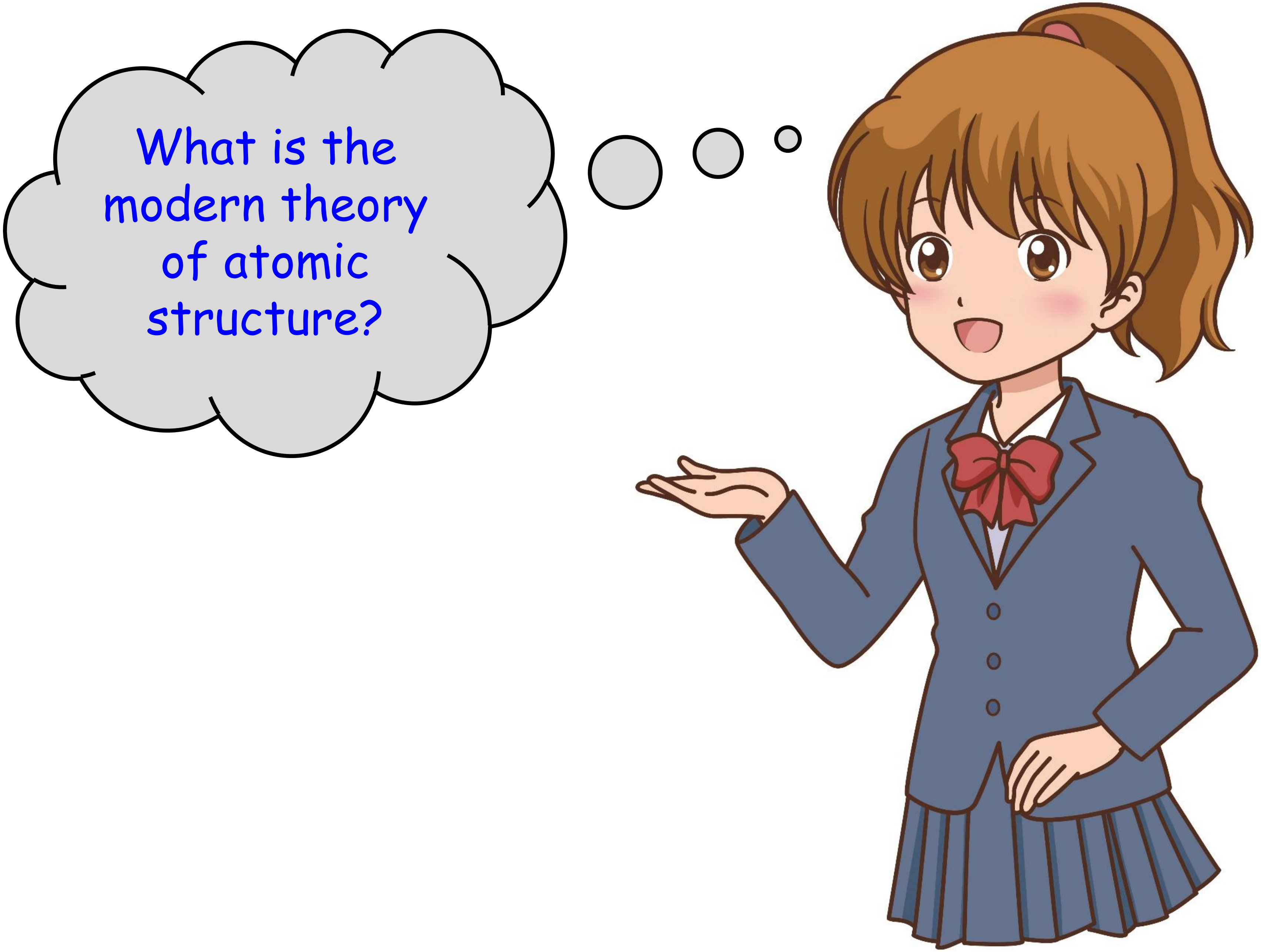
atomic number atomic symbol name relative atomic mass

Group																	
1	2											13	14	15	16	17	18
							1 H hydrogen 1.0										2 He helium 4.0
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	4	5	6	7	8	9	10	11	12	13 Al aluminium 27.0	14 Si silicon 28.1	15 P phosphorus 31.0	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 –		114 Fl flerovium –		116 Lv livermorium –		

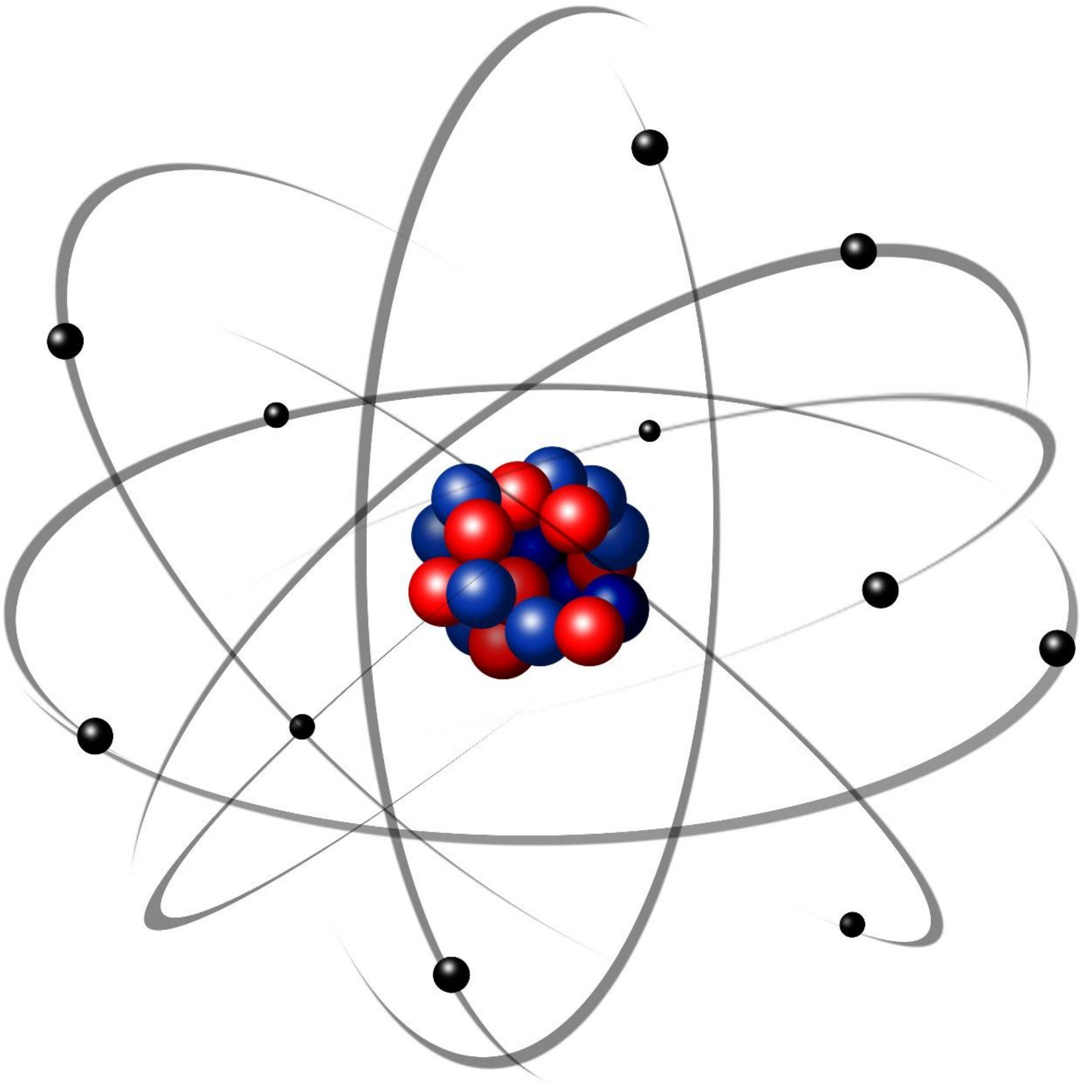
lanthanoids	57 La lanthanum 138.9	58 Ce cerium 140.1	59 Pr praseodymium 140.9	60 Nd neodymium 144.2	61 Pm promethium –	62 Sm samarium 150.4	63 Eu europium 152.0	64 Gd gadolinium 157.3	65 Tb terbium 158.9	66 Dy dysprosium 162.5	67 Ho holmium 164.9	68 Er erbium 167.3	69 Tm thulium 168.9	70 Yb ytterbium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

A' Level Periodic Table of the Chemical Elements

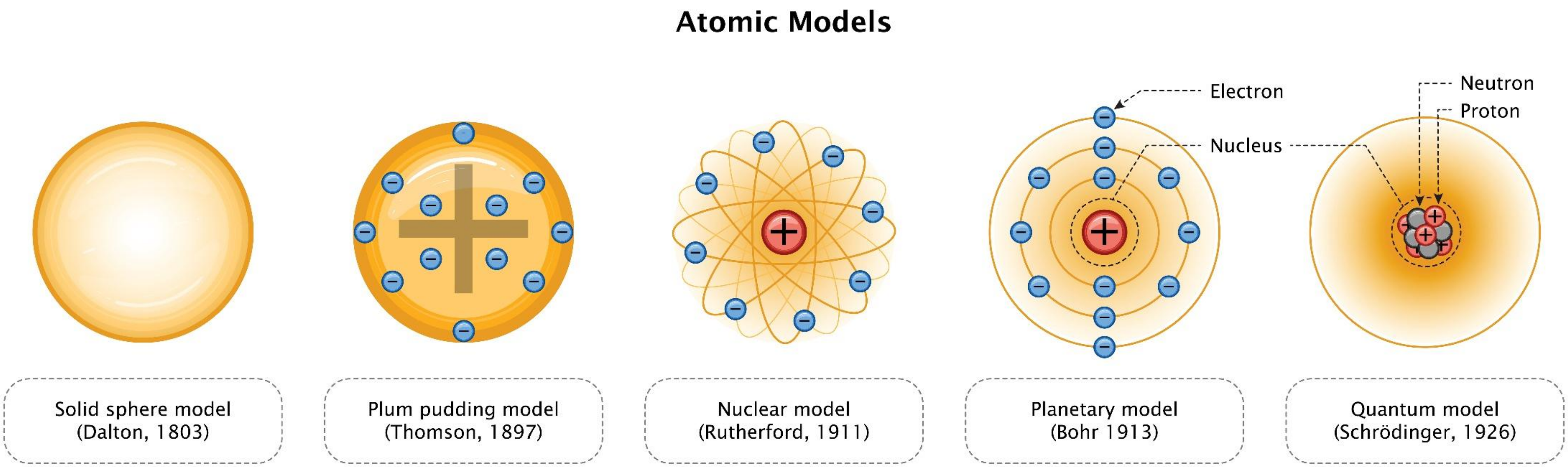
- The *Groups* are numbered in Arabic numerals from 1 to 18.
- The *relative atomic masses* are all written to one decimal place.
- The *atomic number* is written *above* the symbol of the chemical element, while the *relative atomic mass* is written *below* the symbol of the chemical element.


What is the modern theory of atomic structure?

Introduction to Atomic Structure



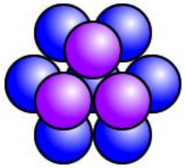
Introduction to Atomic Structure



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.

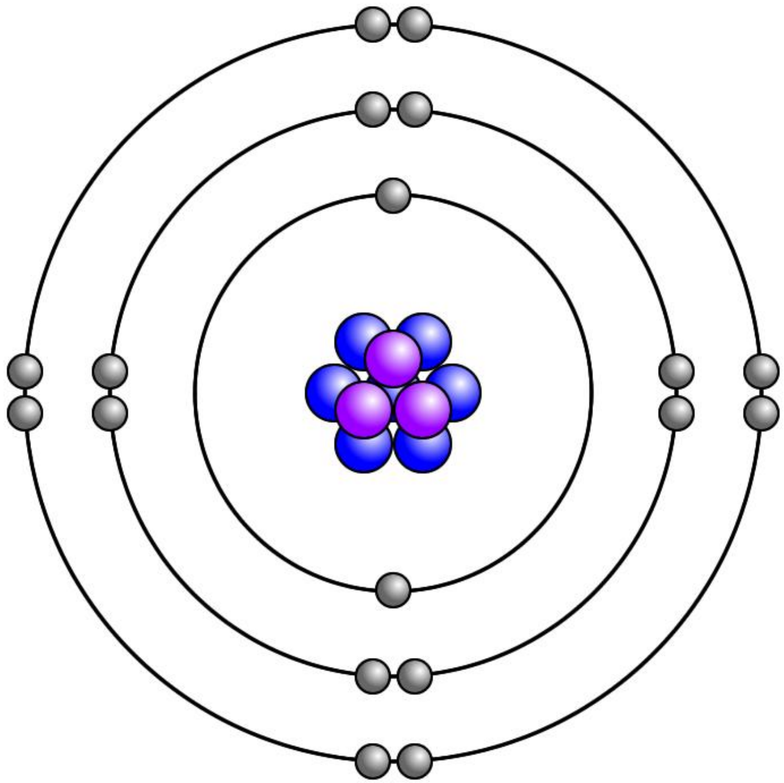
Introduction to Atomic Structure



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

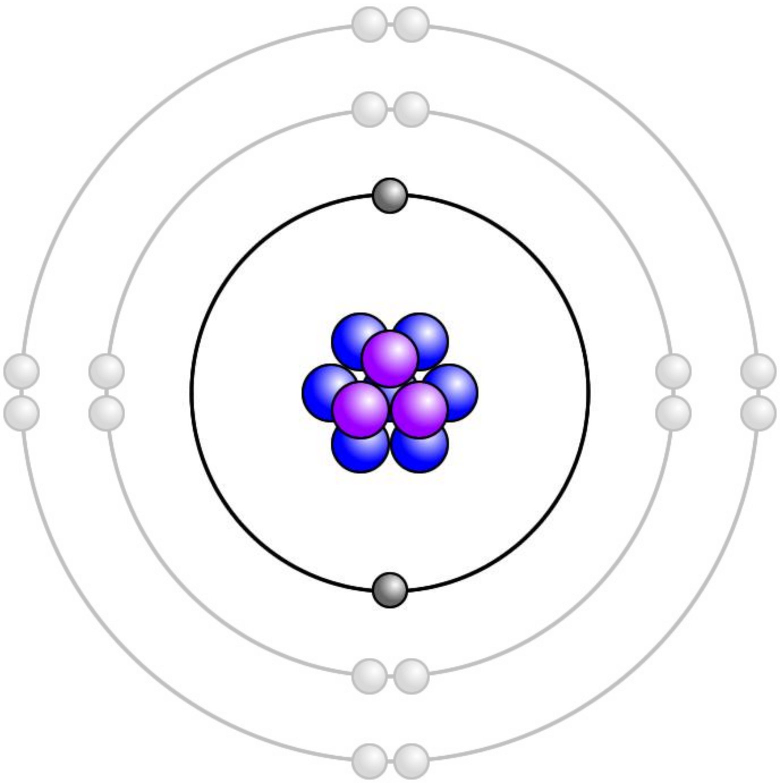
Introduction to Atomic Structure



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

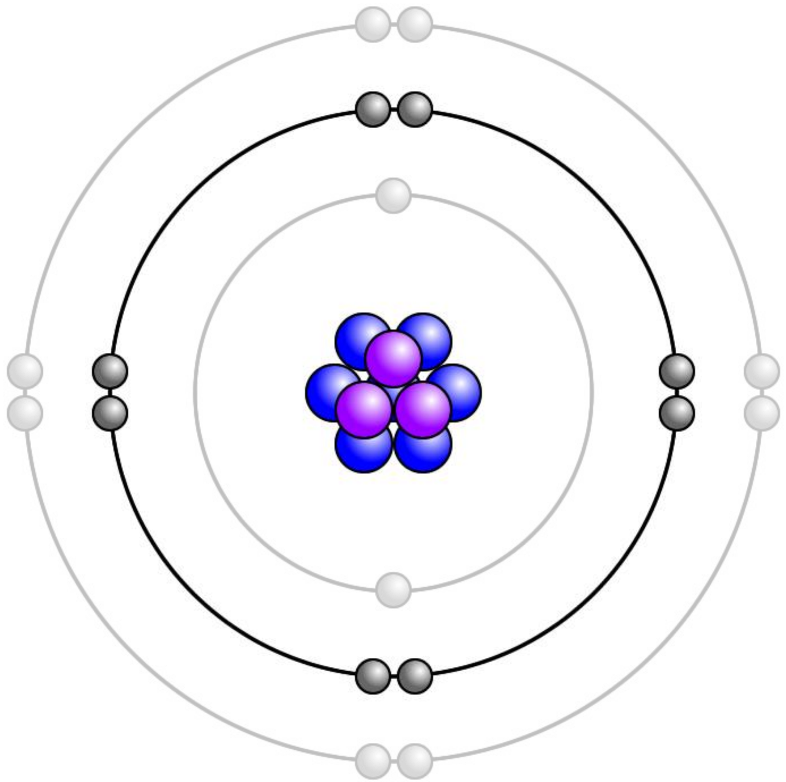
- Relative mass of an electron = $^{1}/_{1836}$
- Relative charge on an electron = –1

Introduction to Atomic Structure



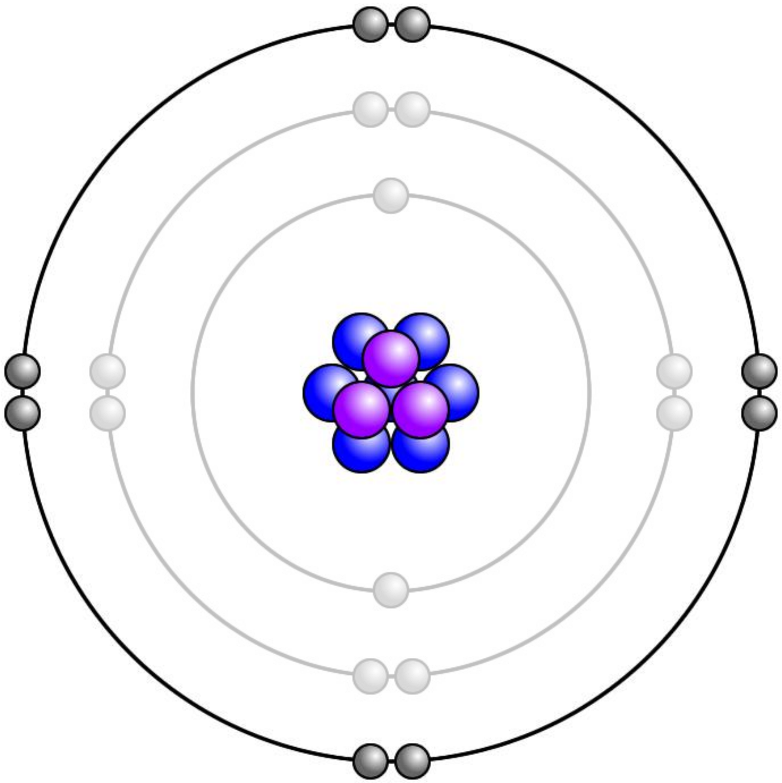
- The *inner electron shell* can hold a maximum number of **2** electrons.
- The *second electron shell* can hold a maximum number of **8** electrons.
- The *third electron shell* can hold a maximum number of **8** electrons.

Introduction to Atomic Structure



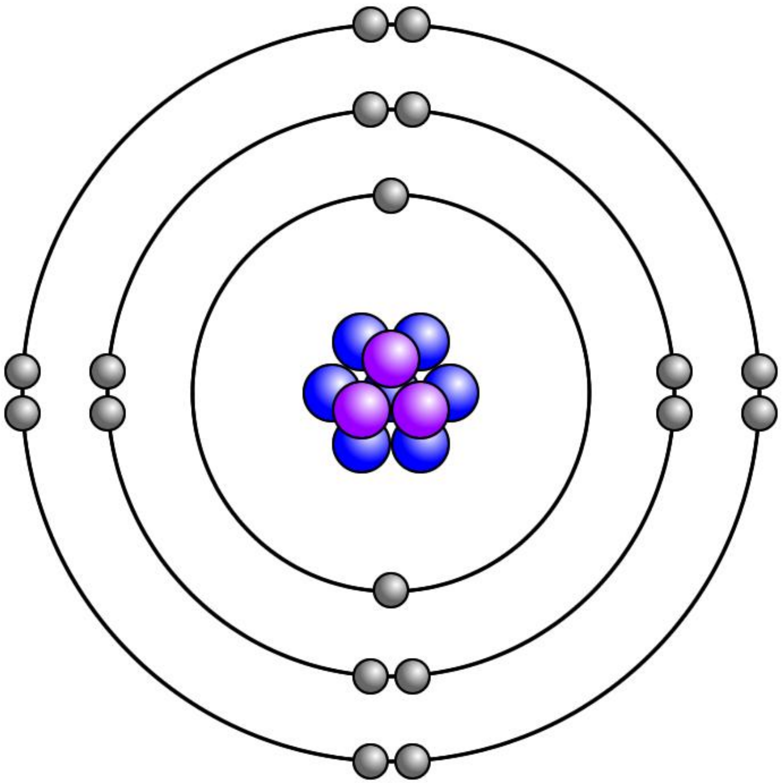
- The *inner electron shell* can hold a maximum number of **2** electrons.
- The *second electron shell* can hold a maximum number of **8** electrons.
- The *third electron shell* can hold a maximum number of **8** electrons.

Introduction to Atomic Structure



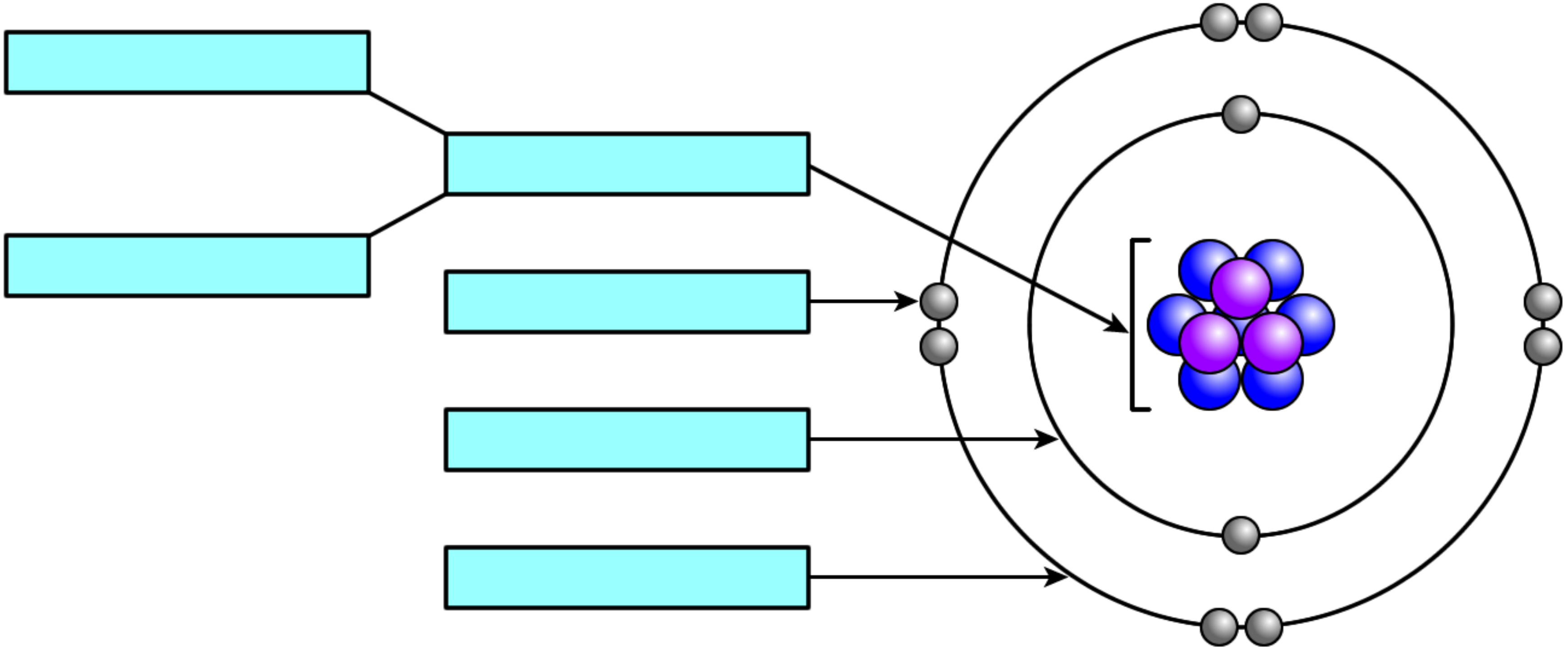
- The *inner electron shell* can hold a maximum number of **2** electrons.
- The *second electron shell* can hold a maximum number of **8** electrons.
- The *third electron shell* can hold a maximum number of **8** electrons.

Introduction to Atomic Structure

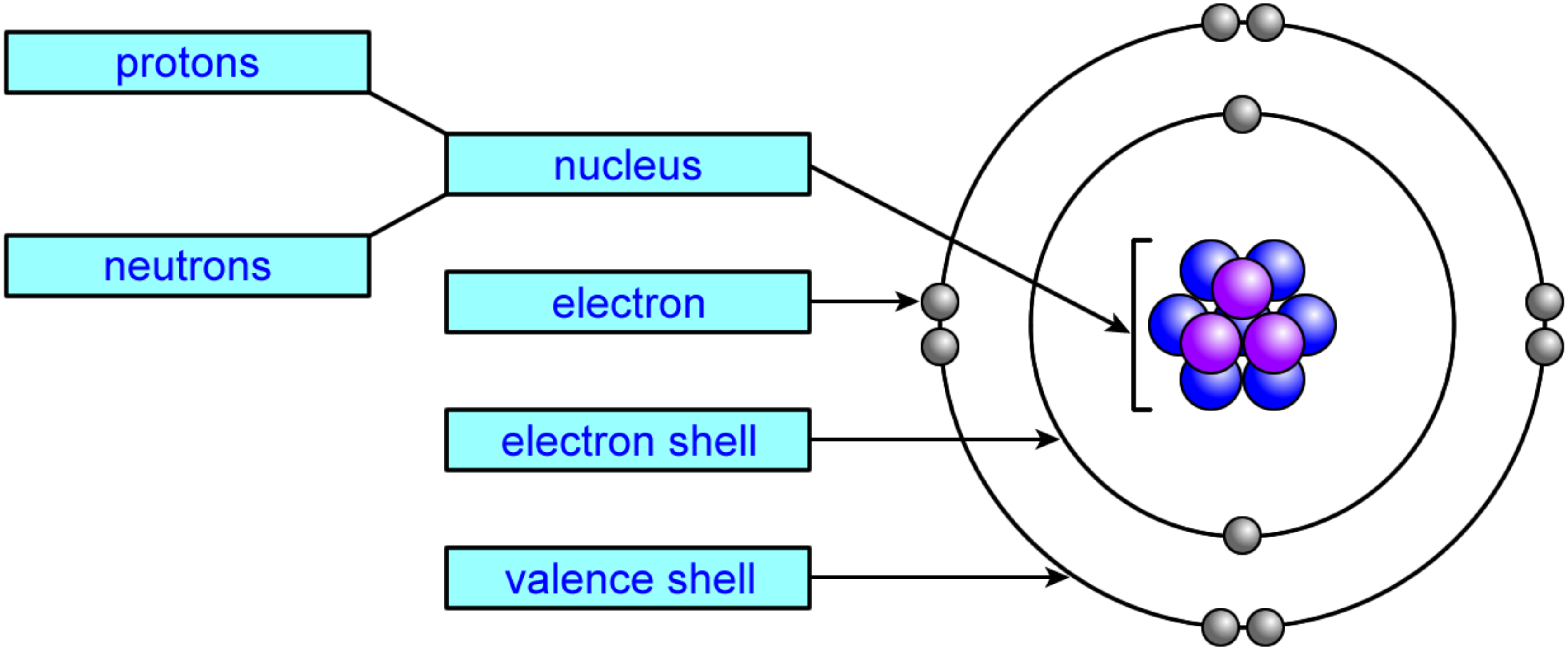


- Atoms are *electrically neutral* because the number of electrons (which carry a charge of –1) orbiting the nucleus equals the number of protons (which carry a charge of +1) located in the nucleus.

Introduction to Atomic Structure



Introduction to Atomic Structure



Introduction to Atomic Structure

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

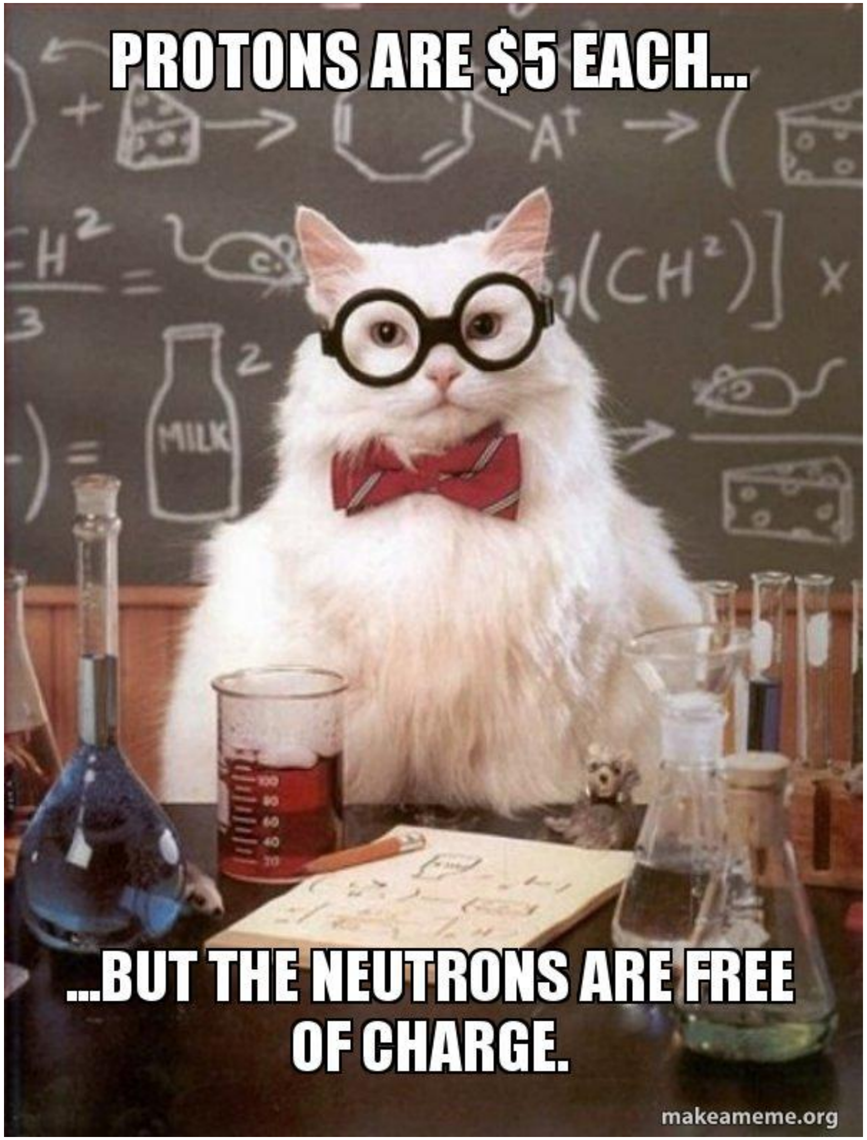
Particle	Particle's Location in the Atom	Particle's Relative Charge	Particle's Relative Mass*
Proton	Nucleus	+1	
Neutron	Nucleus	0	
Electron	Orbits the Nucleus	–1	

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	$^{1}/_{1836}$

*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

$$^{7}_{3}\mathrm{Li}$$

- 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

Nuclide Notation

- Mass (Nucleon) Number
  number of protons + neutrons ↓

$$^{19}_{9}\text{F}$$

- Symbol of the element ←
  F = *fluorine*

↑
- 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

Nuclide Notation

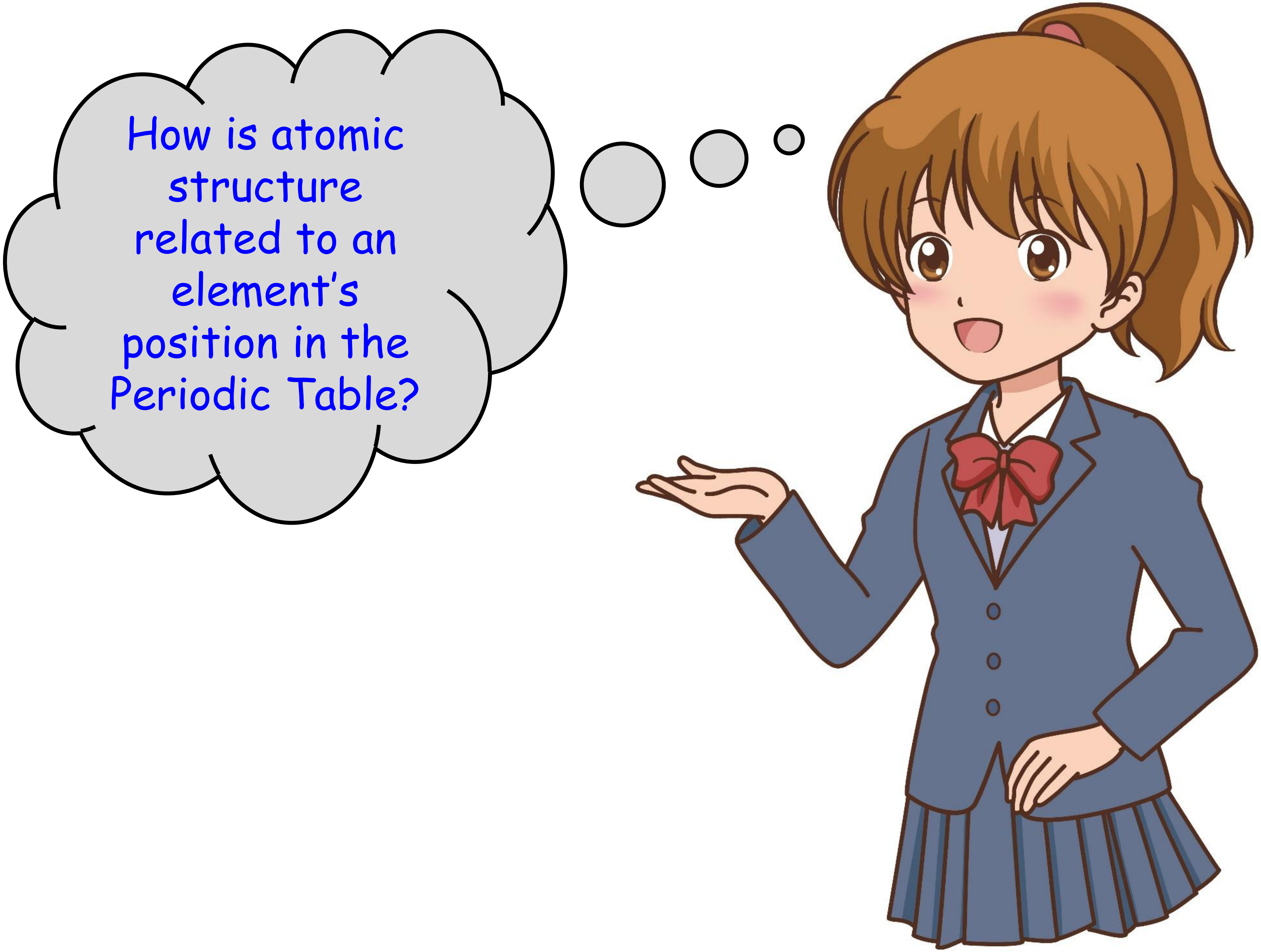
$$^{23}_{11}\mathrm{Na}$$

- Symbol of the element
  Na = *sodium*

- Atomic (Proton) Number
  number of protons

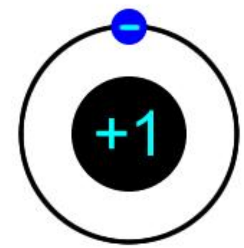
- number of *neutrons* = mass number – atomic number
  number of neutrons = 23 – 11 = **12**

- 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 $^{1}_{1}H$ Atomic Number 1

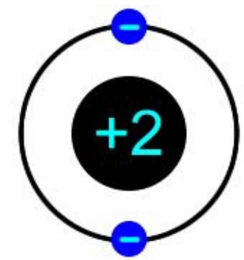
Period	Group 1	Group 2	Group 13	Group 14	Group 15	Group 16	Group 17	Group 18
1	1 Hydrogen H							2 Helium He
2	3 Lithium Li	4 Beryllium Be	5 Boron B	6 Carbon C	7 Nitrogen N	8 Oxygen O	9 Fluorine F	10 Neon Ne
3	11 Sodium Na	12 Magnesium Mg	13 Aluminium Al	14 Silicon Si	15 Phosphorus P	16 Sulphur S	17 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 electron configuration of hydrogen = 1
- The Mass Number equals number of protons (1) + number of neutrons (0) = 1

Helium $^{4}_{2}He$ Atomic Number 2

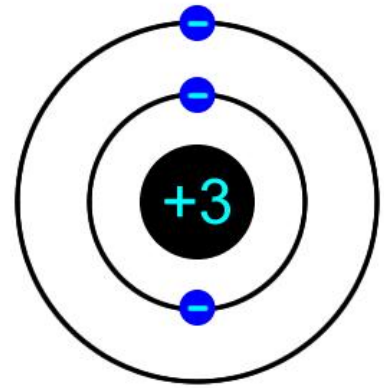
Period \ Group	1	2	13	14	15	16	17	18
1	1 Hydrogen H							2 Helium He
2	3 Lithium Li	4 Beryllium Be	5 Boron B	6 Carbon C	7 Nitrogen N	8 Oxygen O	9 Fluorine F	10 Neon Ne
3	11 Sodium Na	12 Magnesium Mg	13 Aluminium Al	14 Silicon Si	15 Phosphorus P	16 Sulphur S	17 Chlorine Cl	18 Argon Ar
4	19 Potassium 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 shell of helium = 2 (complete)
- The electron configuration of helium = 2
- The Mass Number equals number of protons (2) + number of neutrons (2) = 4

Lithium $^{7}_{3}Li$ Atomic Number 3

Period	Group 1	Group 2	Group 13	Group 14	Group 15	Group 16	Group 17	Group 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						

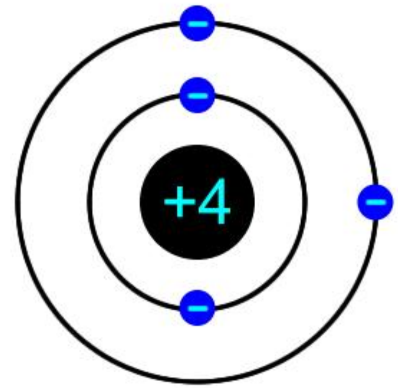


- The Atomic Number equals the number of protons = 3
- For a neutral atom, the number electrons equals the number of protons = 3
- The number of electron shell equals the Period number = 2
- The number of valence electrons can be derived from the Group number (**1**) = 1
- The electron configuration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

Beryllium $^{9}_{4}Be$ Atomic Number 4

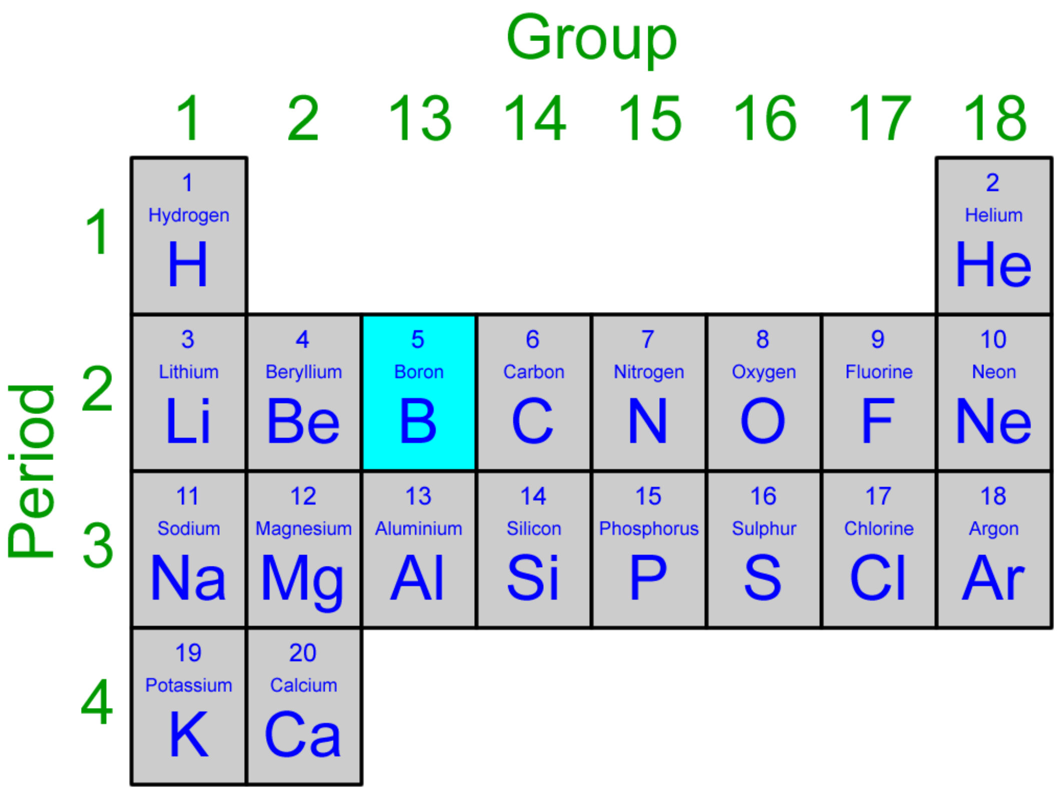
Group

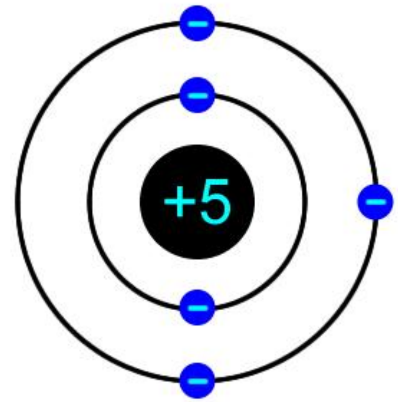
Period	1	2	13	14	15	16	17	18
1	1 Hydrogen H							2 Helium He
2	3 Lithium Li	4 Beryllium Be	5 Boron B	6 Carbon C	7 Nitrogen N	8 Oxygen O	9 Fluorine F	10 Neon Ne
3	11 Sodium Na	12 Magnesium Mg	13 Aluminium Al	14 Silicon Si	15 Phosphorus P	16 Sulphur S	17 Chlorine Cl	18 Argon Ar
4	19 Potassium K	20 Calcium 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 electron configuration of beryllium = 2,2
- The Mass Number equals number of protons (4) + number of neutrons (5) = 9

Boron $^{11}_{5}B$ Atomic Number 5

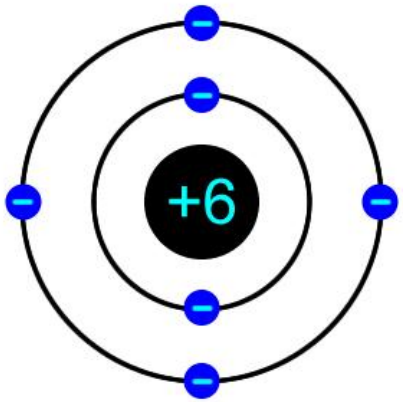




- The Atomic Number equals the number of protons = 5
- For a neutral atom, the number electrons equals the number of protons = 5
- The number of electron shell equals the Period number = 2
- The number of valence electrons can be derived from the Group number (1**3**) = 3
- The electron configuration of boron = 2,3
- The Mass Number equals number of protons (5) + number of neutrons (6) = 11

Carbon $^{12}_{6}C$ Atomic Number 6

Period \ 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						

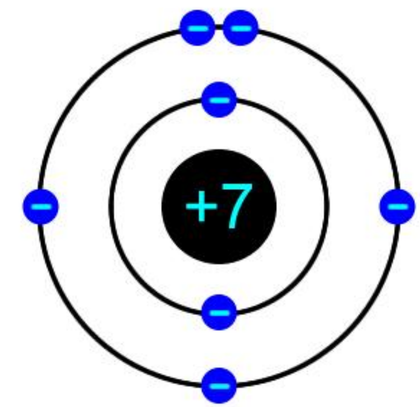


- The Atomic Number equals the number of protons = 6
- For a neutral atom, the number electrons equals the number of protons = 6
- The number of electron shell equals the Period number = 2
- The number of valence electrons can be derived from the Group number (1**4**) = 4
- The electron configuration of carbon = 2,4
- The Mass Number equals number of protons (6) + number of neutrons (6) = 12

Nitrogen $^{14}_{7}\mathrm{N}$ Atomic Number 7

Group

Period	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						

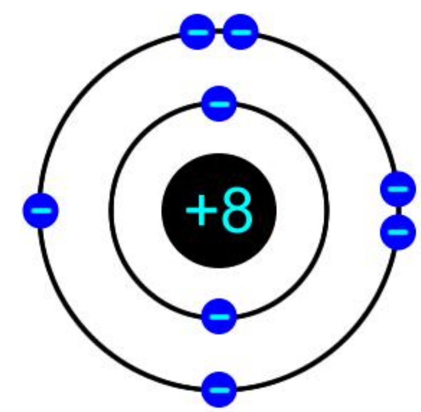


- The Atomic Number equals the number of protons = 7
- For a neutral atom, the number electrons equals the number of protons = 7
- The number of electron shell equals the Period number = 2
- The number of valence electrons can be derived from the Group number (1**5**) = 5
- The electron configuration of nitrogen = 2,5
- The Mass Number equals number of protons (7) + number of neutrons (7) = 14

Oxygen $^{16}_{8}O$ Atomic Number 8

Group

Period	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						

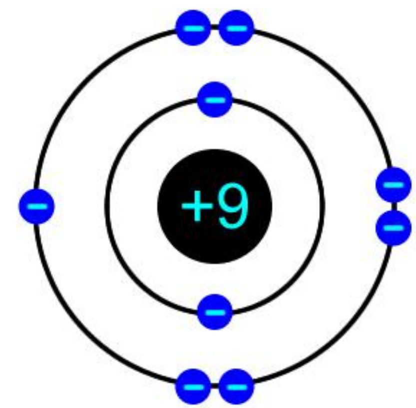


- The Atomic Number equals the number of protons = 8
- For a neutral atom, the number electrons equals the number of protons = 8
- The number of electron shell equals the Period number = 2
- The number of valence electrons can be derived from the Group number (1**6**) = 6
- The electron configuration of oxygen = 2,6
- The Mass Number equals number of protons (8) + number of neutrons (8) = 16

Fluorine $^{19}_{9}F$ Atomic Number 9

Group

Period	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						

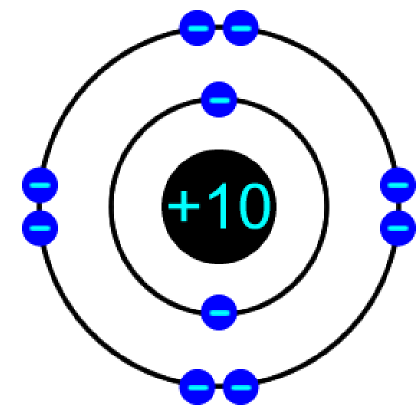


- The Atomic Number equals the number of protons = 9
- For a neutral atom, the number electrons equals the number of protons = 9
- The number of electron shell equals the Period number = 2
- The number of valence electrons can be derived from the Group number (1**7**) = 7
- The electron configuration of fluorine = 2,7
- The Mass Number equals number of protons (9) + number of neutrons (10) = 19

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

Group

Period	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						

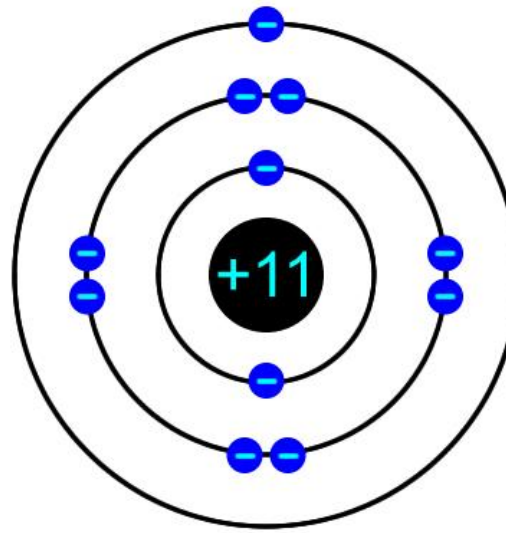


- The Atomic Number equals the number of protons = 10
- For a neutral atom, the number electrons equals the number of protons = 10
- The number of electron shell equals the Period number = 2
- The number of valence electrons can be derived from the Group number (1**8**) = 8 (complete)
- The electron configuration of neon = 2,8
- The Mass Number equals number of protons (10) + number of neutrons (10) = 20

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

Group

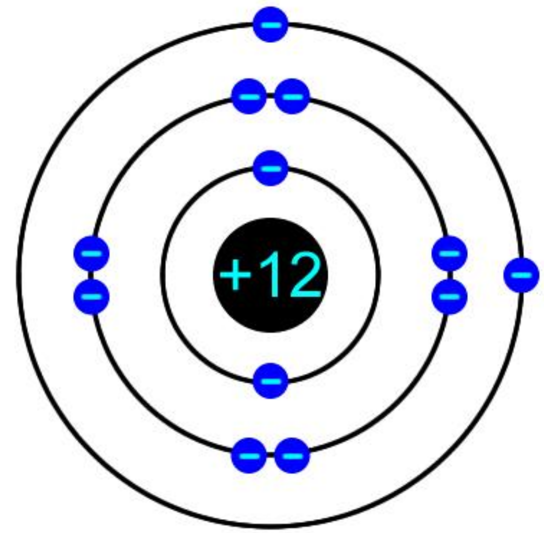
Period	1	2	13	14	15	16	17	18
1	1 Hydrogen H							2 Helium He
2	3 Lithium Li	4 Beryllium Be	5 Boron B	6 Carbon C	7 Nitrogen N	8 Oxygen O	9 Fluorine F	10 Neon Ne
3	11 Sodium Na	12 Magnesium Mg	13 Aluminium Al	14 Silicon Si	15 Phosphorus P	16 Sulphur S	17 Chlorine Cl	18 Argon Ar
4	19 Potassium K	20 Calcium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



- The Atomic Number equals the number of protons = 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 electron configuration of sodium = 2,8,1
- The Mass Number equals number of protons (11) + number of neutrons (12) = 23

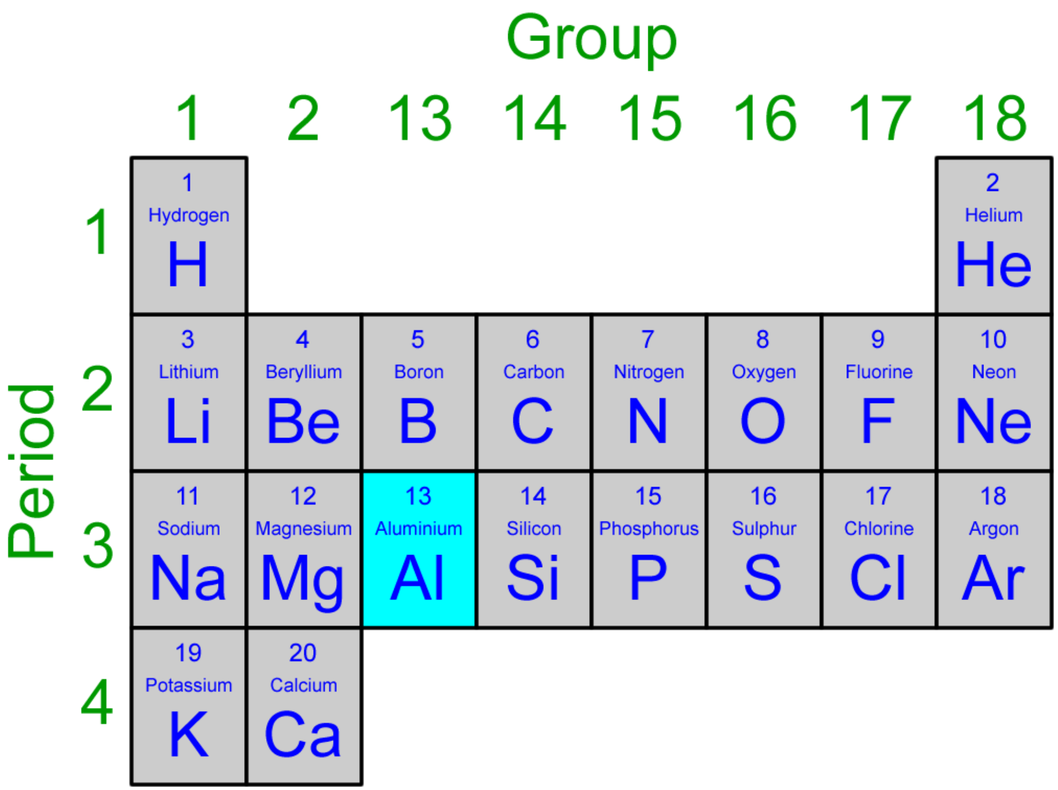
Magnesium $^{24}_{12}Mg$ Atomic Number 12

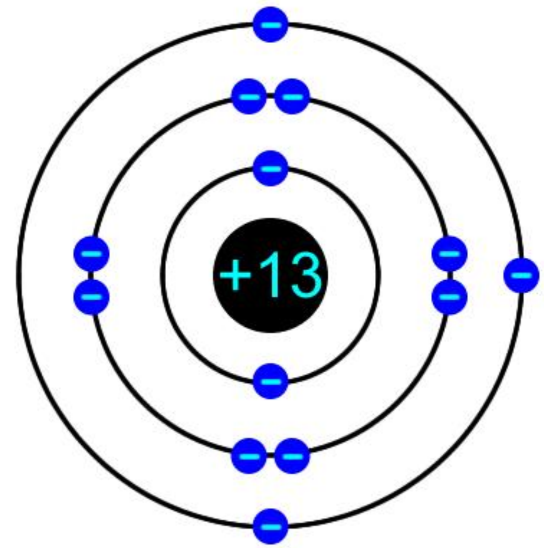
Period \ Group	1	2	13	14	15	16	17	18
1	1 Hydrogen H							2 Helium He
2	3 Lithium Li	4 Beryllium Be	5 Boron B	6 Carbon C	7 Nitrogen N	8 Oxygen O	9 Fluorine F	10 Neon Ne
3	11 Sodium Na	12 Magnesium Mg	13 Aluminium Al	14 Silicon Si	15 Phosphorus P	16 Sulphur S	17 Chlorine Cl	18 Argon Ar
4	19 Potassium 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 electron configuration of magnesium = 2,8,2
- The Mass Number equals number of protons (12) + number of neutrons (12) = 24

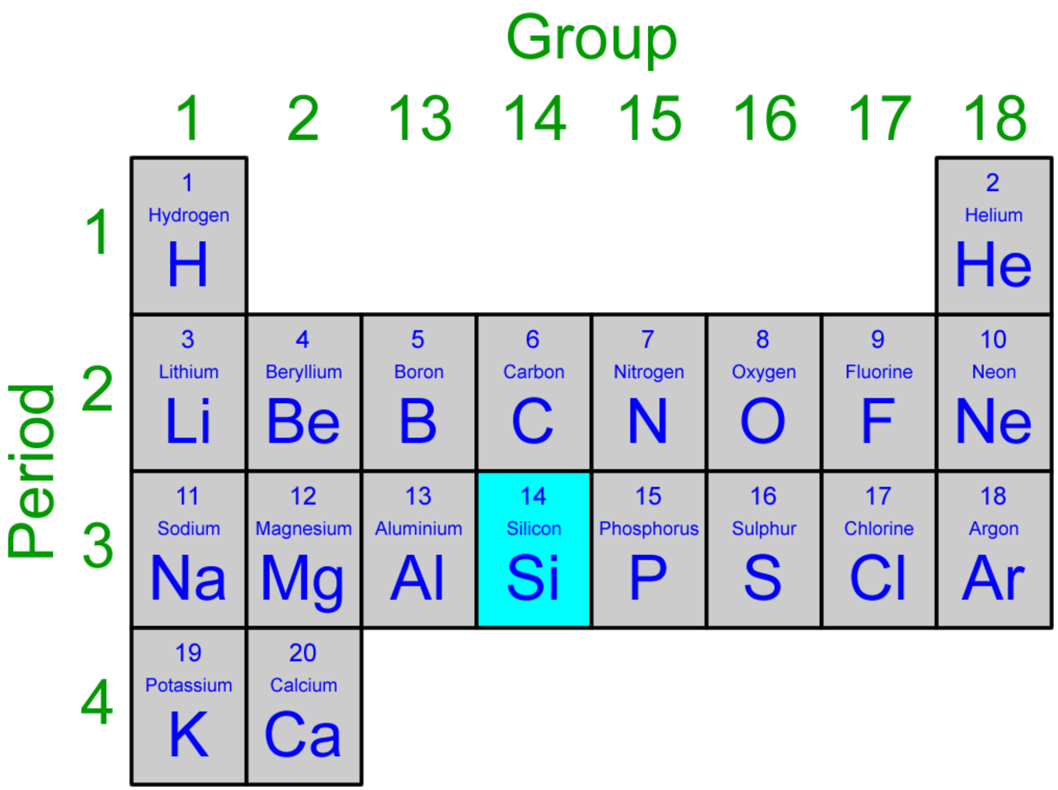
Aluminium $^{27}_{13}Al$ Atomic Number 13

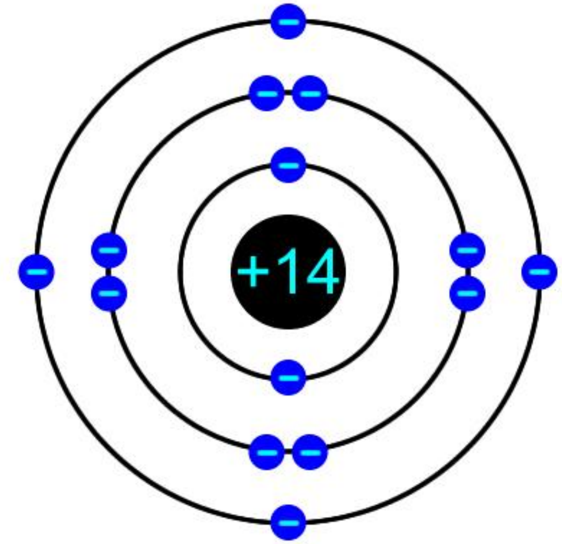




- The Atomic Number equals the number of protons = 13
- For a neutral atom, the number electrons equals the number of protons = 13
- The number of electron shell equals the Period number = 3
- The number of valence electrons can be derived from the Group number (1**3**) = 3
- The electron configuration of aluminium = 2,8,3
- The Mass Number equals number of protons (13) + number of neutrons (14) = 27

Silicon $^{28}_{14}\text{Si}$ Atomic Number 14



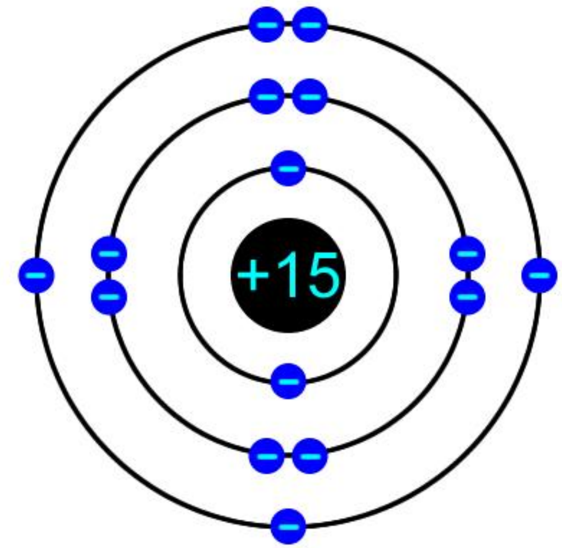


- The Atomic Number equals the number of protons = 14
- For a neutral atom, the number electrons equals the number of protons = 14
- The number of electron shell equals the Period number = 3
- The number of valence electrons can be derived from the Group number (1**4**) = 4
- The electron configuration of silicon = 2,8,4
- The Mass Number equals number of protons (14) + number of neutrons (14) = 28

Phosphorus $^{31}_{15}P$ Atomic Number 15

Group

Period	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						

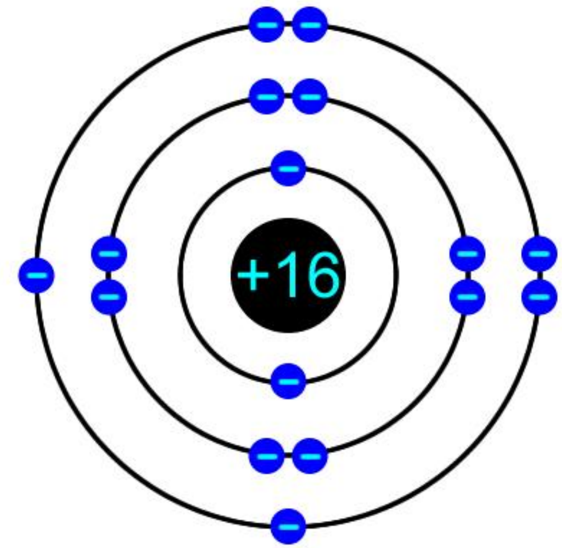


- The Atomic Number equals the number of protons = 15
- For a neutral atom, the number electrons equals the number of protons = 15
- The number of electron shell equals the Period number = 3
- The number of valence electrons can be derived from the Group number (1**5**) = 5
- The electron configuration of phosphorus = 2,8,5
- The Mass Number equals number of protons (15) + number of neutrons (16) = 31

Sulfur $^{32}_{16}S$ Atomic Number 16

Group

Period	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						

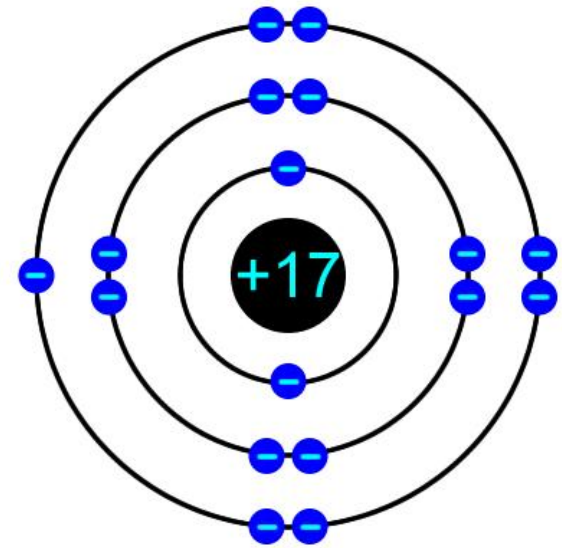


- The Atomic Number equals the number of protons = 16
- For a neutral atom, the number electrons equals the number of protons = 16
- The number of electron shell equals the Period number = 3
- The number of valence electrons can be derived from the Group number (1**6**) = 6
- The electron configuration of sulfur = 2,8,6
- The Mass Number equals number of protons (16) + number of neutrons (16) = 32

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

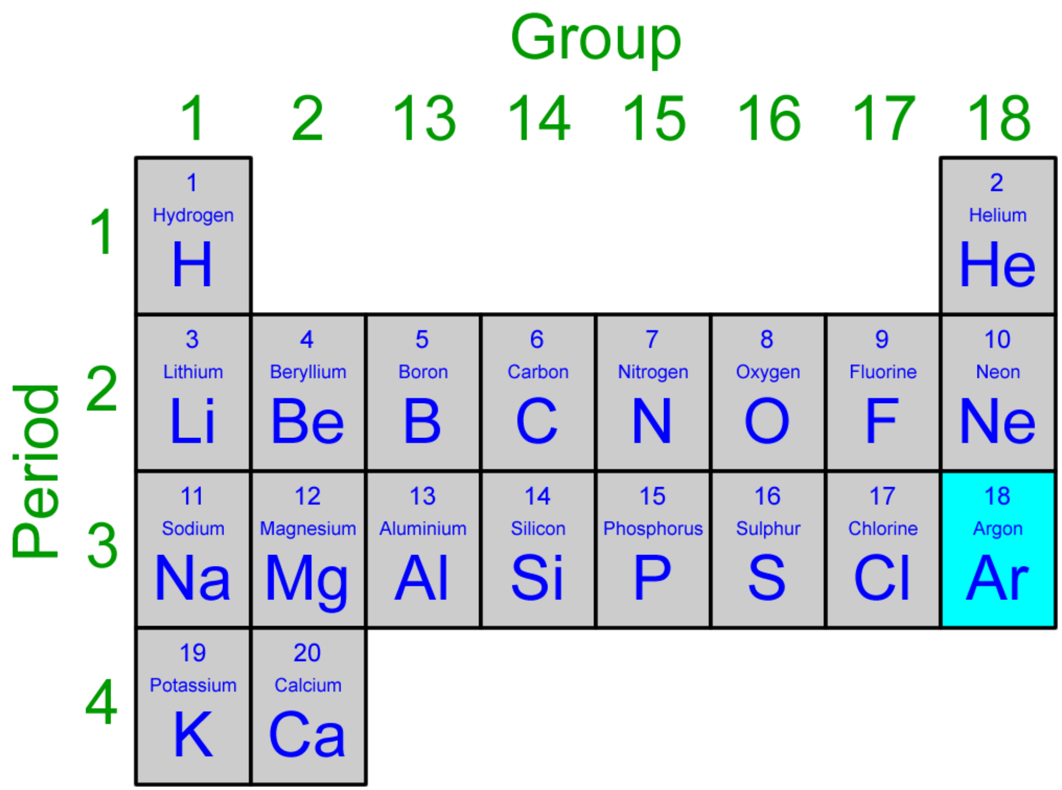
Group

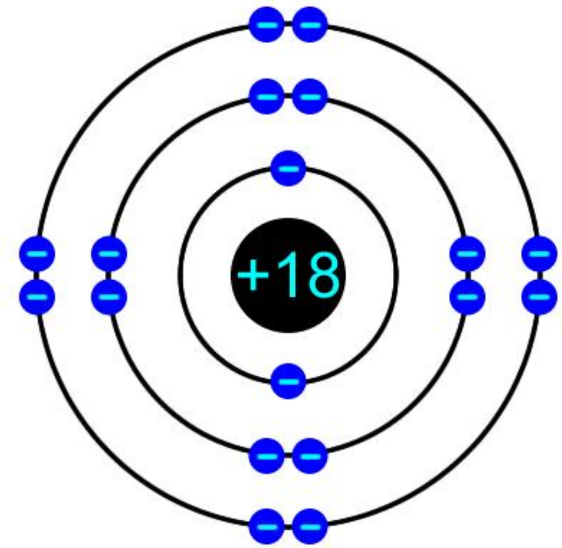
Period	1	2	13	14	15	16	17	18
1	1 Hydrogen H							2 Helium He
2	3 Lithium Li	4 Beryllium Be	5 Boron B	6 Carbon C	7 Nitrogen N	8 Oxygen O	9 Fluorine F	10 Neon Ne
3	11 Sodium Na	12 Magnesium Mg	13 Aluminium Al	14 Silicon Si	15 Phosphorus P	16 Sulphur S	17 Chlorine Cl	18 Argon Ar
4	19 Potassium K	20 Calcium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



- The Atomic Number equals the number of protons = 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 (1**7**) = 7
- The electron configuration of chlorine = 2,8,7
- The Mass Number equals number of protons (17) + number of neutrons (18) = 35

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



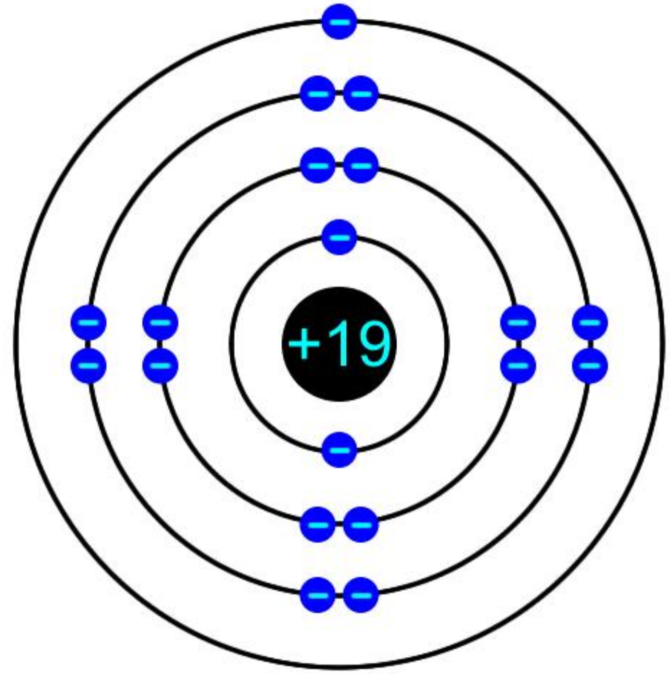


- The Atomic Number equals the number of protons = 18
- For a neutral atom, the number electrons equals the number of protons = 18
- The number of electron shell equals the Period number = 3
- The number of valence electrons can be derived from the Group number (1**8**) = 8 (complete)
- The electron configuration of argon = 2,8,8
- The Mass Number equals number of protons (18) + number of neutrons (22) = 40

Potassium $^{39}_{19}K$ Atomic Number 19

Group

Period	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						

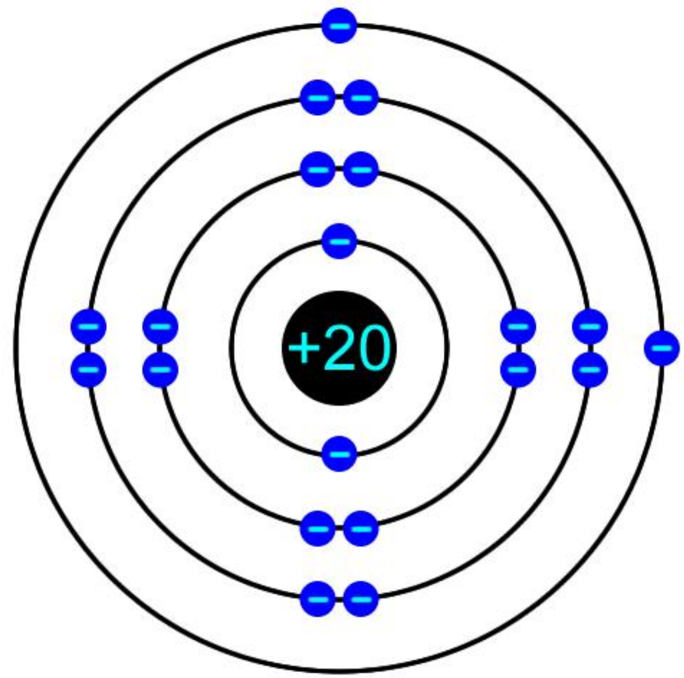


- The Atomic Number equals the number of protons = 19
- For a neutral atom, the number electrons equals the number of protons = 19
- The number of electron shell equals the Period number = 4
- The number of valence electrons can be derived from the Group number (**1**) = 1
- The electron configuration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

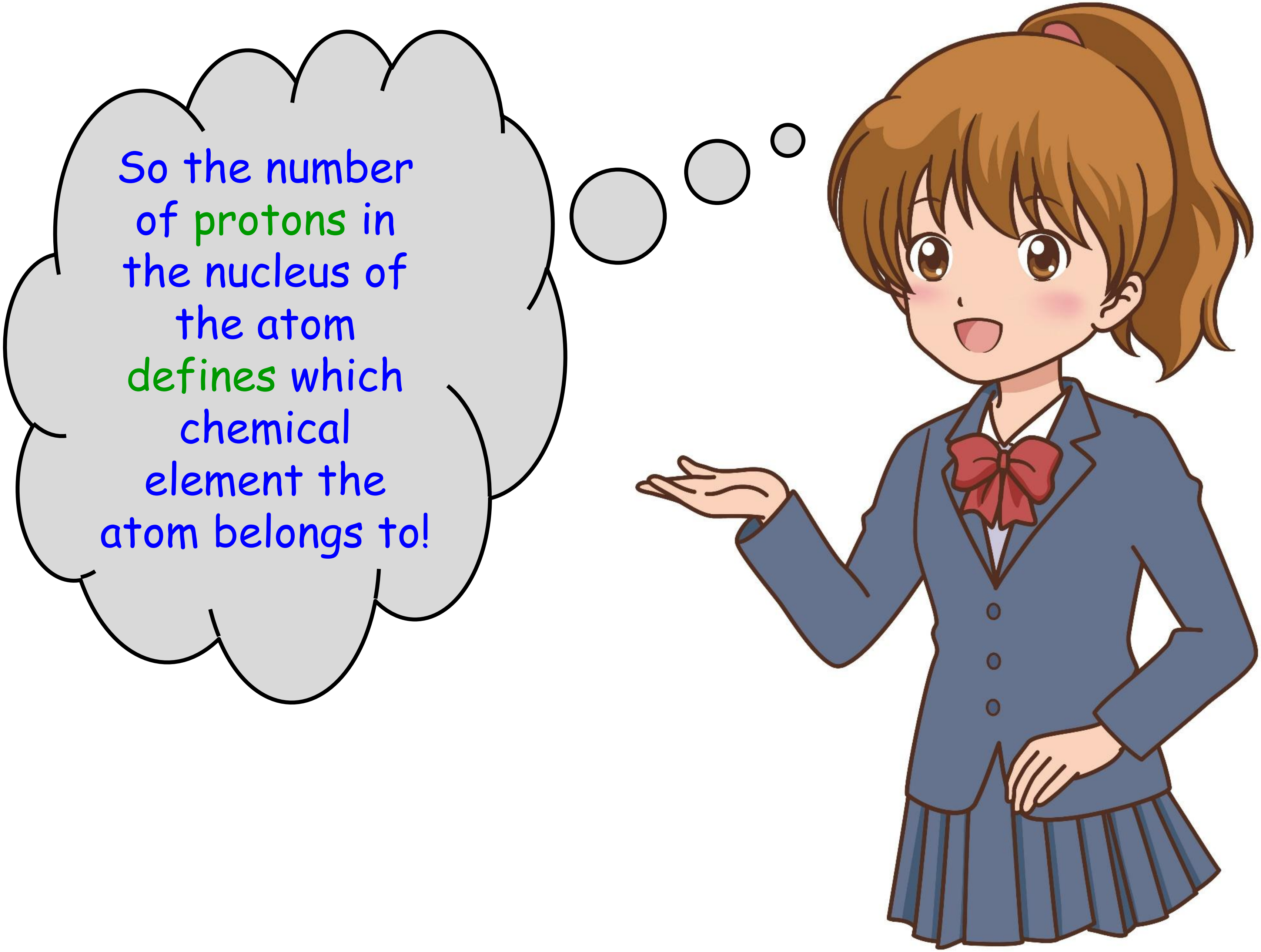
Calcium $^{40}_{20}Ca$ Atomic Number 20

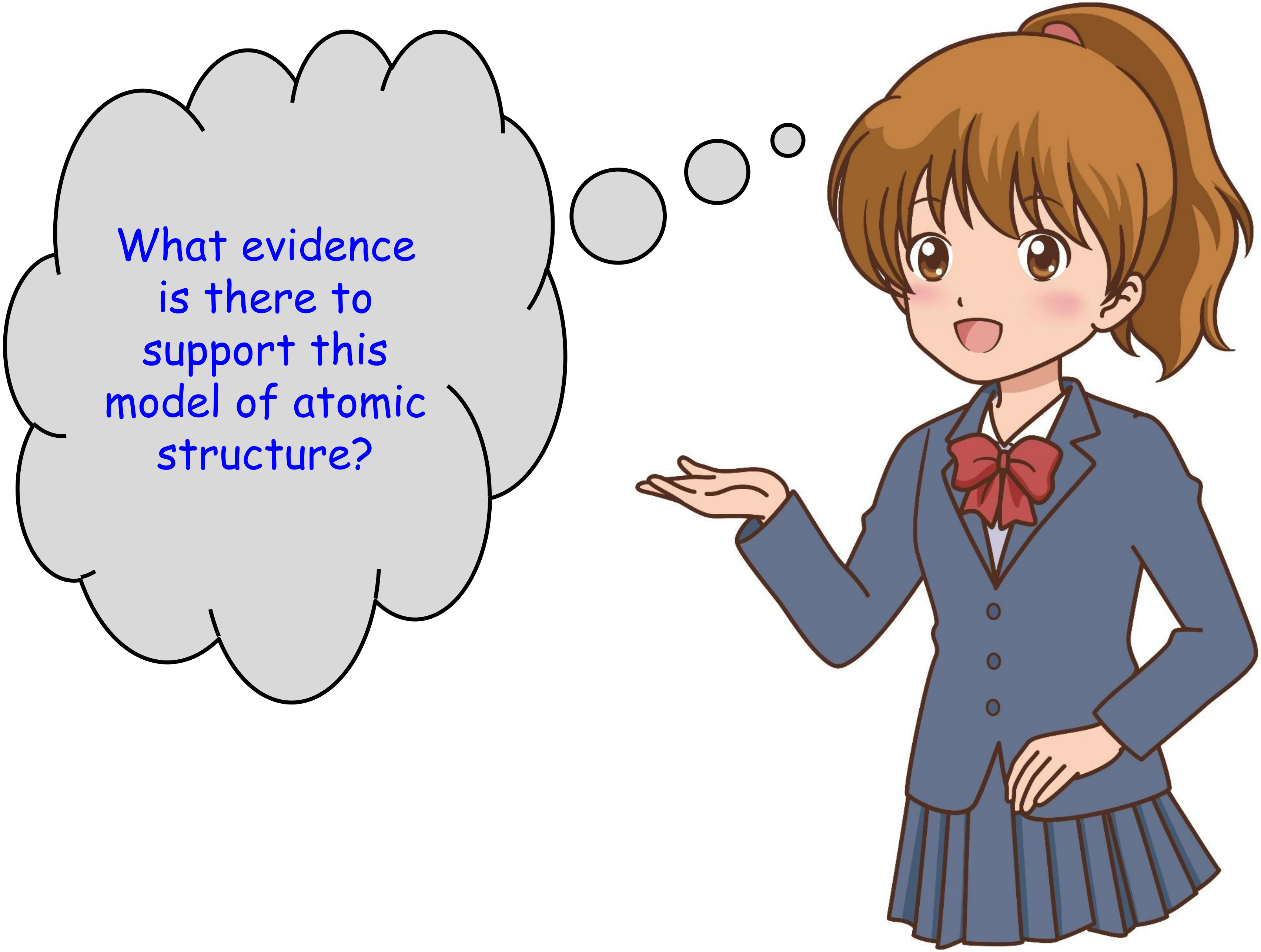
Group

Period	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						



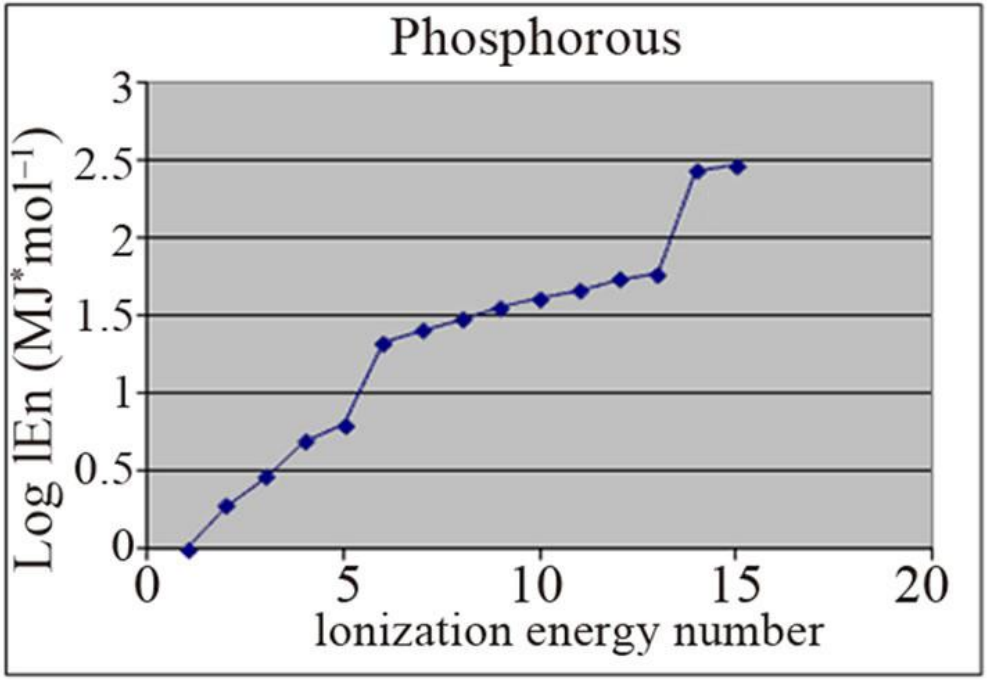
- The Atomic Number equals the number of protons = 20
- For a neutral atom, the number electrons equals the number of protons = 20
- The number of electron shell equals the Period number = 4
- The number of valence electrons can be derived from the Group number (**2**) = 2
- The electron configuration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

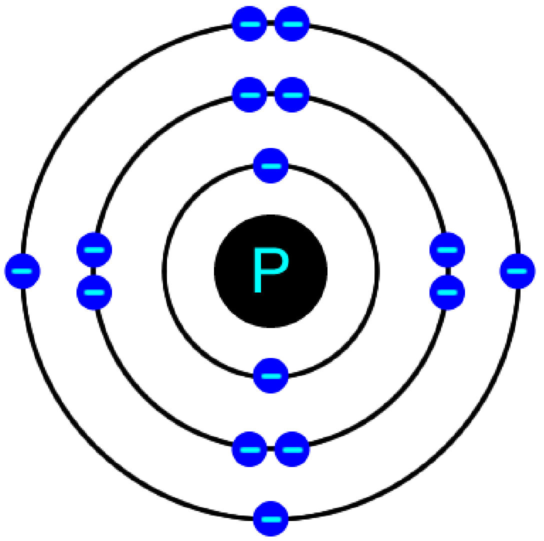

So the number of protons in the nucleus of the atom defines which chemical element the atom belongs to!


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

Introduction to Atomic Structure

Evidence of Atomic Structure – Successive Ionisation Energies

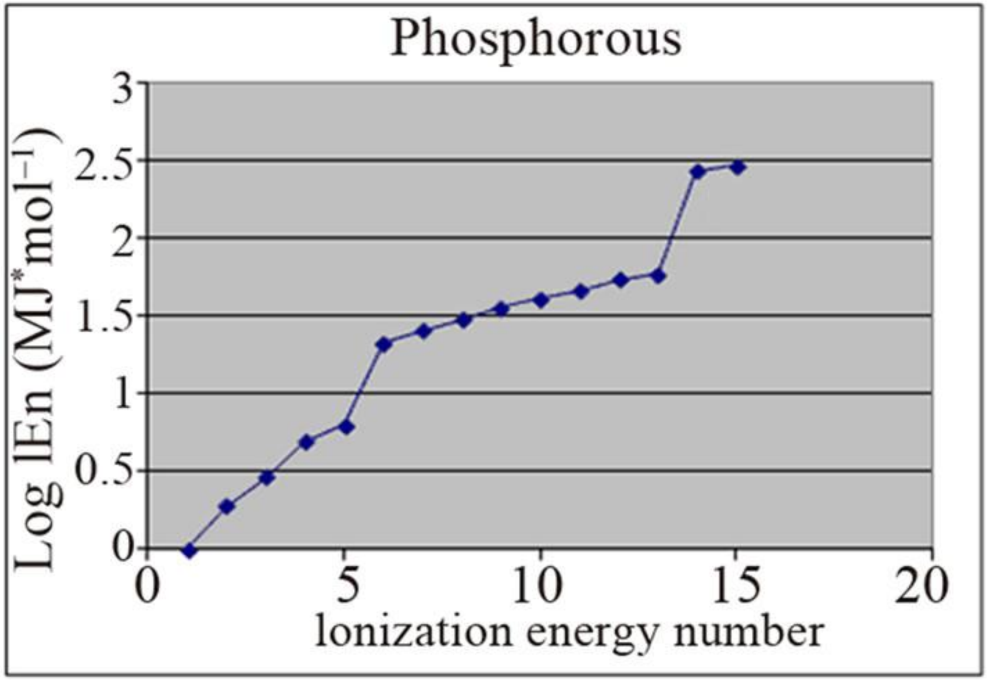


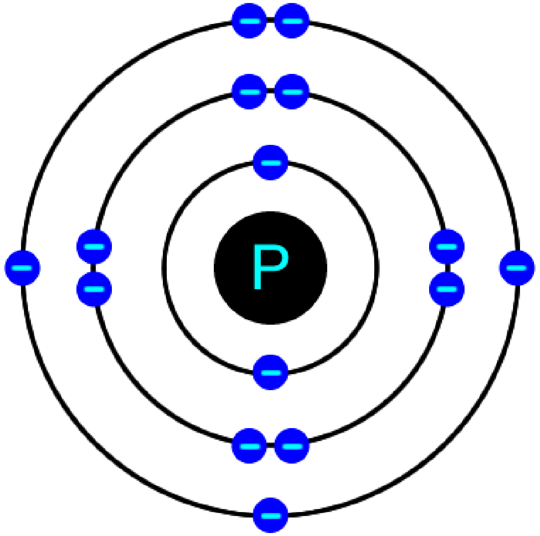


- Energy is *required* to *remove* an electron from an atom. The energy is used to overcome the *electrostatic force of attraction* between the negatively charged electron and positively charged nucleus.

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.

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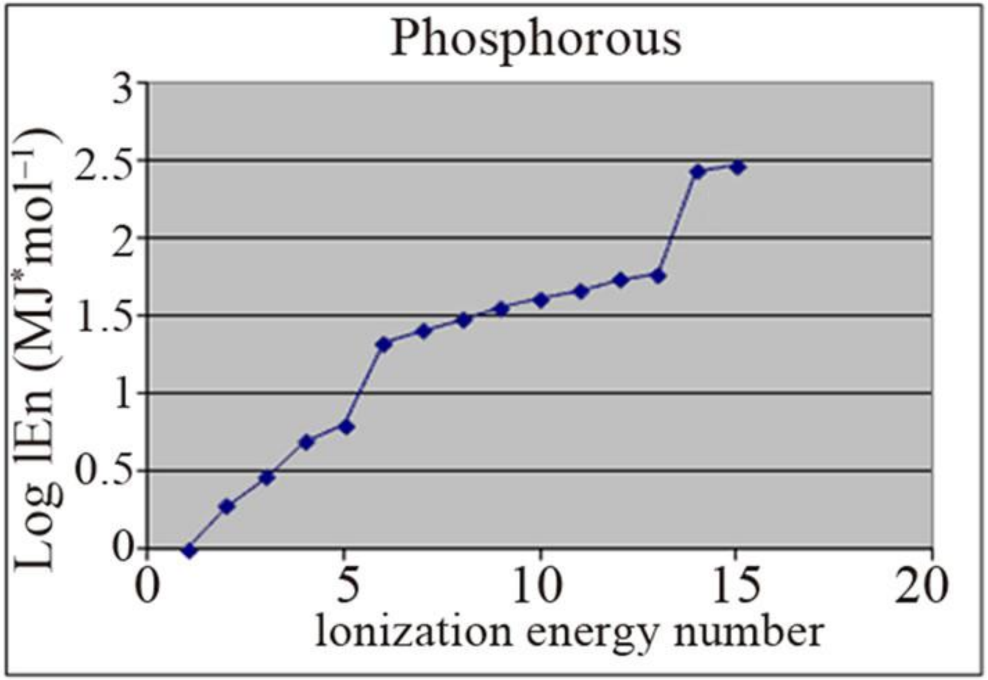
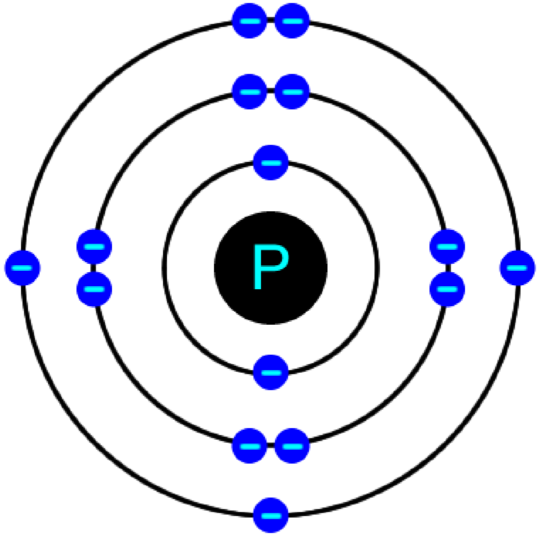
Evidence of Atomic Structure – Successive Ionisation Energies





1st Ionisation: $P(g) \rightarrow P^{+}(g) + e^{-}$

2nd Ionisation: $P^{+}(g) \rightarrow P^{2+}(g) + e^{-}$

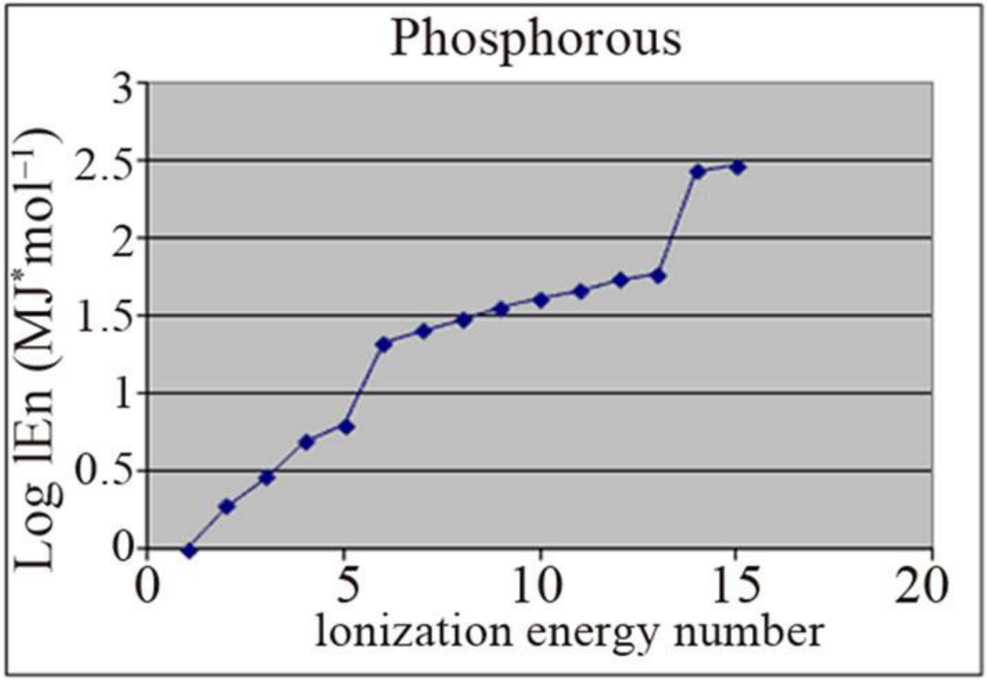
3rd Ionisation: $P^{2+}(g) \rightarrow P^{3+}(g) + e^{-}$

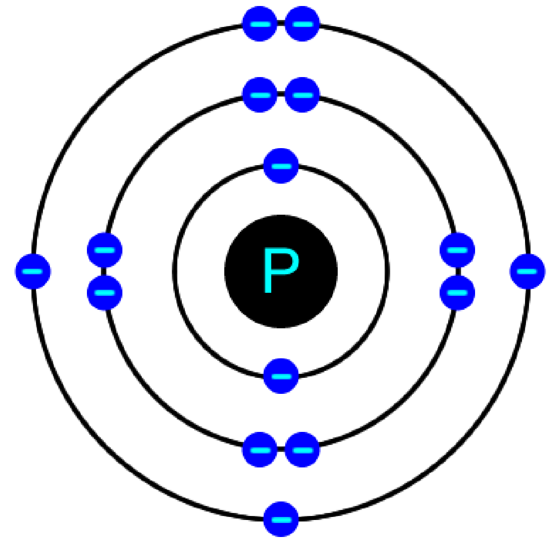
4th Ionisation: $P^{3+}(g) \rightarrow P^{4+}(g) + e^{-}$

…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:

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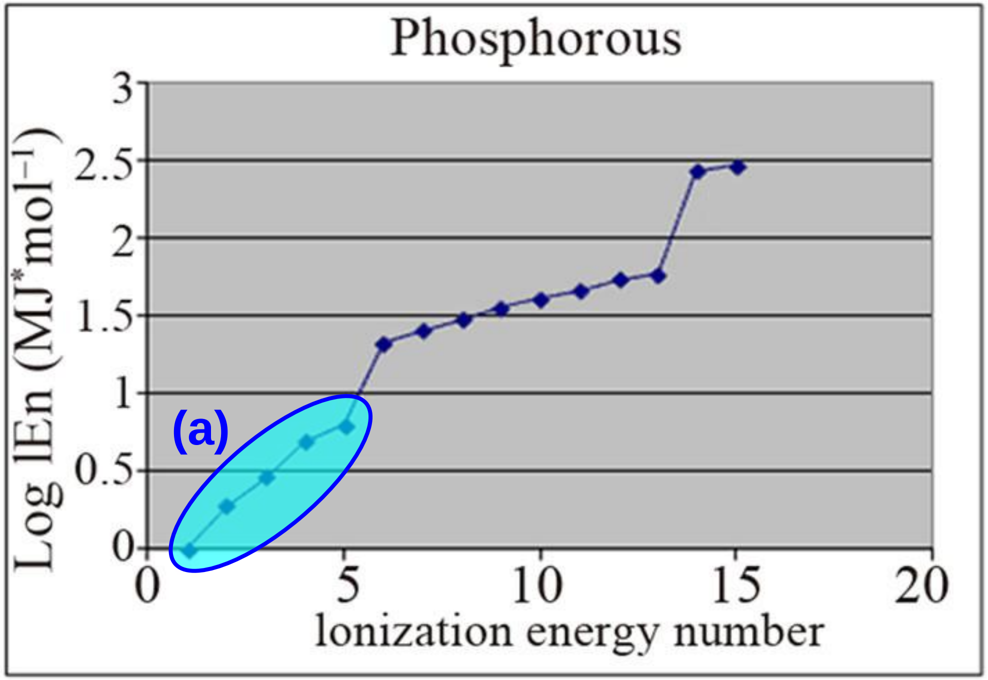
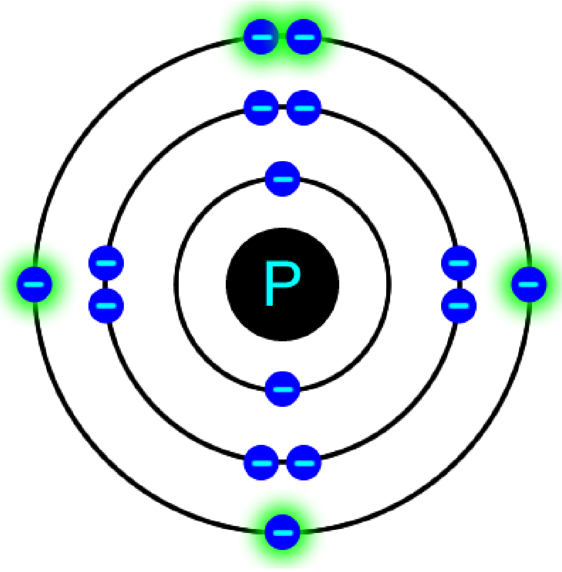
Evidence of Atomic Structure – Successive Ionisation Energies





- 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 shell*.

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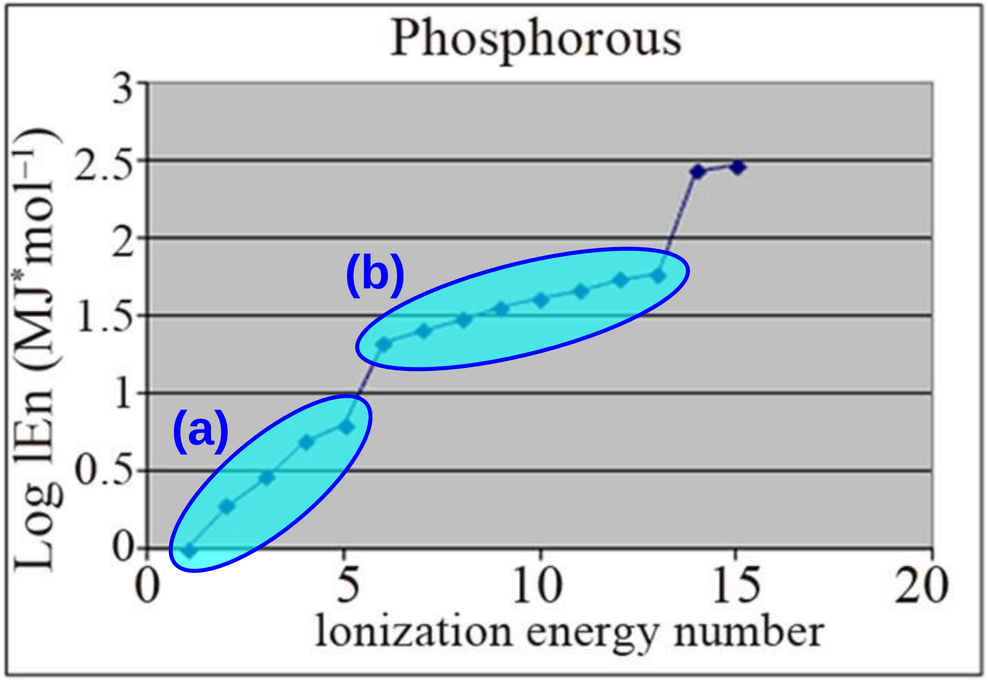
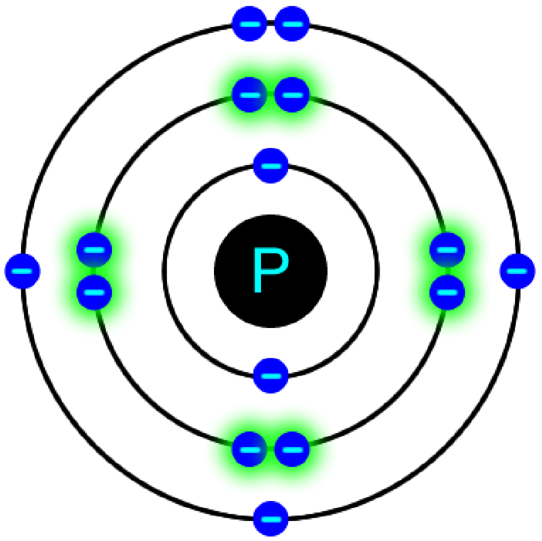
Evidence of Atomic Structure – Successive Ionisation Energies





- 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 shell*.

(b) Removal of the *eight electrons* in the *second shell*.

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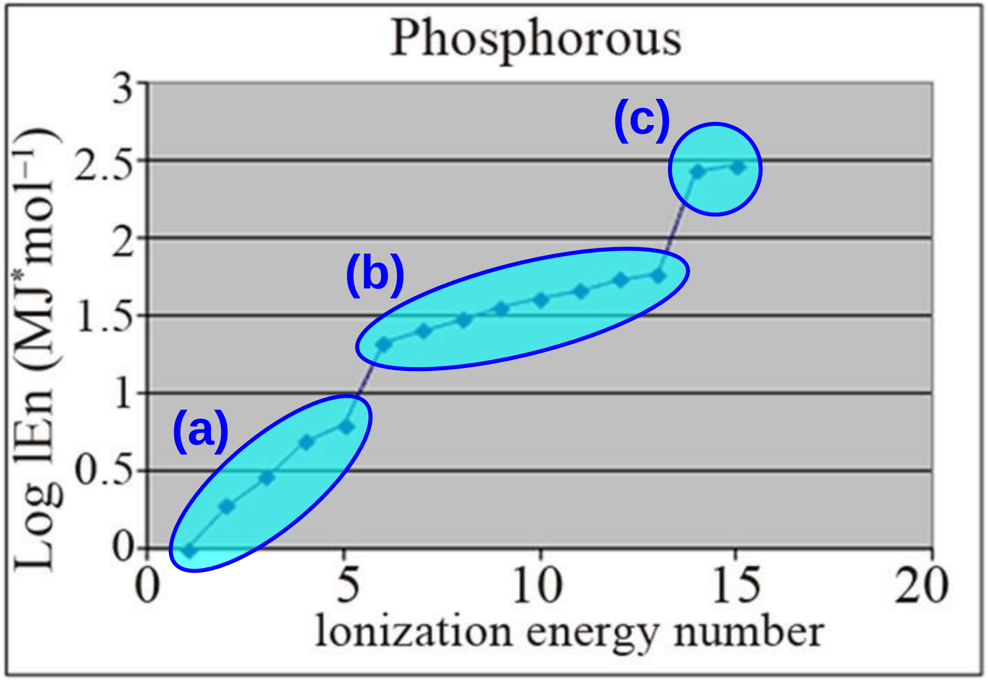
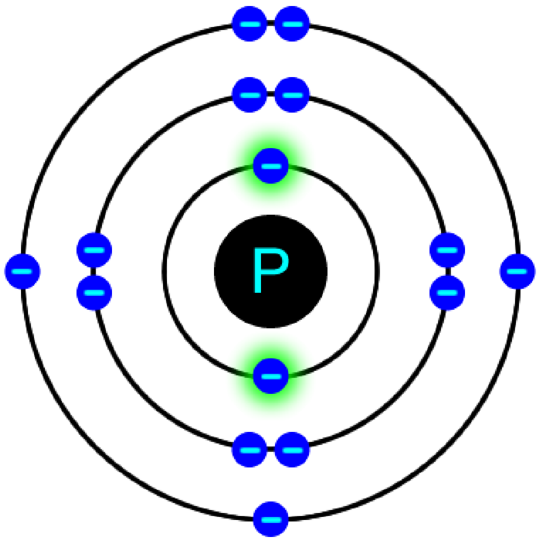
Evidence of Atomic Structure – Successive Ionisation Energies





- 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 shell*.

(b) Removal of the *eight electrons* in the *second shell*.

(c) Removal of the *two electrons* in the *inner shell*.

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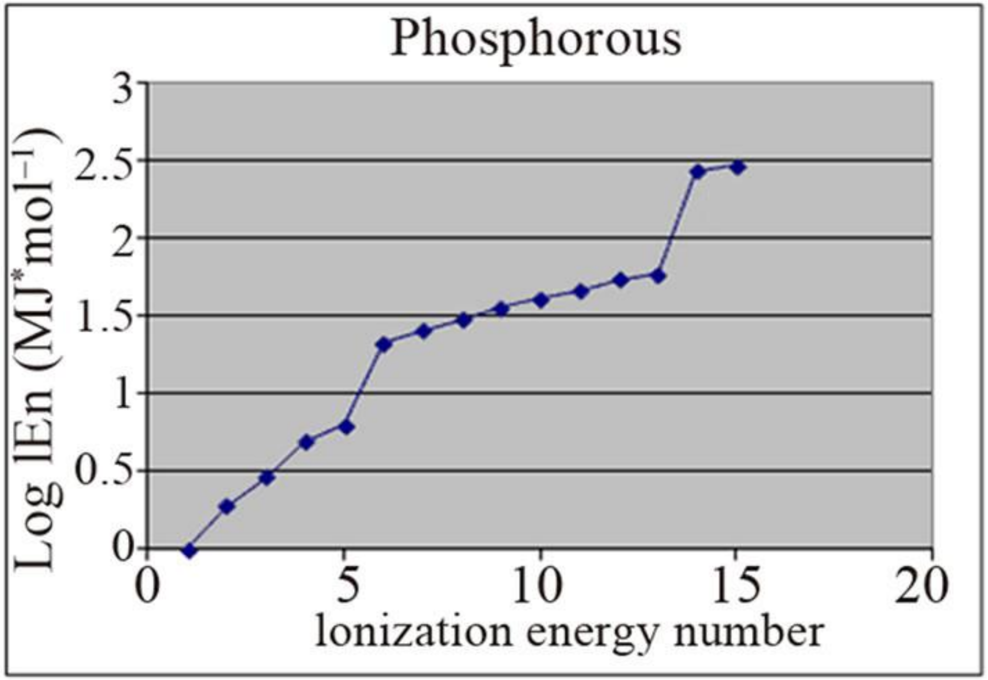
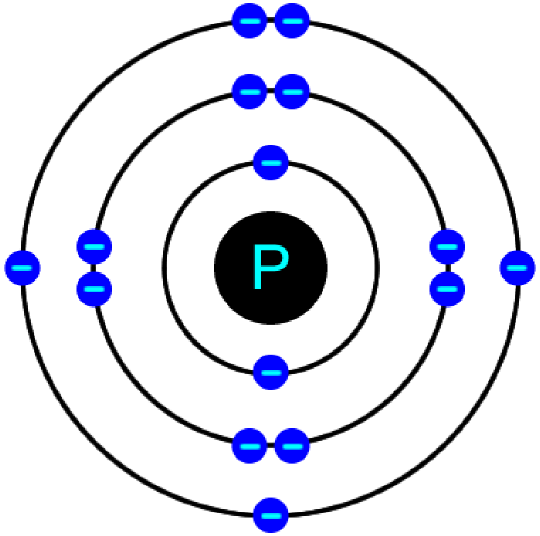
Evidence of Atomic Structure – Successive Ionisation Energies





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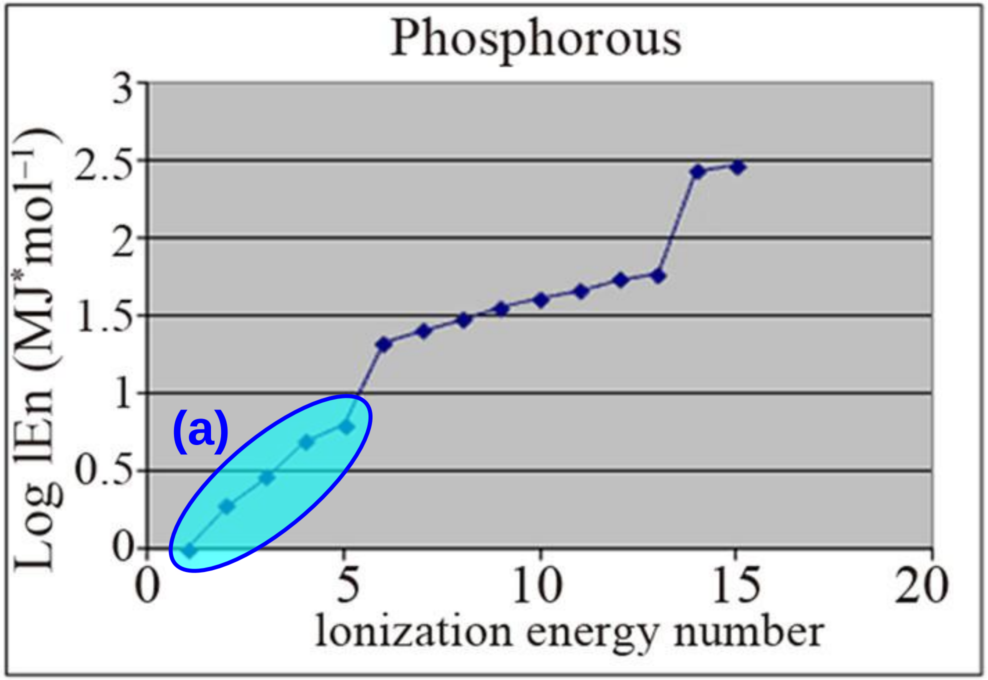
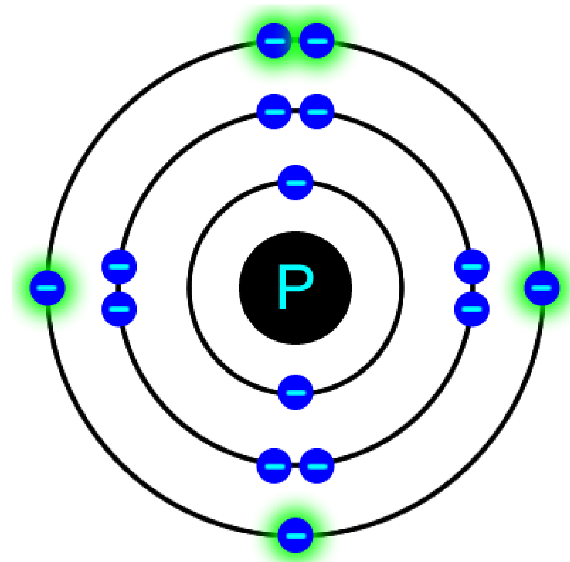
Evidence of Atomic Structure – Successive Ionisation Energies





(a) The *five electrons* in the *valence 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.

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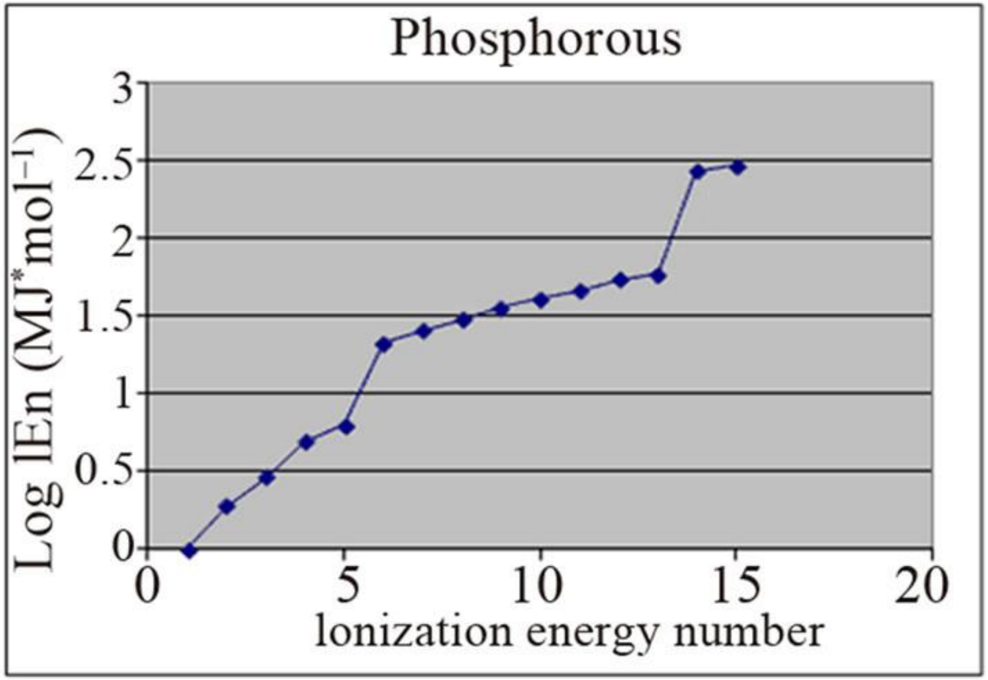
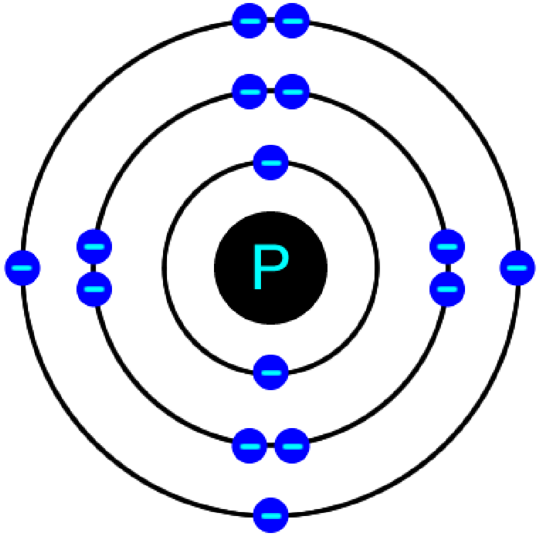
Evidence of Atomic Structure – Successive Ionisation Energies





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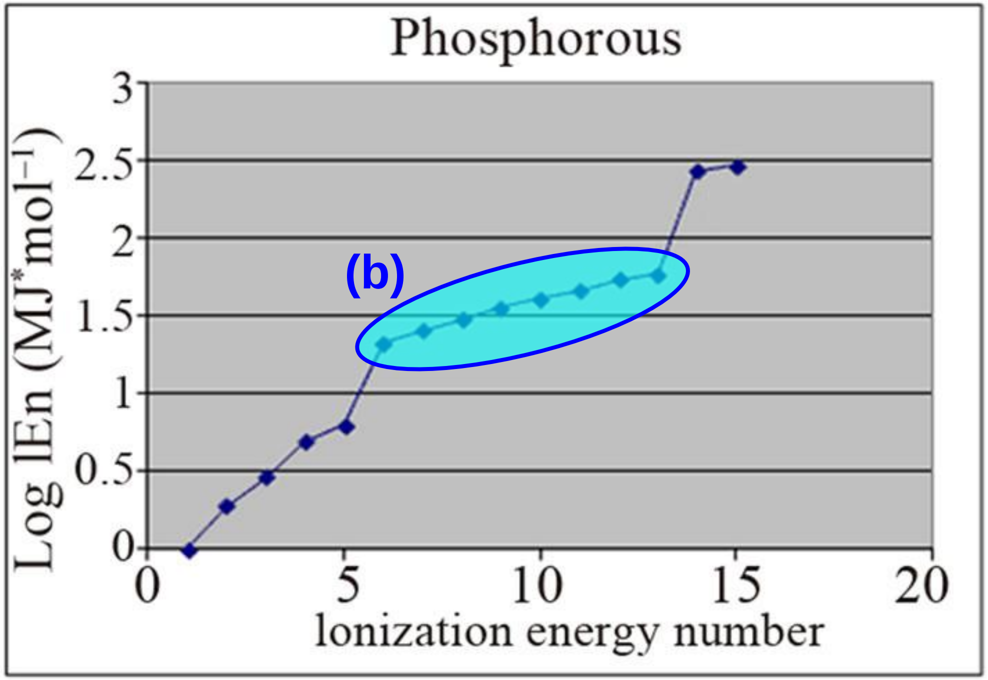
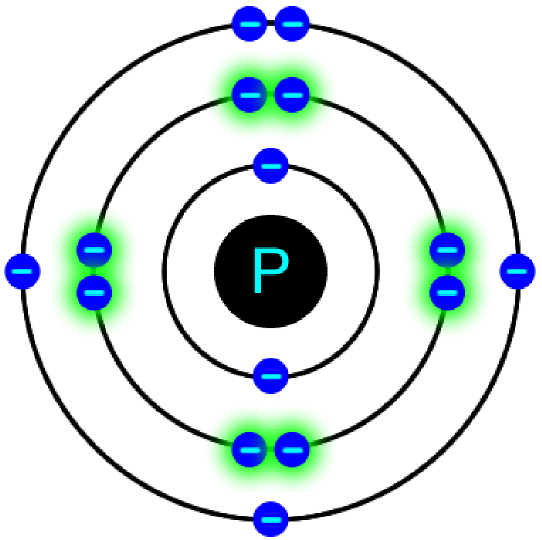
Evidence of Atomic Structure – Successive Ionisation Energies





(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.

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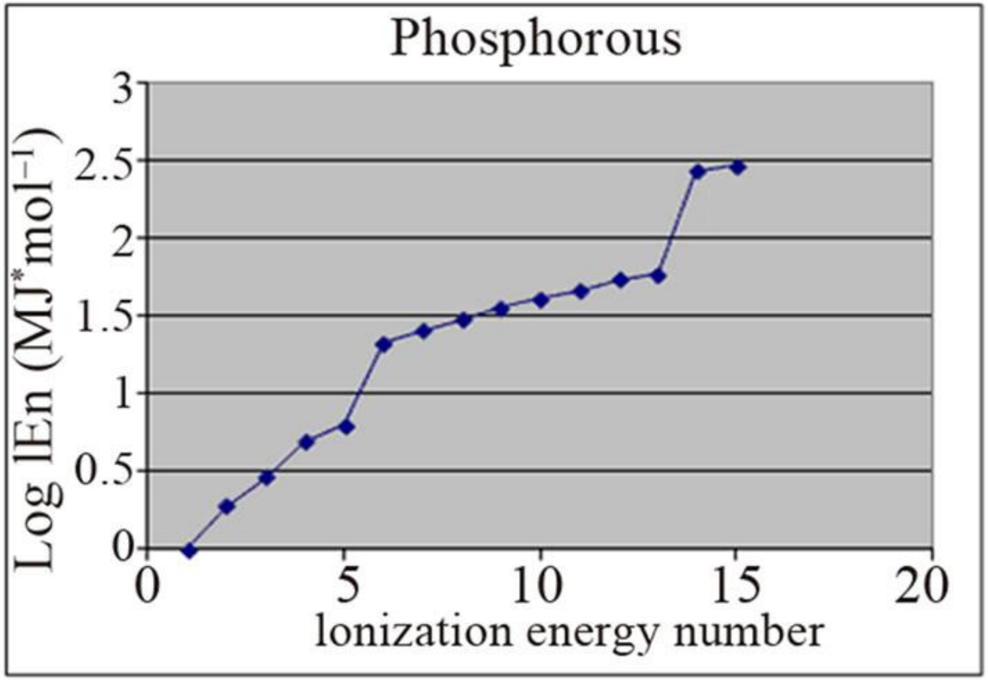
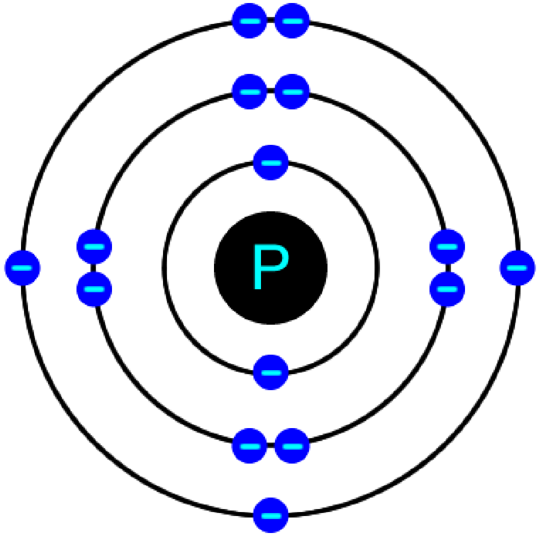
Evidence of Atomic Structure – Successive Ionisation Energies





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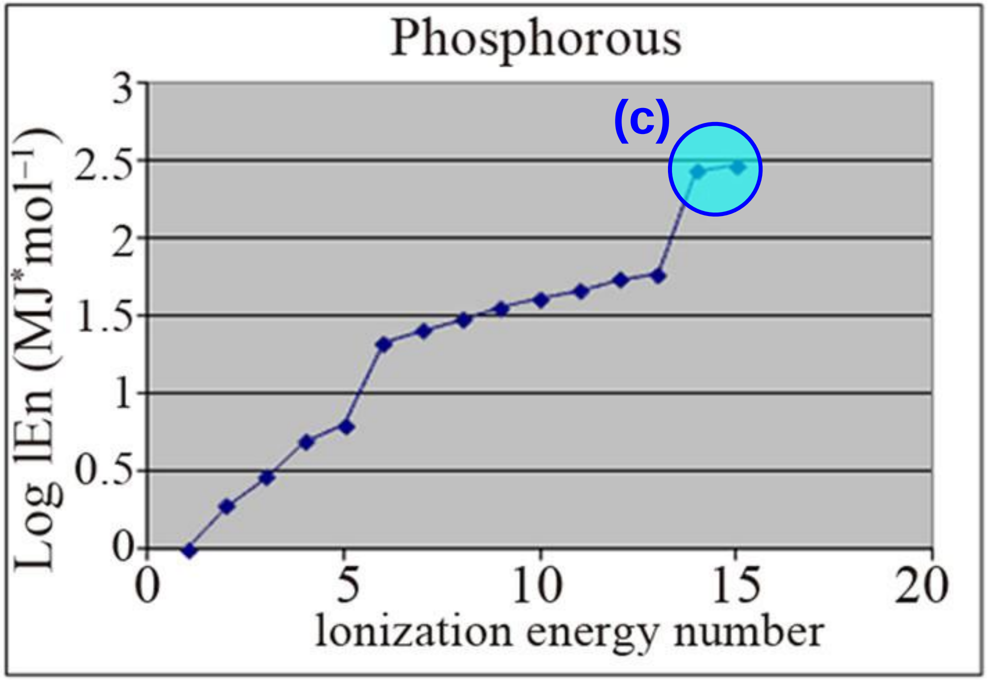
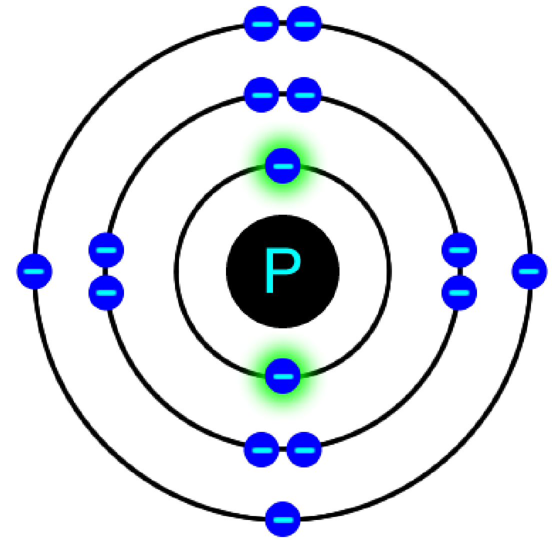
Evidence of Atomic Structure – Successive Ionisation Energies





(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.

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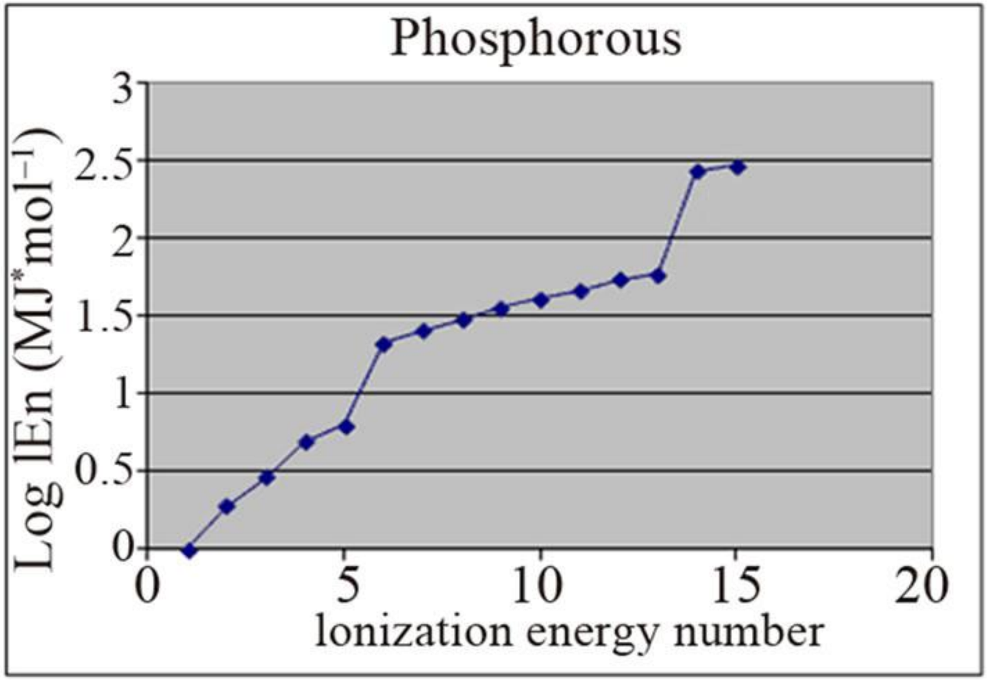
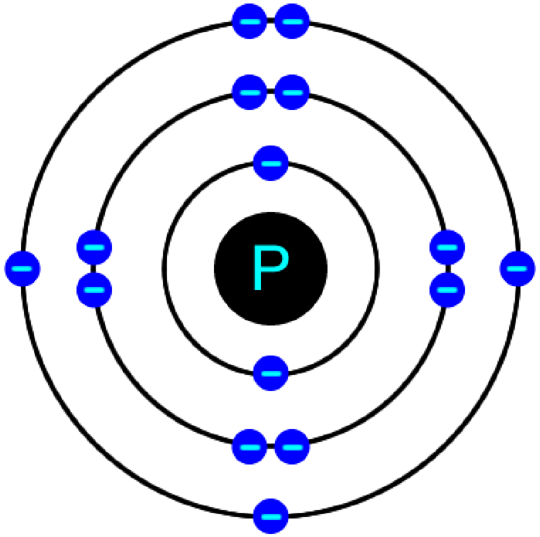
Evidence of Atomic Structure – Successive Ionisation Energies





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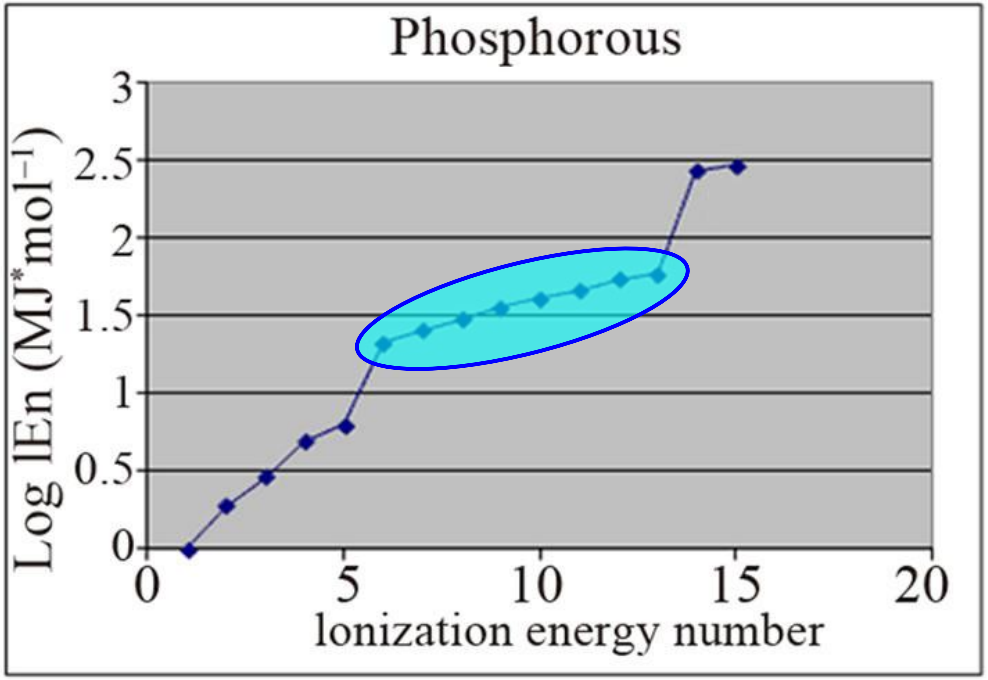
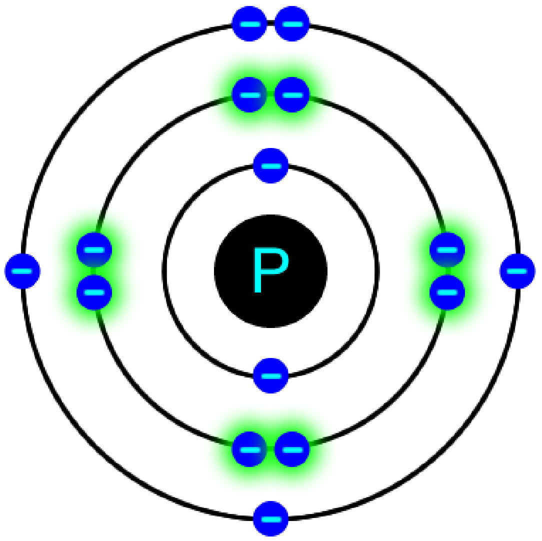
Evidence of Atomic Structure – Successive Ionisation Energies

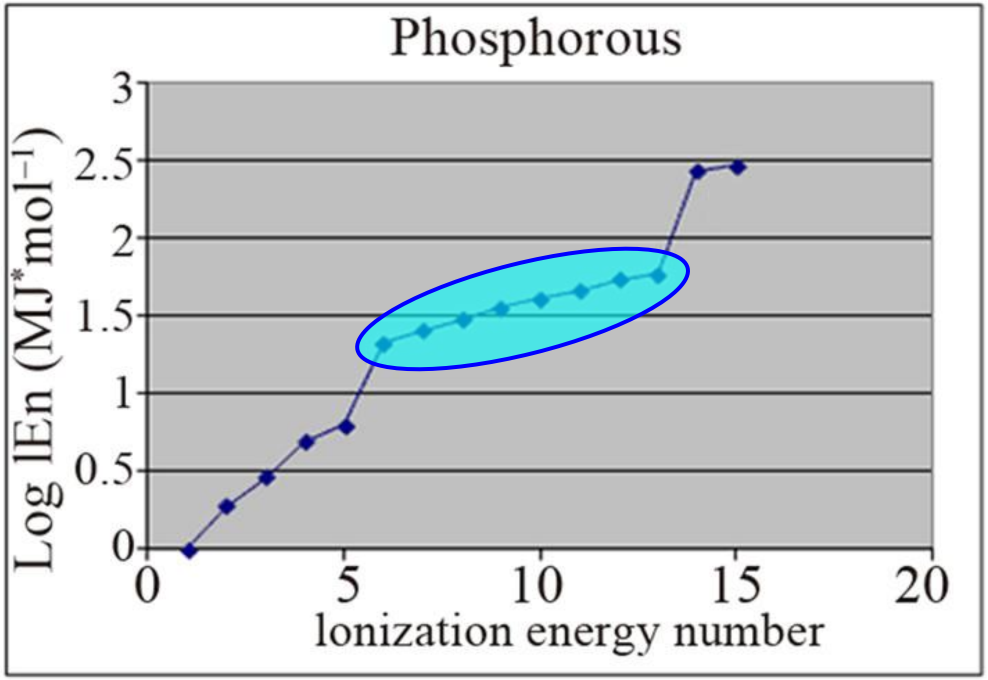


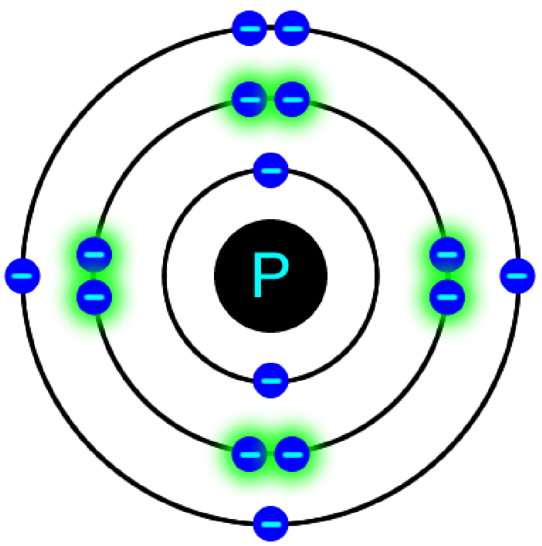


Note: The energy required to remove electrons from the *same* electron shell *gradually increases*. This is because the phosphorus ion is steadily becoming *more positively charged*, hence *more energy is required* to overcome the *stronger electrostatic force of attraction* between the negatively charged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

7th Ionisation: $P^{6+}(g) \rightarrow P^{7+}(g) + e^-$

8th Ionisation: $P^{7+}(g) \rightarrow P^{8+}(g) + e^-$

9th Ionisation: $P^{8+}(g) \rightarrow P^{9+}(g) + e^-$

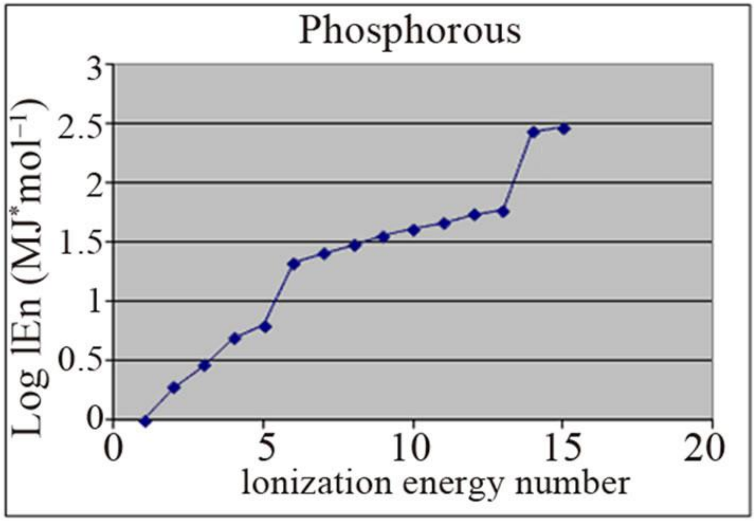
10th Ionisation: $P^{9+}(g) \rightarrow P^{10+}(g) + e^-$

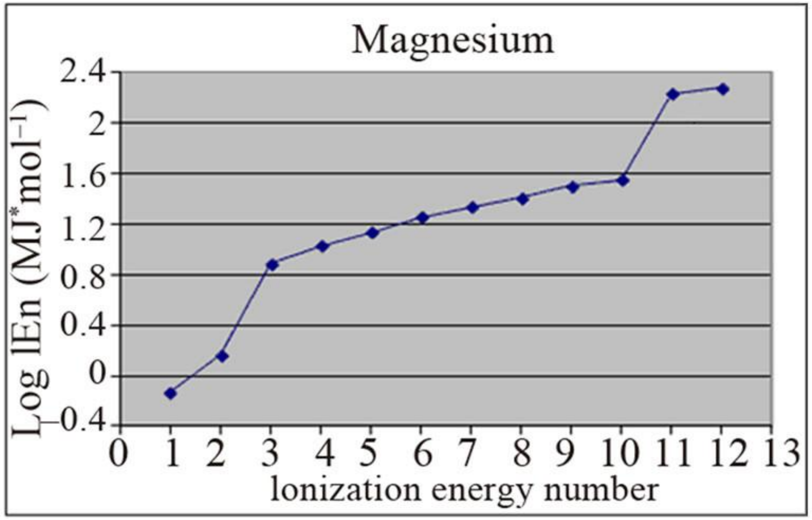
…and so on…

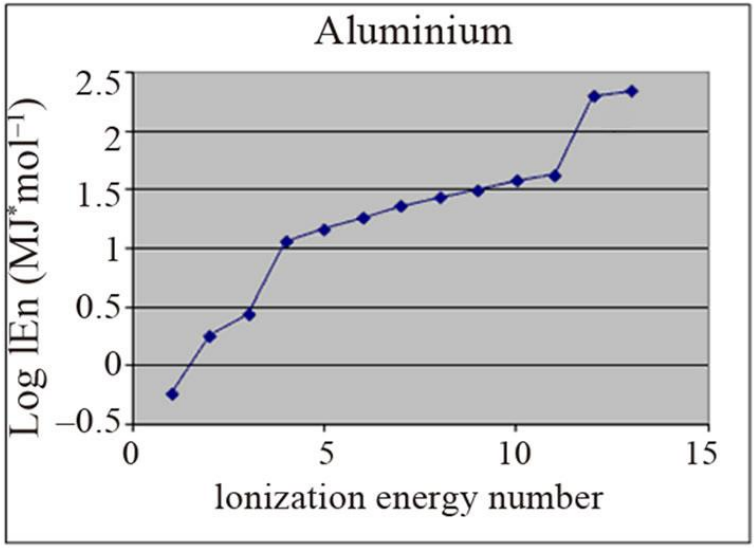
Increasing electrostatic force of attraction between e^- and ion.

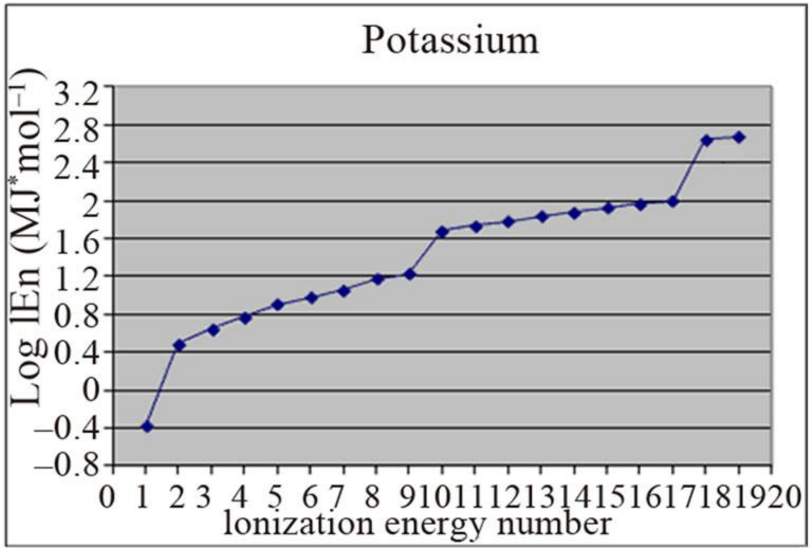
Introduction to Atomic Structure

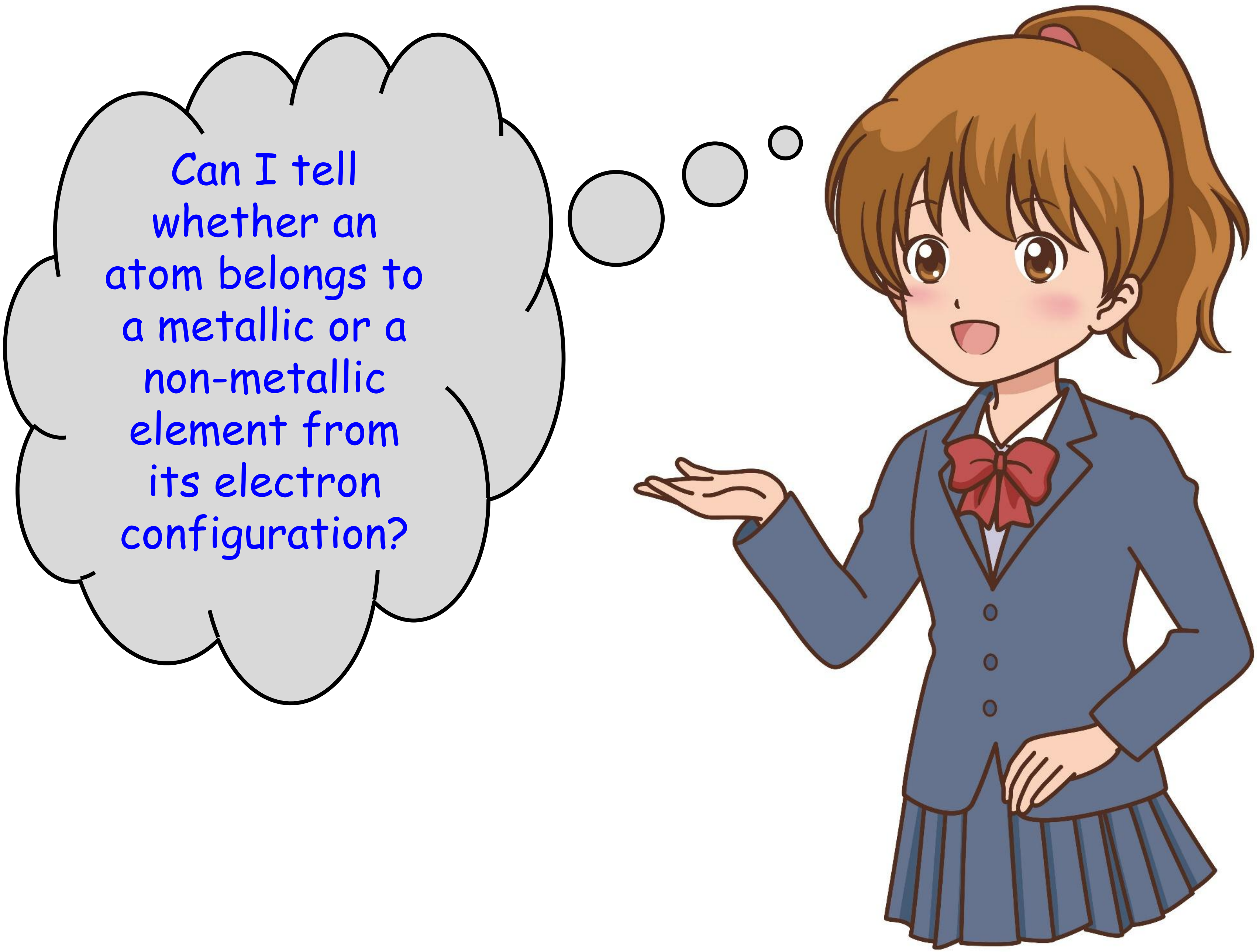
Evidence of Atomic Structure – Successive Ionisation Energies







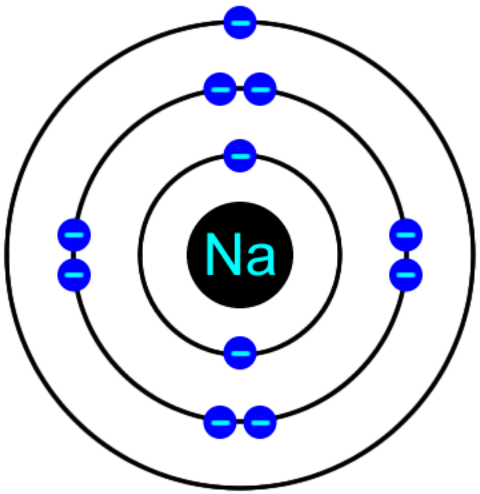


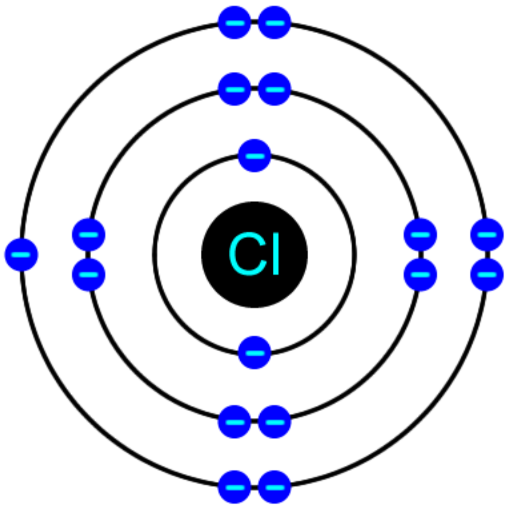

Can I tell whether an atom belongs to a metallic or a non-metallic element from its electron configuration?

Introduction to Atomic Structure

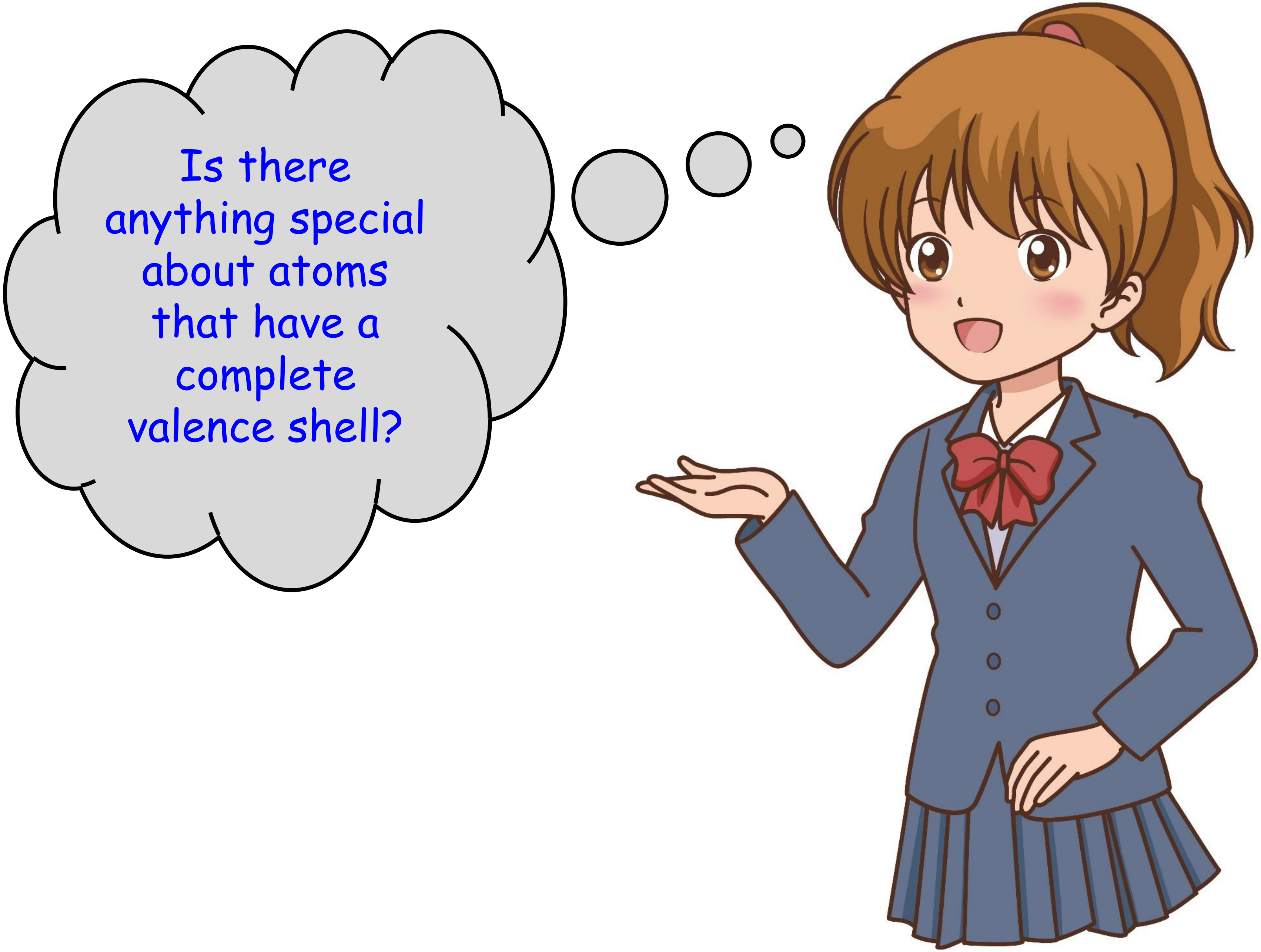
- Atoms of *metallic elements* tend to have only *one*, *two* or *three* electrons in their valence shell.
- Atoms of *non-metallic elements* tend to have *four*, *five, six, seven* or *eight* electrons in their valence shell.
- **Note**: This is only a *general rule* and there are some exceptions.

Introduction to Atomic Structure



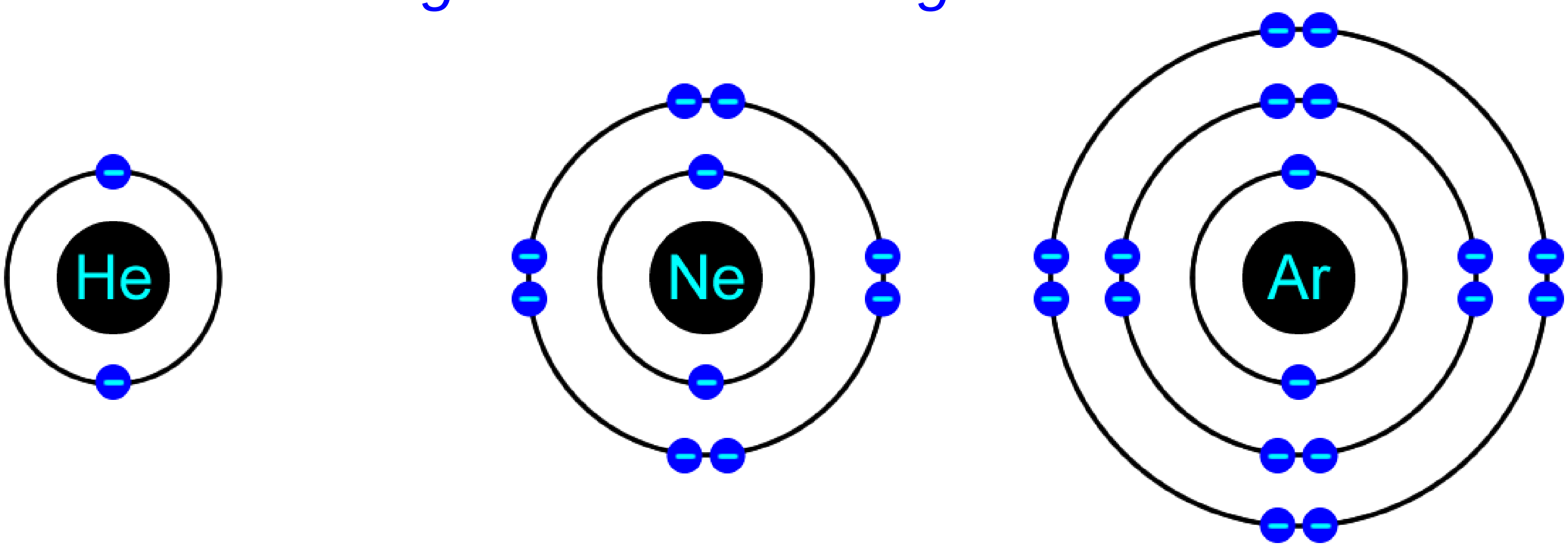


- An atom of *sodium* only has *one* electron in its valence shell, making it a *metal*.
- An atom of *chlorine* has *seven* electrons in its valence shell, making it a *non-metal*.


Is there anything special about atoms that have a complete valence shell?

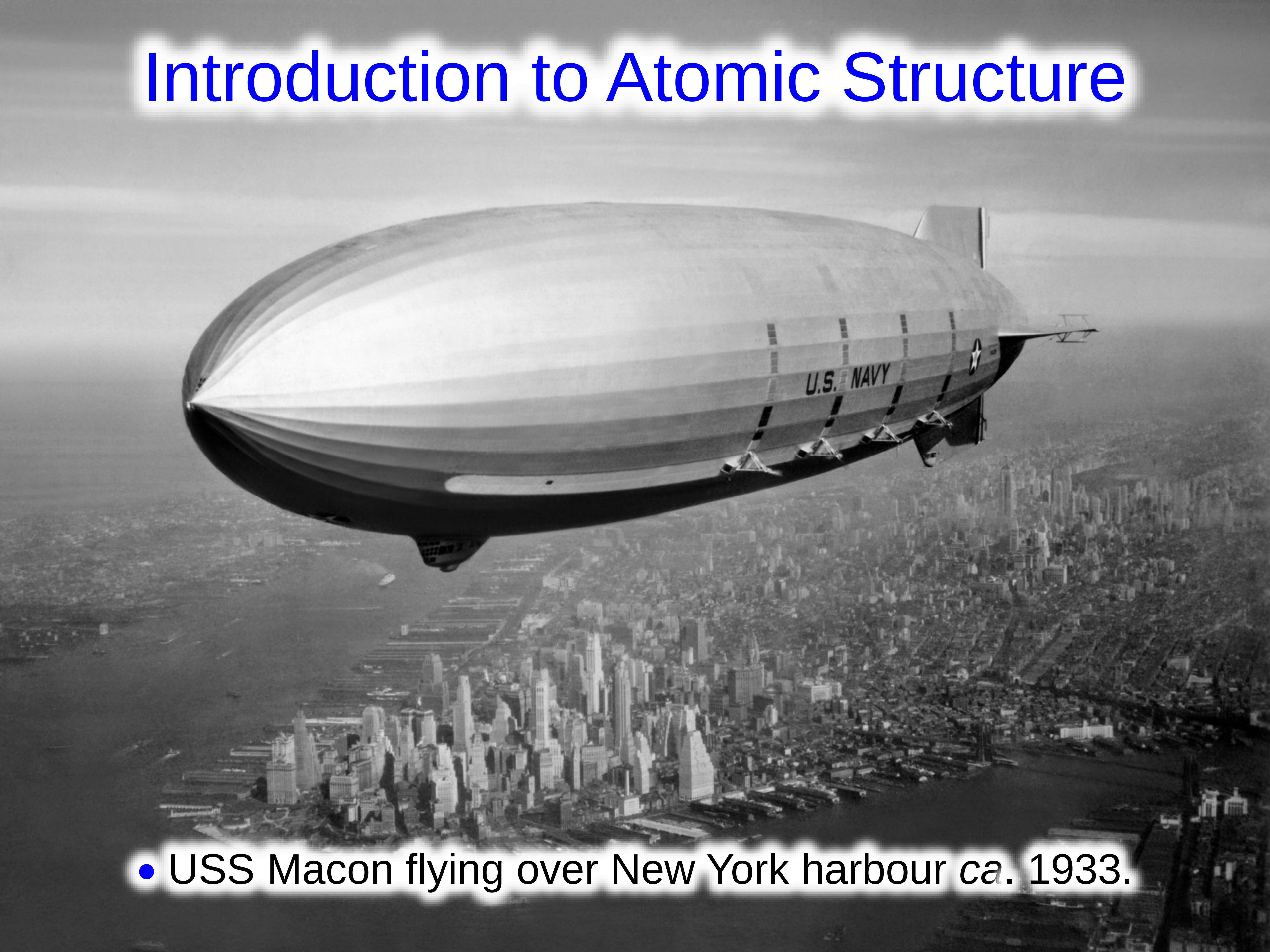
Introduction to Atomic Structure

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



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

Introduction to Atomic Structure

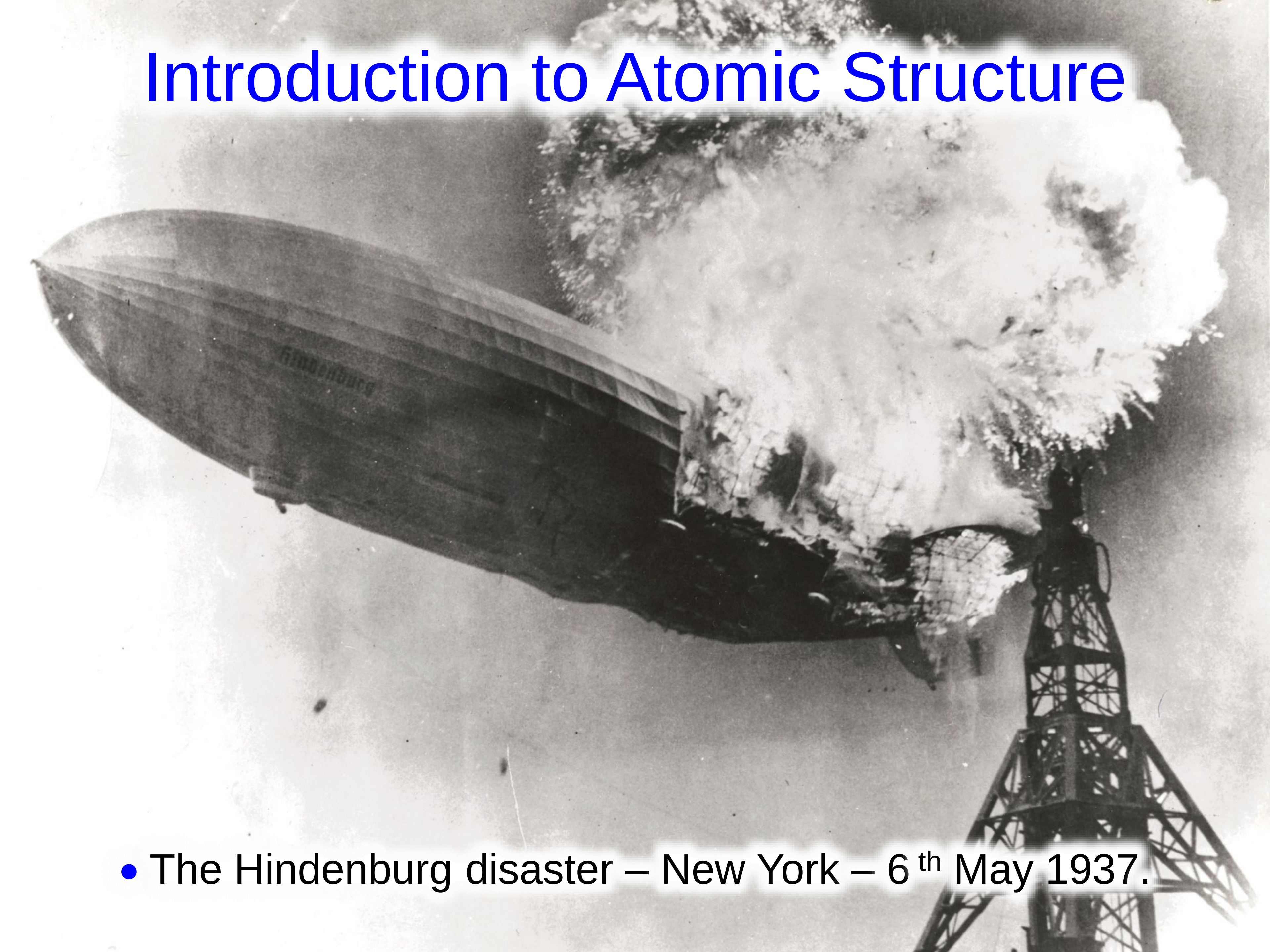


- USS Macon flying over New York harbour *ca*. 1933.

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 – 6 th May 1937.

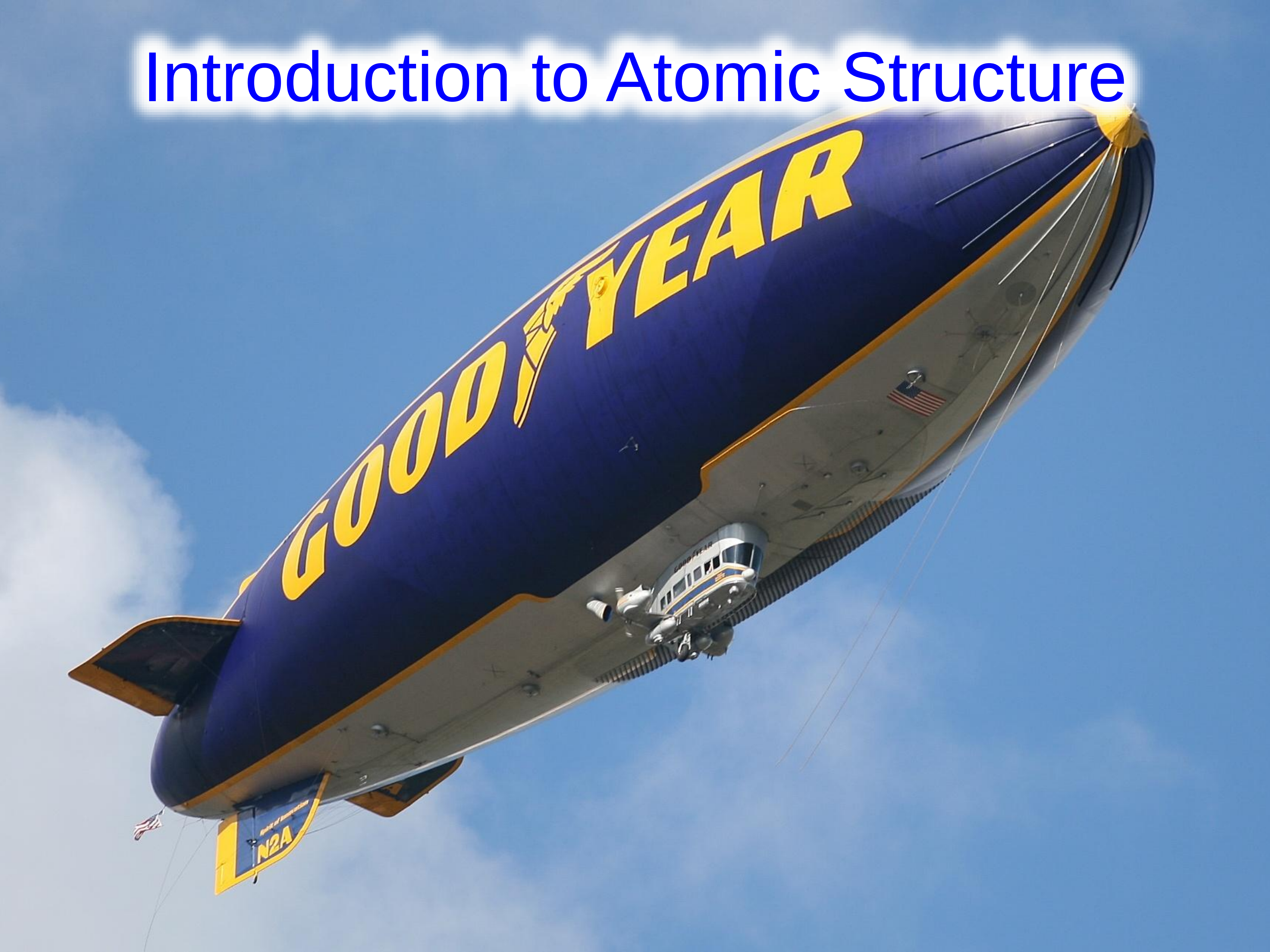
Introduction to Atomic Structure

- 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.

$$\text{hydrogen} + \text{oxygen} \rightarrow \text{water}$$

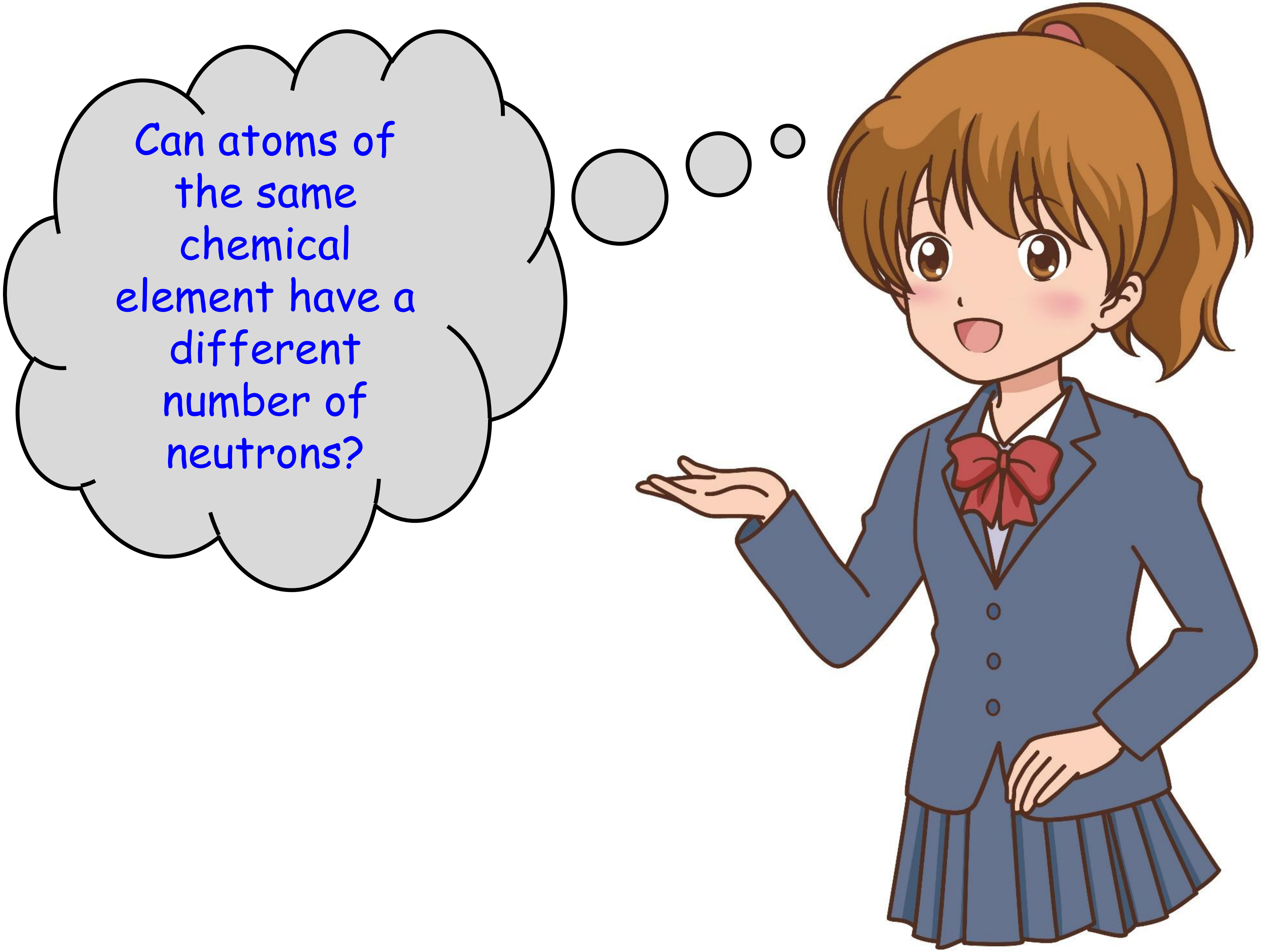
$$2H_2(g) + O_2(g) \rightarrow 2H_2O(l)$$

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.


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

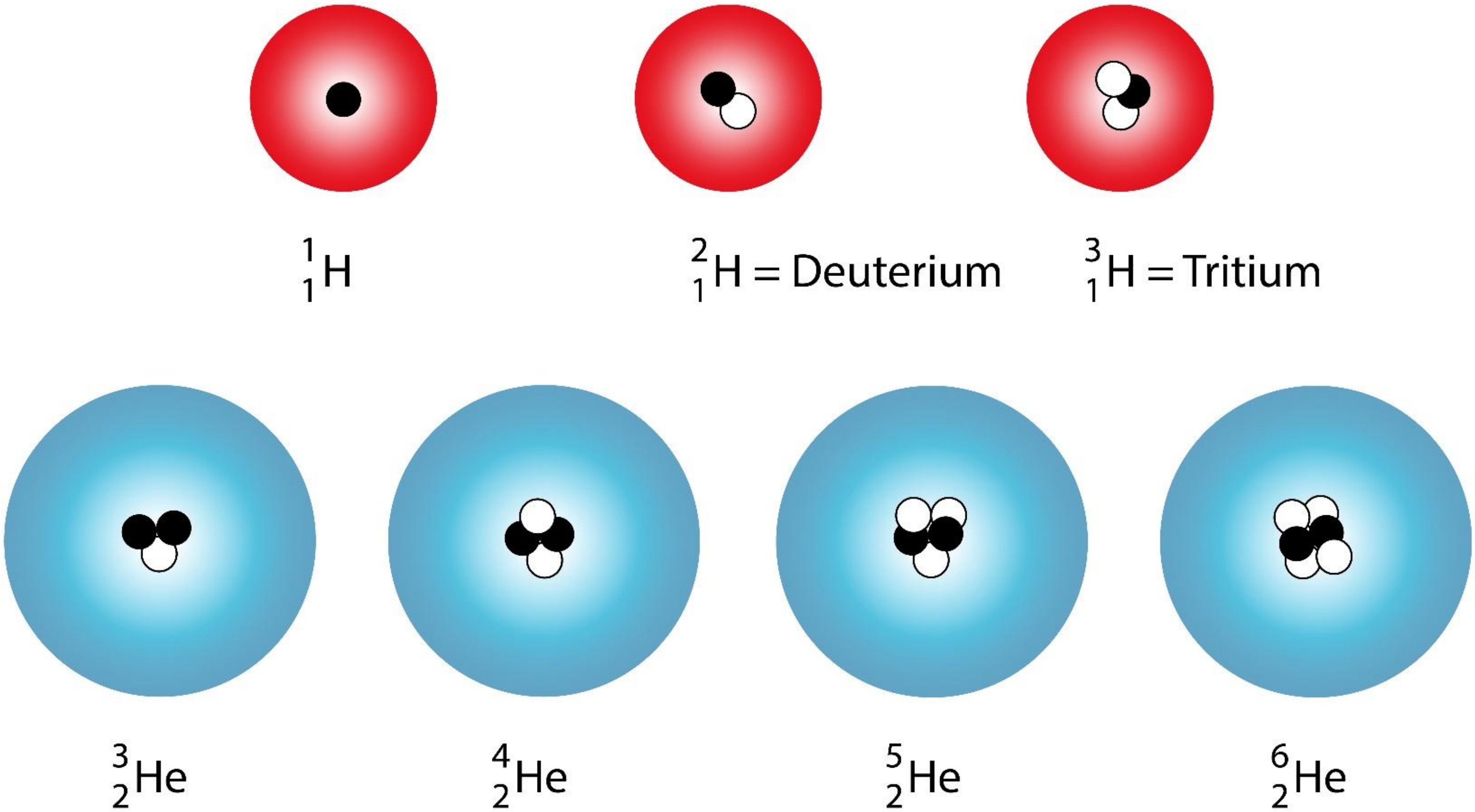
Isotopes

- *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 electron 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



Isotopes

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

Isotopes

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

Isotopes

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

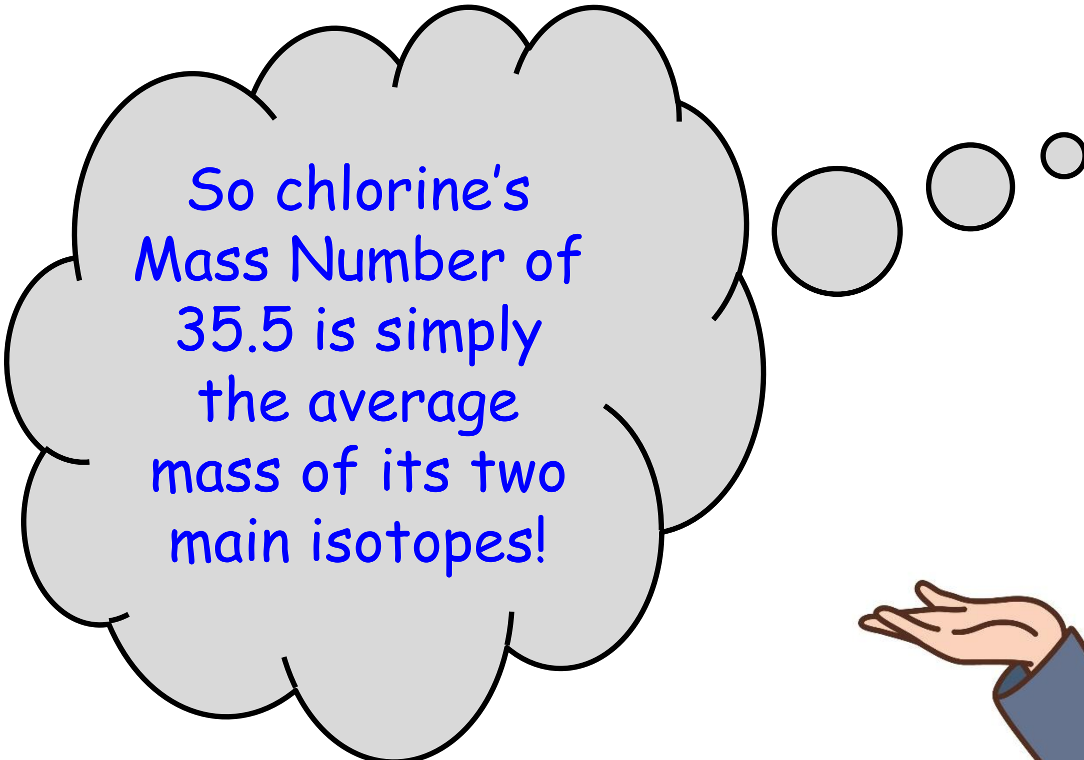
Isotopes

Isotope of Carbon	Number of Protons	Number of Electrons	Number of Neutrons
$^{12}_{6}C$	6	6	12 – 6 = 6
$^{13}_{6}C$	6	6	13 – 6 = 7
$^{14}_{6}C$	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?

Isotopes

- Naturally occurring chlorine has two common isotopes, $^{35}_{17}\text{C}l$ and $^{37}_{17}\text{C}l$.
- $^{35}_{17}\text{C}l$ contains 17 protons and 35 – 17 = 18 neutrons while $^{37}_{17}\text{C}l$ contains 17 protons and 37 – 17 = 20 neutrons.
- The two isotopes of chlorine do not occur in equal quantities. 75% is $^{35}_{17}\text{C}l$ and 25% is $^{37}_{17}\text{C}l$.
- This means that if 100 atoms of chlorine were sampled at random, 75 would be $^{35}_{17}\text{C}l$ and 25 would be $^{37}_{17}\text{C}l$.
- 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.).



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

Isotopes

Definition of Relative Atomic Mass

- *Relative atomic mass* is the average mass of the isotopes of a chemical element compared to $^{1}/_{12}{}^{th}$ the mass of a single atom of $^{12}_{6}C$.

- Because it is a *ratio*, relative atomic mass is dimensionless *i.e.* it does not have any units.

- The $^{12}_{6}C$ isotope has a mass of exactly 12.000. Therefore, $^{1}/_{12}{}^{th}$ the mass of a single atom of $^{12}_{6}C$ is exactly 1.000. This is known as the *atomic mass unit*.

- **Note:** The $^{1}_{1}H$ isotope is *not* used as the reference used to define relative atomic mass because it actually has as mass of 1.008 and not 1.000!

Isotopes

Question: There are two naturally occurring isotopes of copper, $^{63}_{29}Cu$ and $^{65}_{29}Cu$.

69.2% of naturally occurring copper is $^{63}_{29}Cu$.
30.8% of naturally occurring copper is $^{65}_{29}Cu$.

Using this information, calculate the relative atomic mass of copper.

Isotopes

Question: There are two naturally occurring isotopes of copper, $^{63}_{29}Cu$ and $^{65}_{29}Cu$.

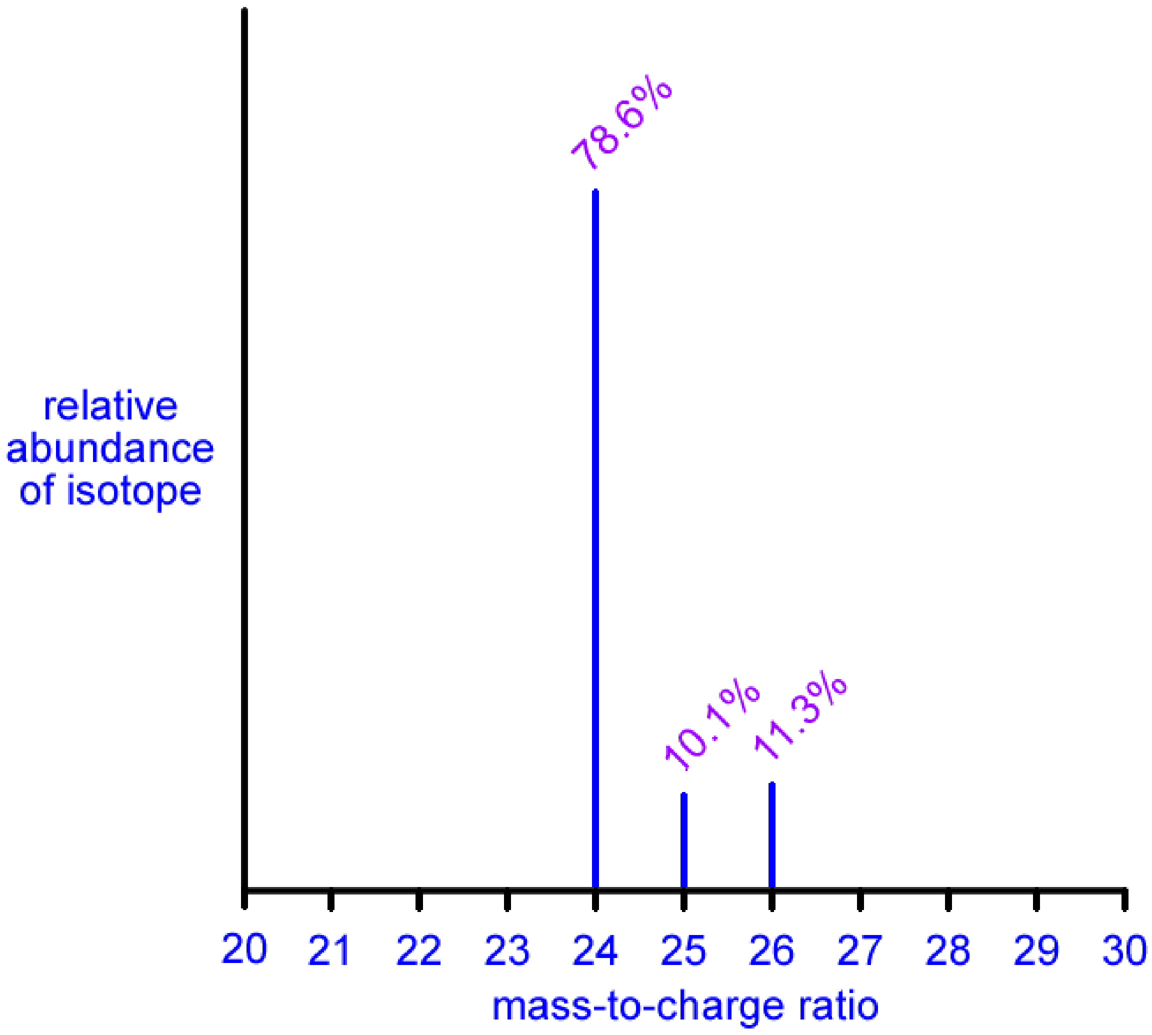
69.2% of naturally occurring copper is $^{63}_{29}Cu$.
30.8% of naturally occurring copper is $^{65}_{29}Cu$.

Using this information, calculate the relative atomic mass of copper.

- Mass of 100 copper atoms = (69.2×63) + (30.8×65).
- Mass of 100 copper atoms = 6361.6.

- Average mass of 1 copper atom = $6361.6 \div 100$.
- Average mass of 1 copper atom = 63.6 (3 s.f.).

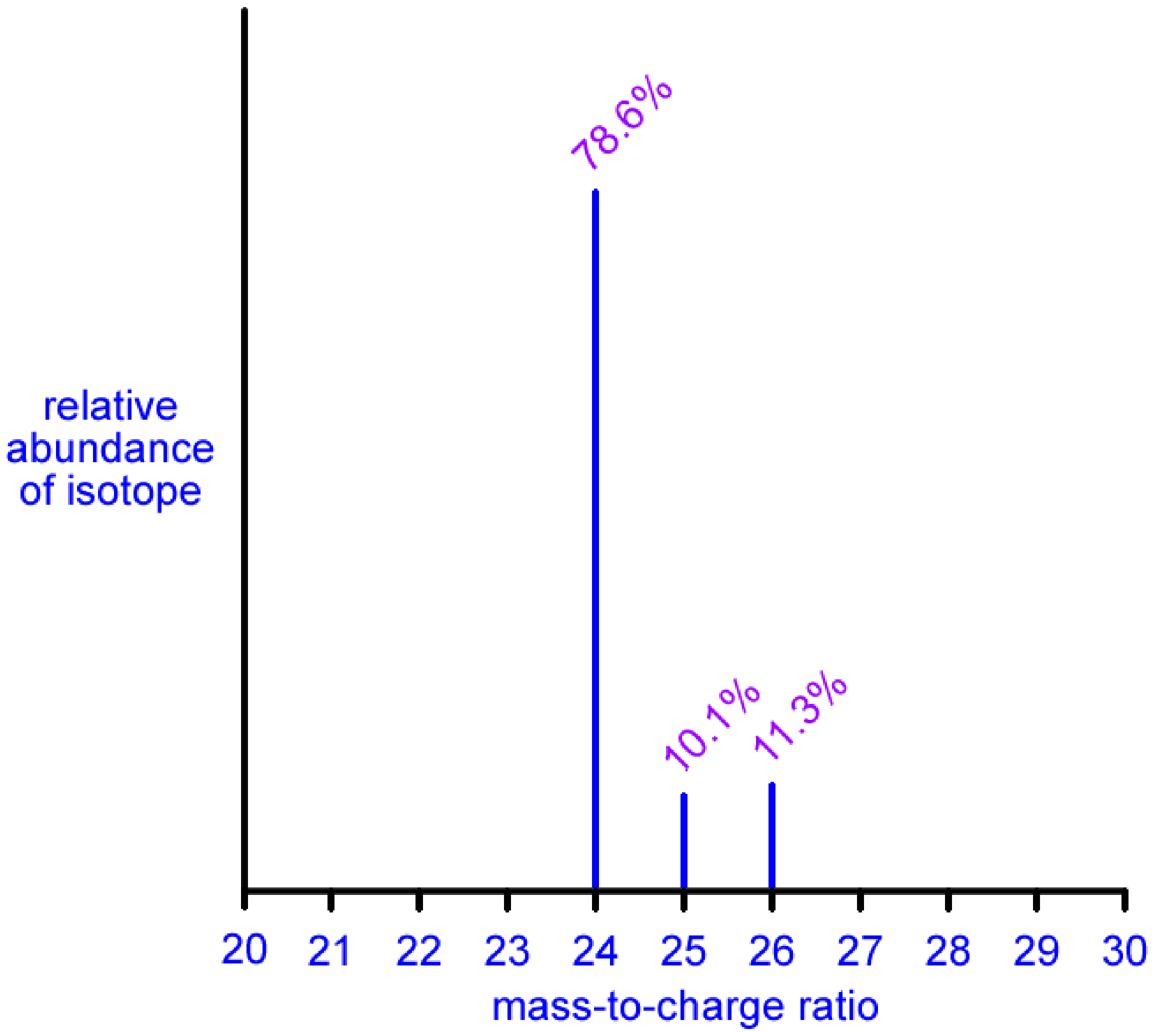
Isotopes



Question:

A sample of naturally occurring magnesium was analysed using a mass spectrometer. The results obtained from the machine are shown. Briefly interpret the results.

Isotopes



Question:

A sample of naturally occurring magnesium was analysed using a mass spectrometer. The results obtained from the machine are shown. Briefly interpret the results.

$^{24}_{12}Mg$ = 78.6%

$^{25}_{12}Mg$ = 10.1%

$^{26}_{12}Mg$ = 11.3%

Isotopes

Question: There are three naturally occurring isotopes of magnesium, $^{24}_{12}Mg$, $^{25}_{12}Mg$ and $^{26}_{12}Mg$.

78.6% of naturally occurring magnesium is $^{24}_{12}Mg$.
10.1% of naturally occurring magnesium is $^{25}_{12}Mg$.
11.3% of naturally occurring magnesium is $^{26}_{12}Mg$.

Using this information, calculate the relative atomic mass of magnesium.

Isotopes

Question: There are three naturally occurring isotopes of magnesium, $^{24}_{12}Mg$, $^{25}_{12}Mg$ and $^{26}_{12}Mg$.

78.6% of naturally occurring magnesium is $^{24}_{12}Mg$.
10.1% of naturally occurring magnesium is $^{25}_{12}Mg$.
11.3% of naturally occurring magnesium is $^{26}_{12}Mg$.

Using this information, calculate the relative atomic mass of magnesium.

- Mass of 100 Mg atoms = $(78.6 \times 24) + (10.1 \times 25) + (11.3 \times 26)$.
- Mass of 100 Mg atoms = 2432.7.

- Average mass of 1 Mg atom = $2432.7 \div 100$.
- Average mass of 1 Mg atom = 24.3 (3 s.f.).

Why do atoms react?

How do atoms react?

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

Essential Understanding for Chemistry

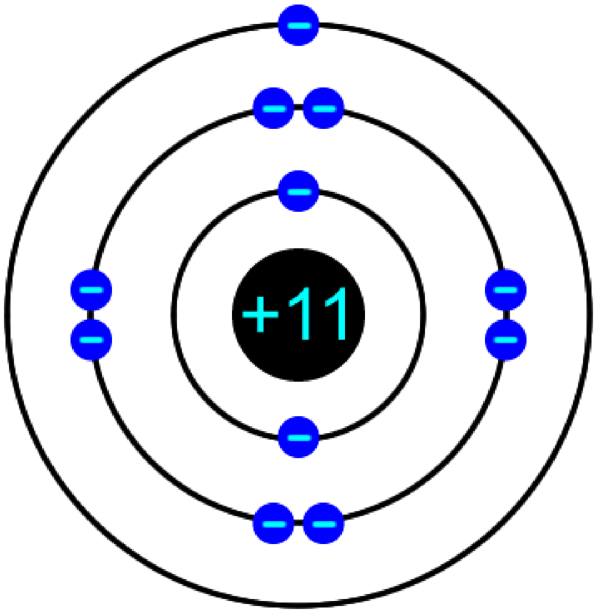
Why do atoms react?

- 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).

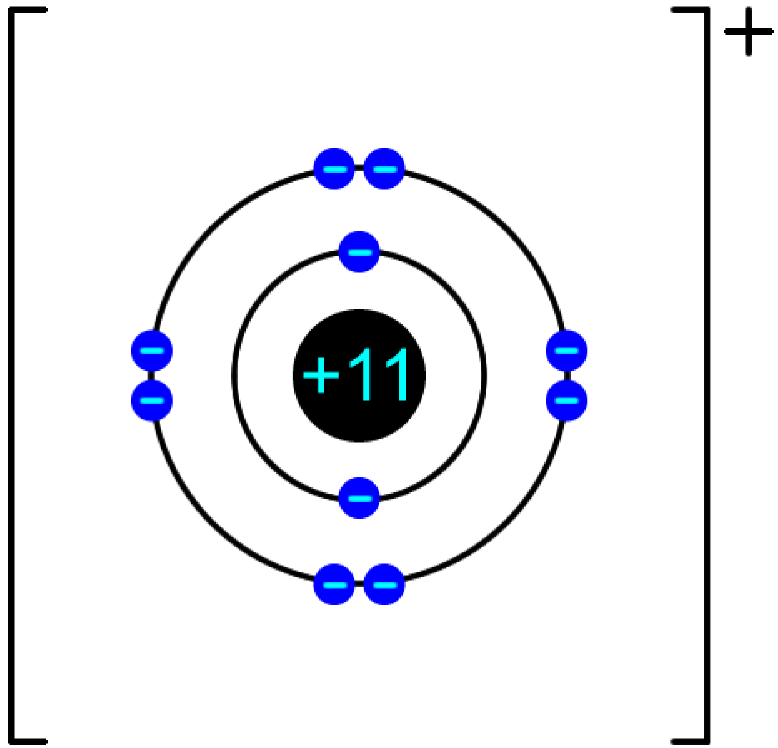
Ions – Anions and Cations



Sodium Atom

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

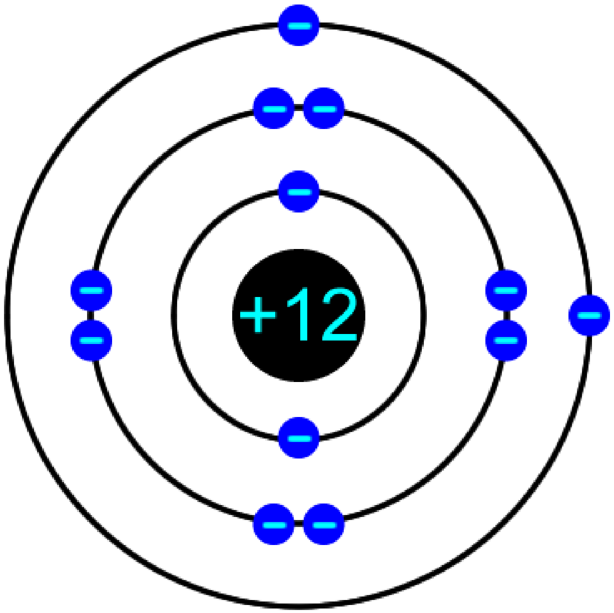
Ions – Anions and Cations



Sodium Ion

…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*.

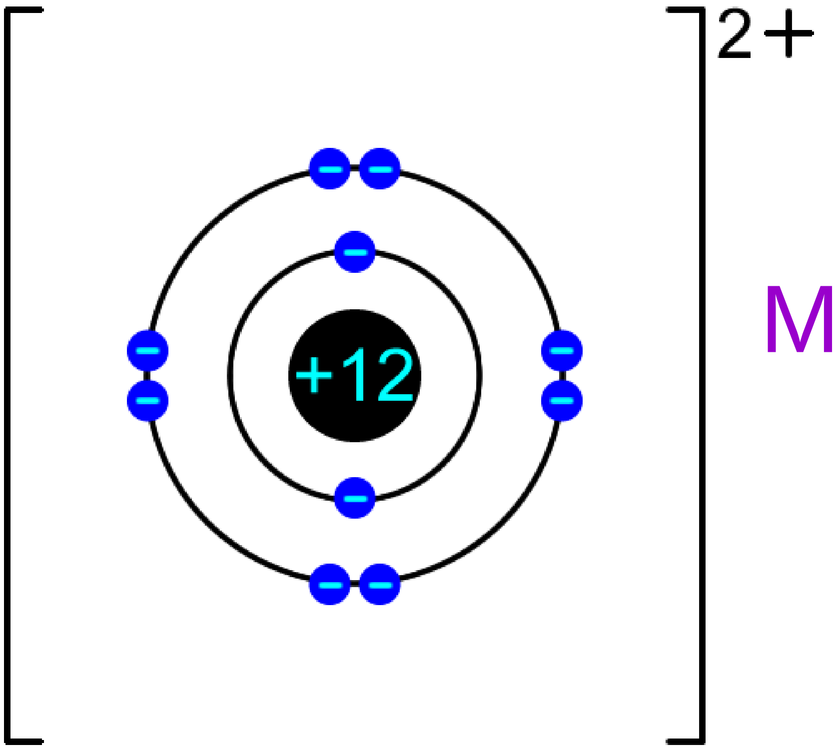
Ions – Anions and Cations



Magnesium Atom

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

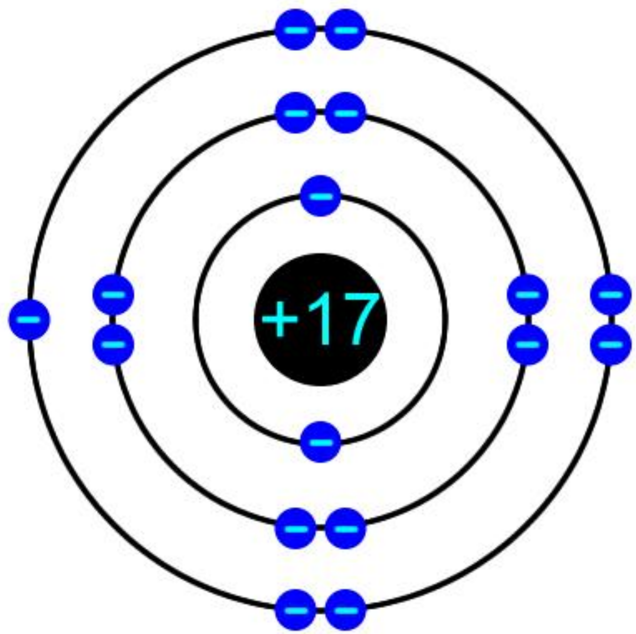
Ions – Anions and Cations



Magnesium Ion

…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*.

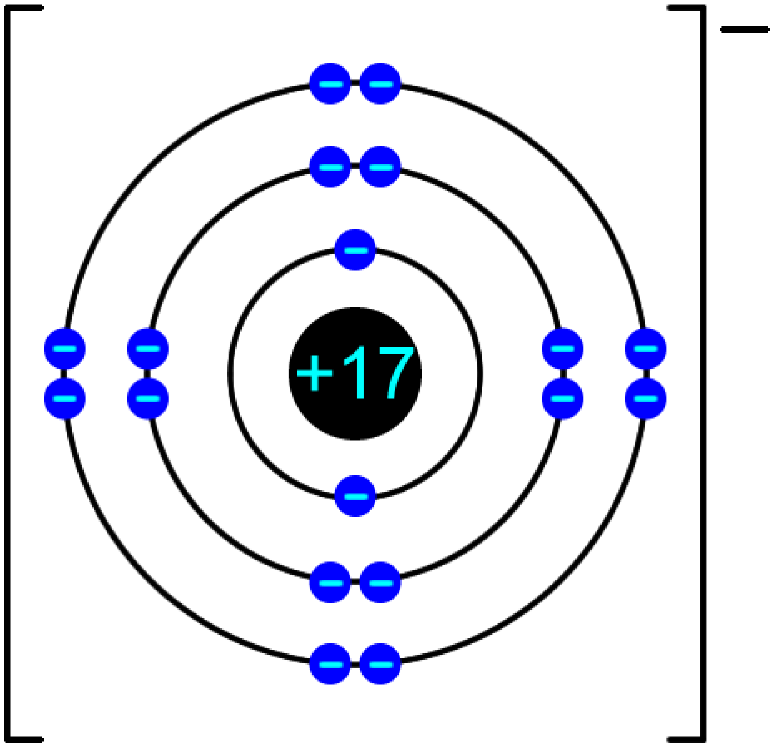
Ions – Anions and Cations



Chlorine Atom

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

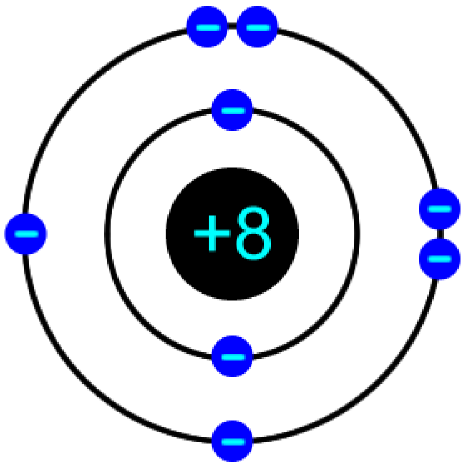
Ions – Anions and Cations



Chloride Ion

…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*.

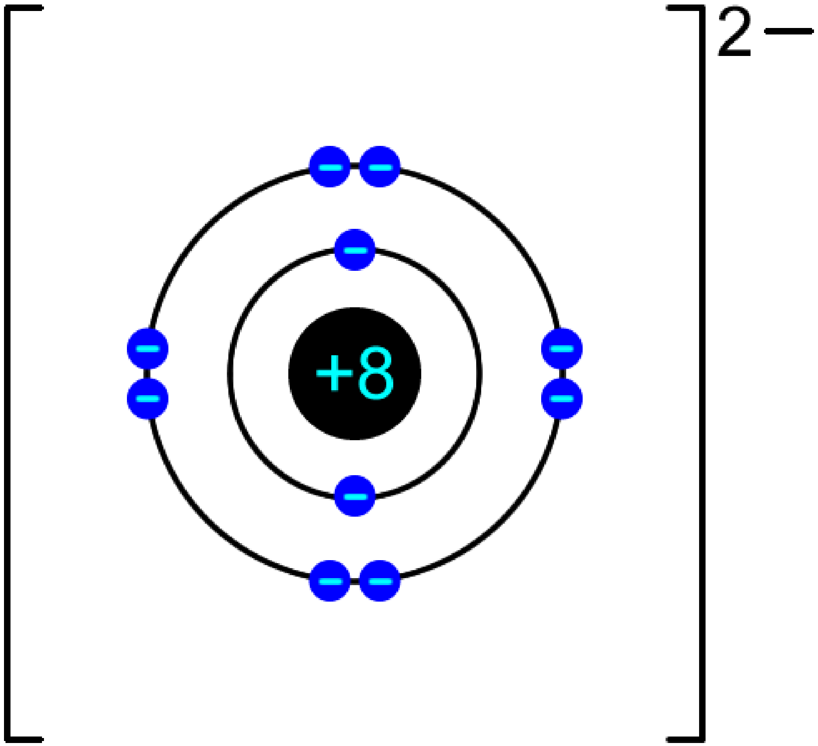
Ions – Anions and Cations



Oxygen
Atom

- Atoms of non-metallic elements *gain* electrons in order to obtain the stable electron 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



Oxide Ion

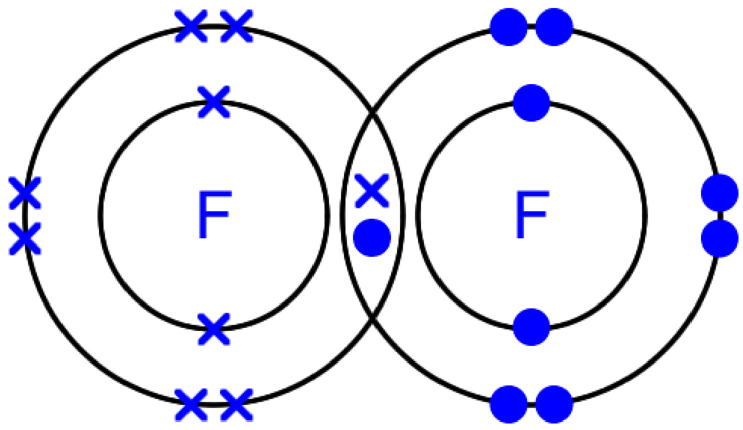
…the resulting particle will have *8 protons* and *10 electrons*. Adding up the charges on all of the protons and all of the electrons *[(+8) + (-10)]* results in an overall charge on the particle of *-2*. The particle is described as an *anion* or *negative ion*.

Simple Covalent Molecules

Two fluorine atoms – 2 × F

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

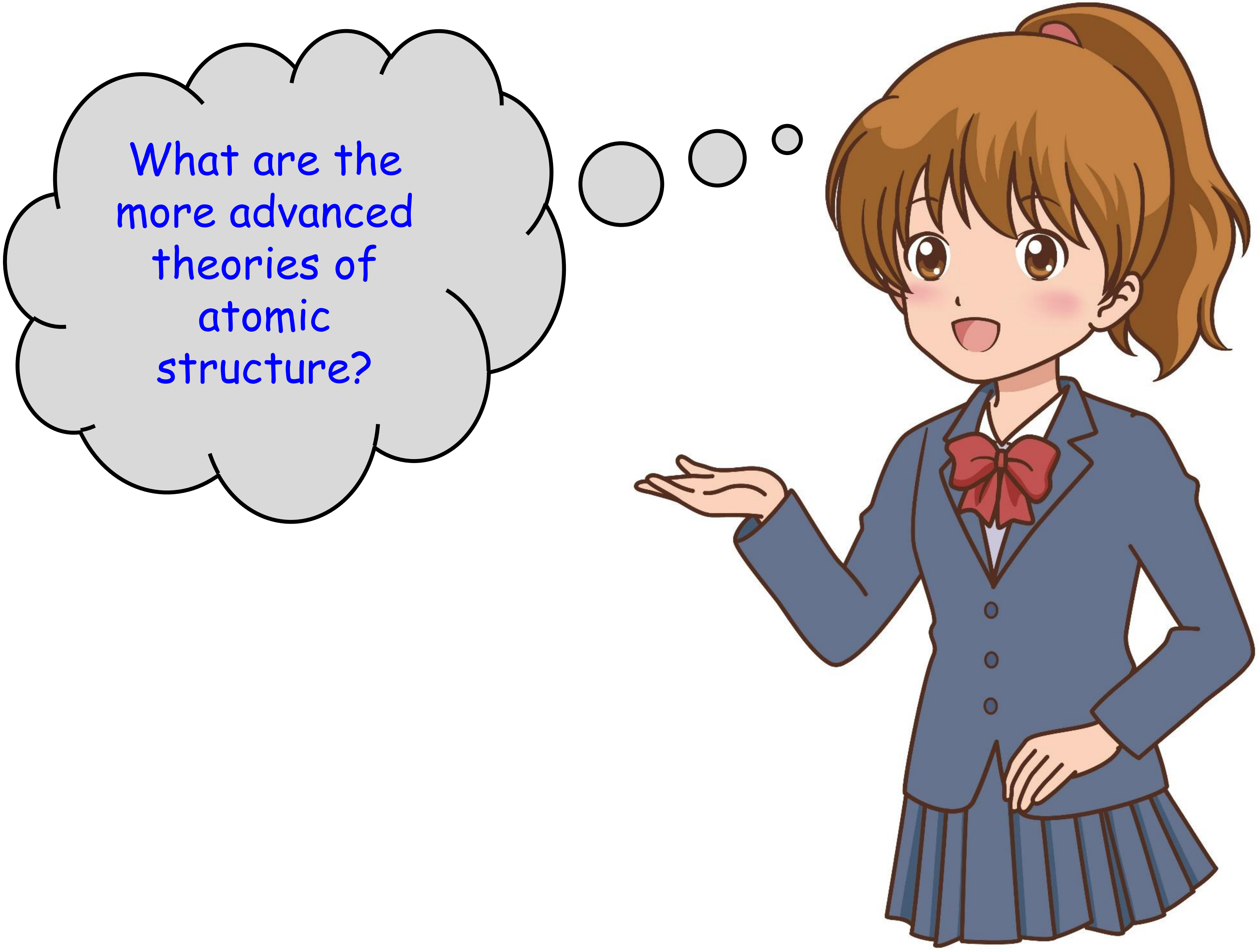
Simple Covalent Molecules



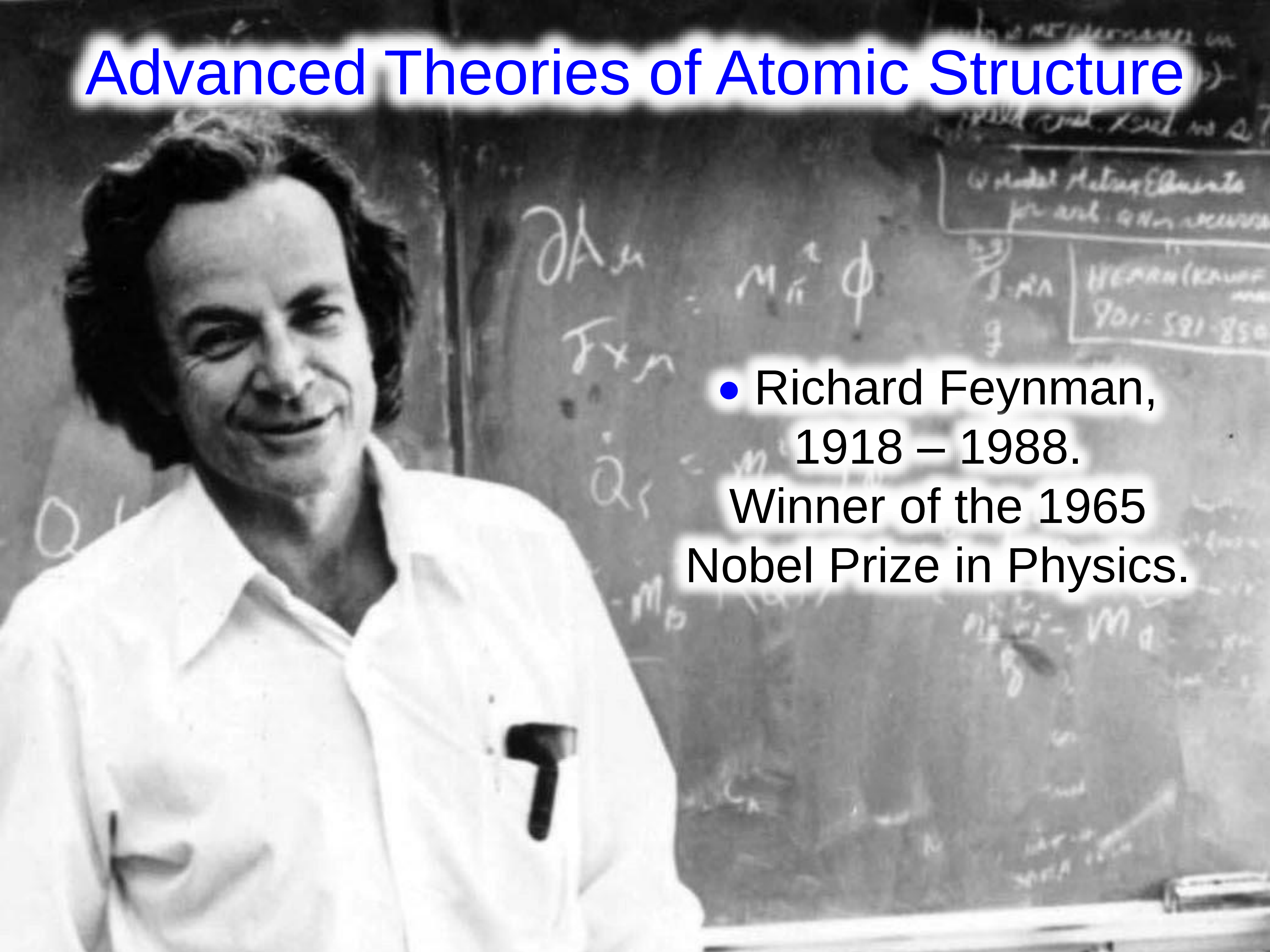
A single molecule of fluorine – F_2

- 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.

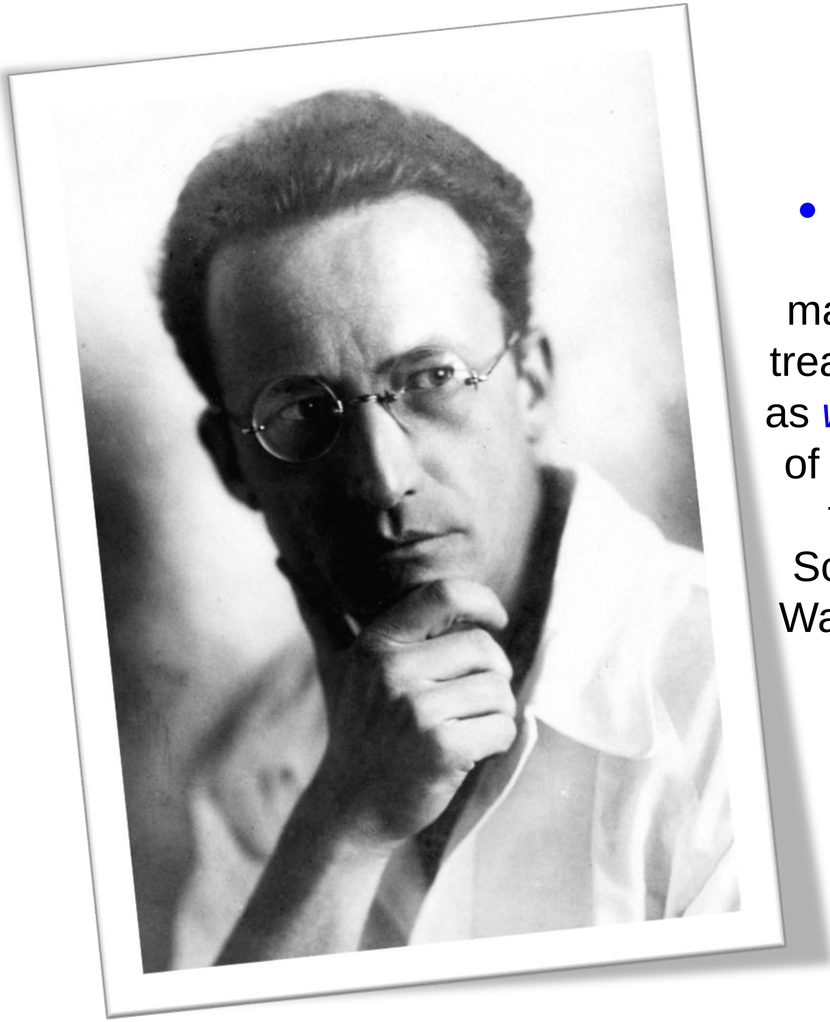
Advanced Theories of Atomic Structure

- 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.

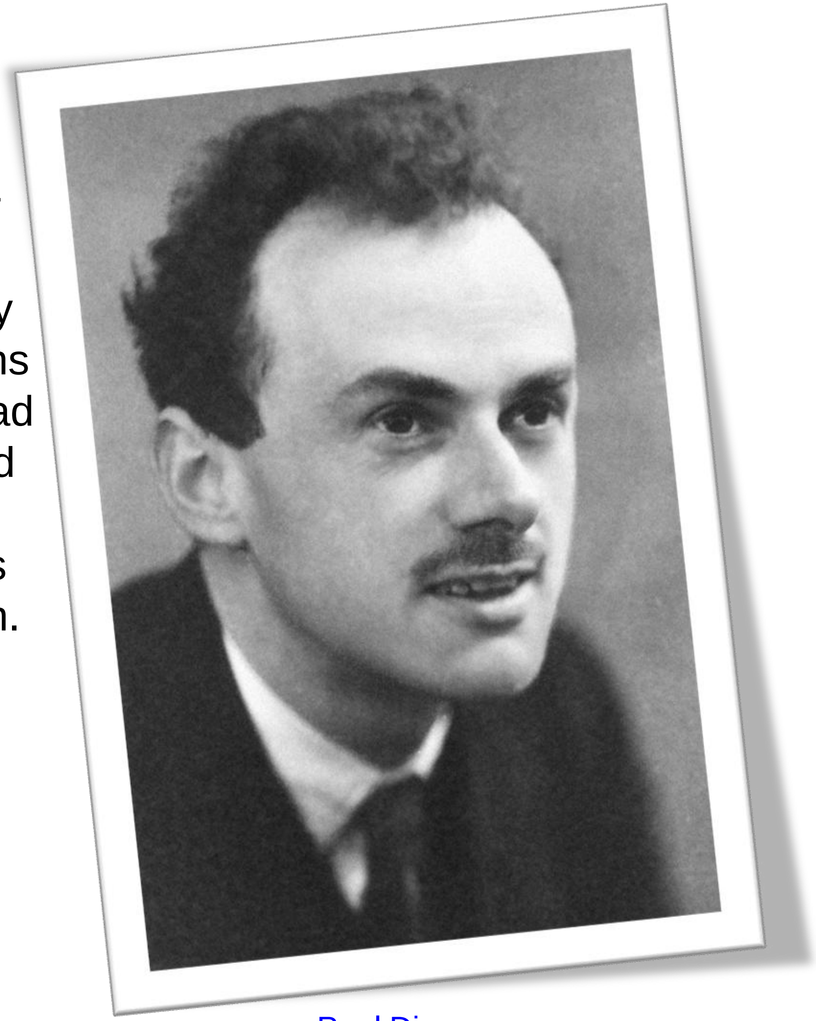
- This is true for a scientific understanding of the atomic orbital structure of the atom. It should be attempted one step at-a-time until the big picture of how electrons orbit the nucleus of the atom starts to emerge.

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:

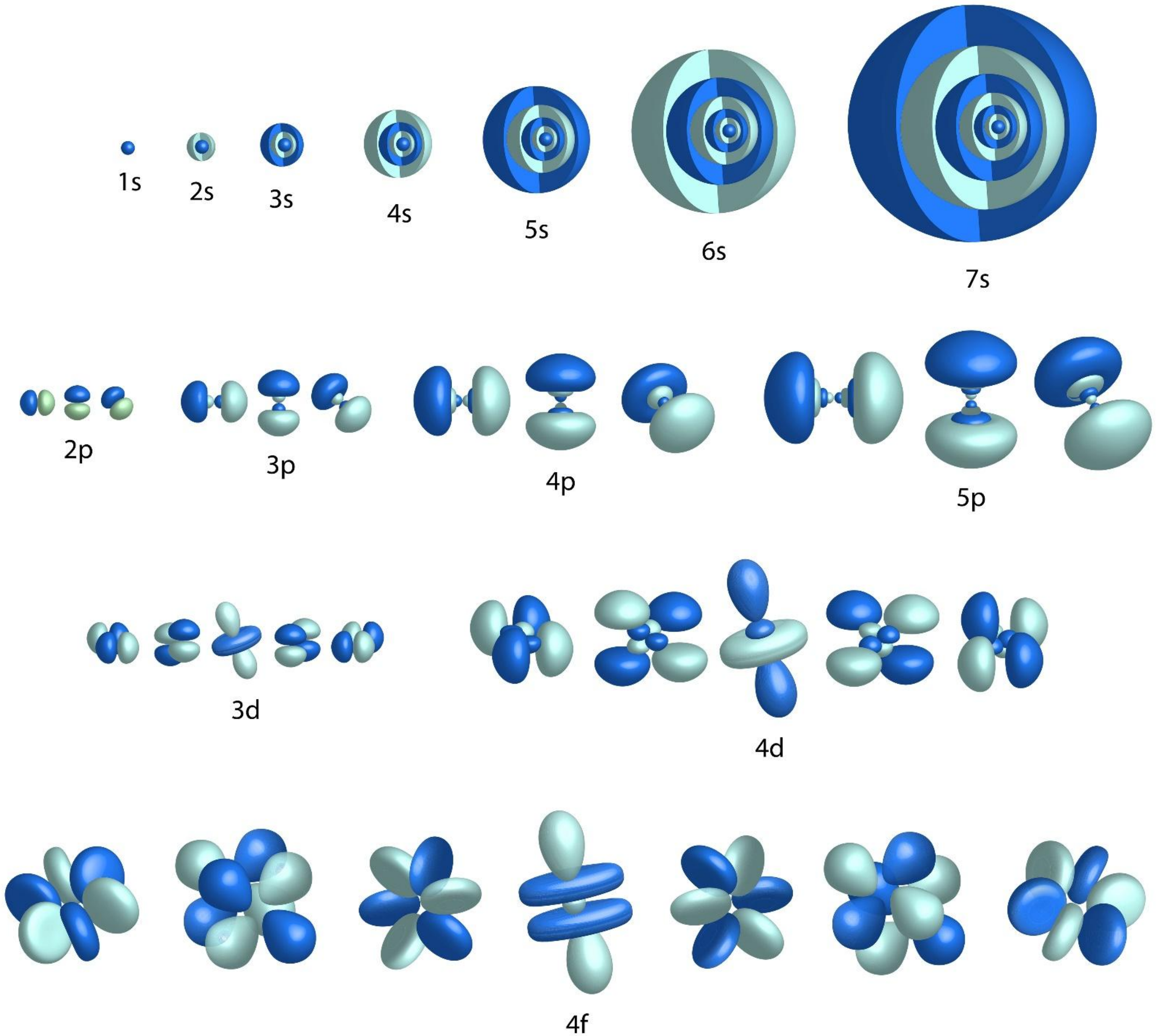
→ First (n): Principle quantum number – the principle quantum shell that the electron occupies.

→ Second (l): The sub-shell within the quantum shell that the electron occupies, *e.g. s-orbitals*, *p-orbitals* or *d-orbitals*.

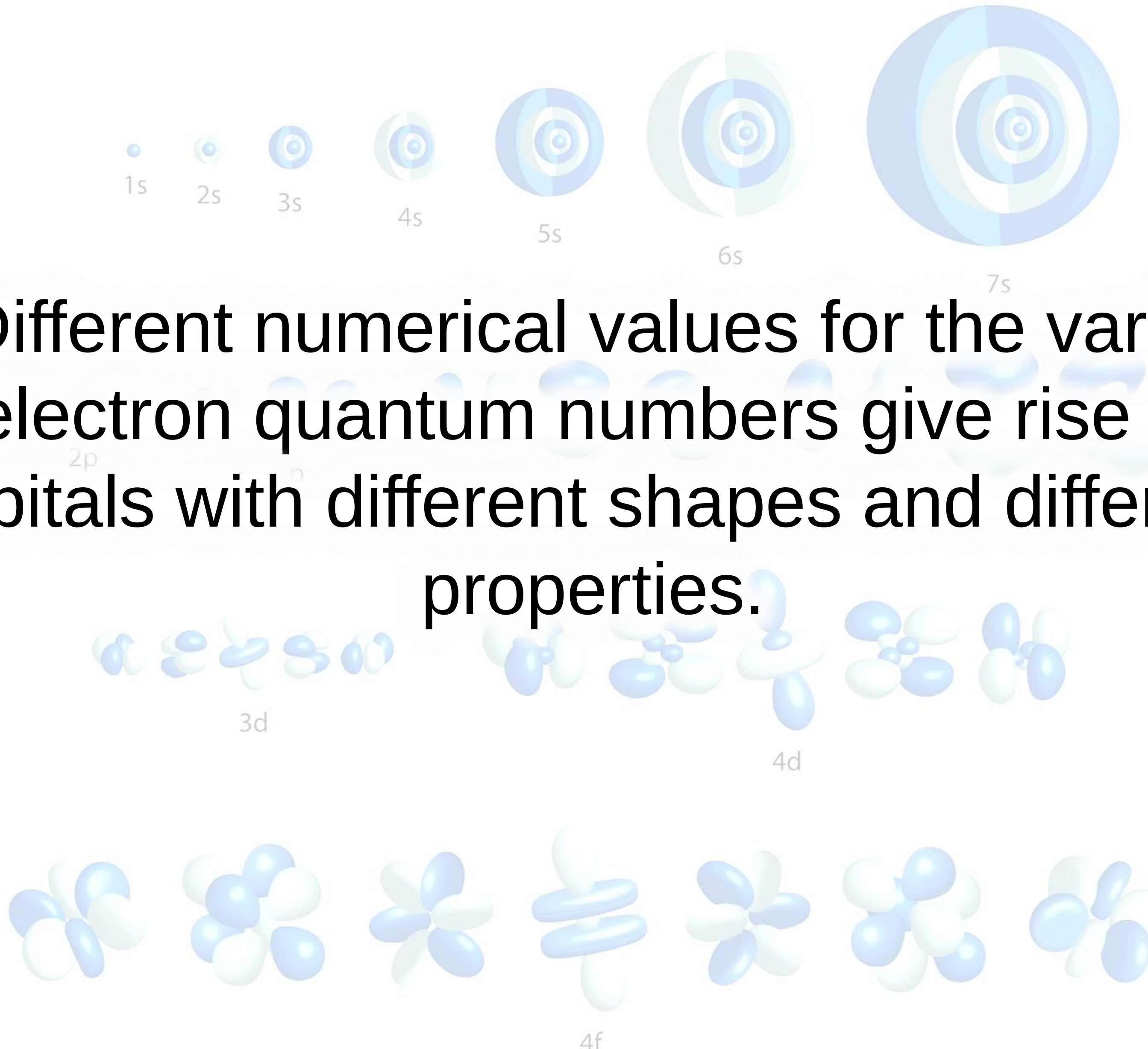
→ Third (m): The orbital within the sub-shell that the electron occupies, *e.g. p-orbitals* are always arranged in groups of three, so if an electron occupies a *p-orbital*, this electron quantum number states exactly which one, the p_x*-orbital*, the p_y*-orbital* or the p_z*-orbital*.

→ 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

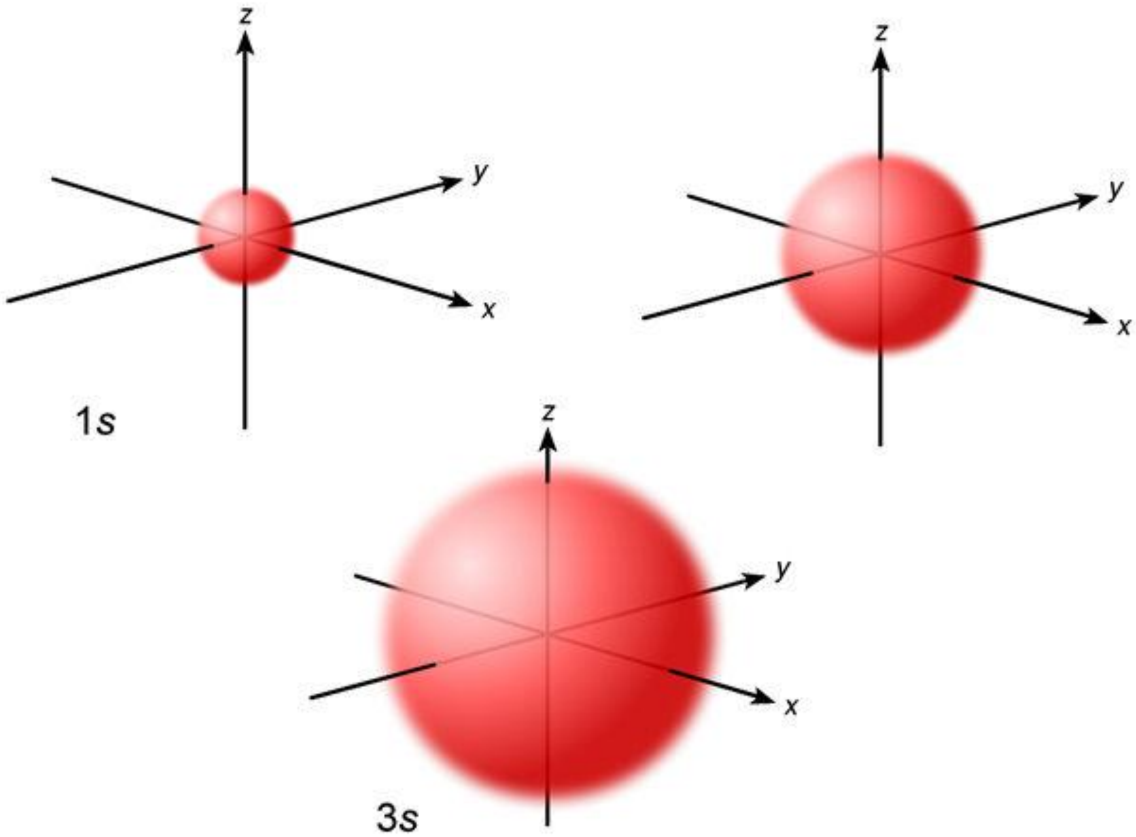


Advanced Theories of Atomic Structure



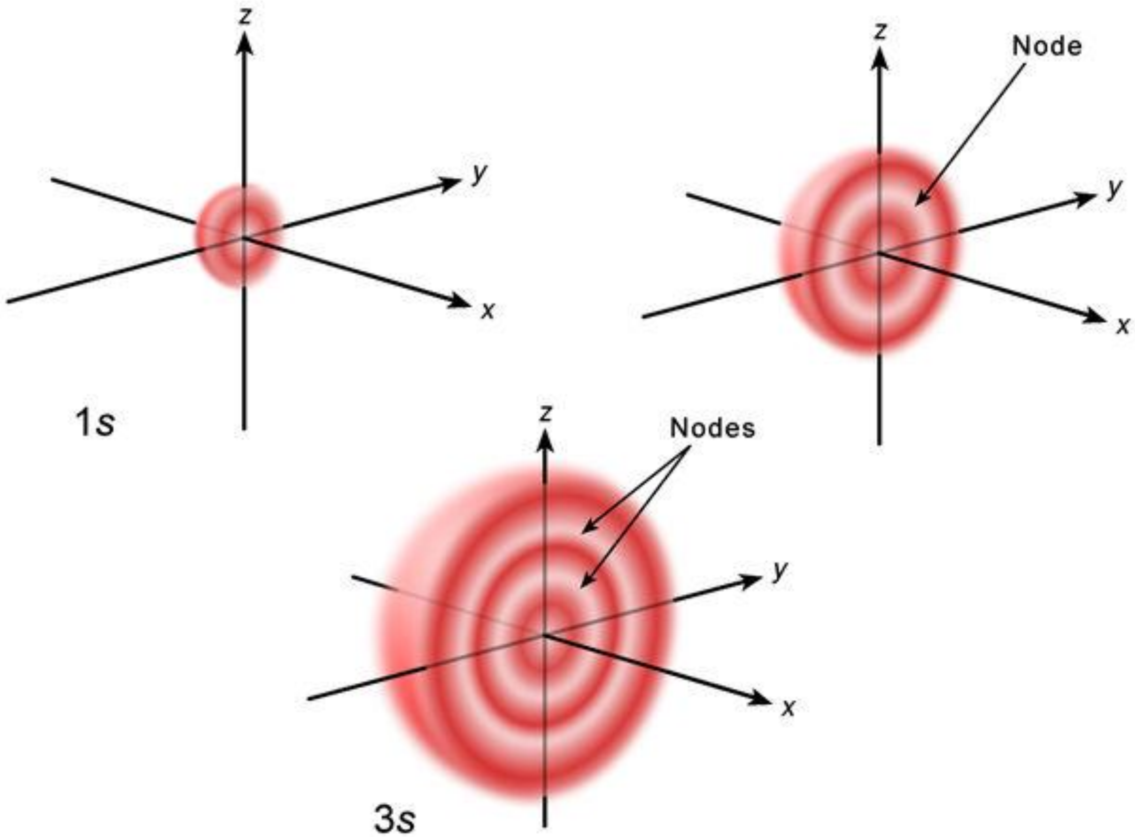
- Different numerical values for the various electron quantum numbers give rise to orbitals with different shapes and different properties.

Advanced Theories of Atomic Structure



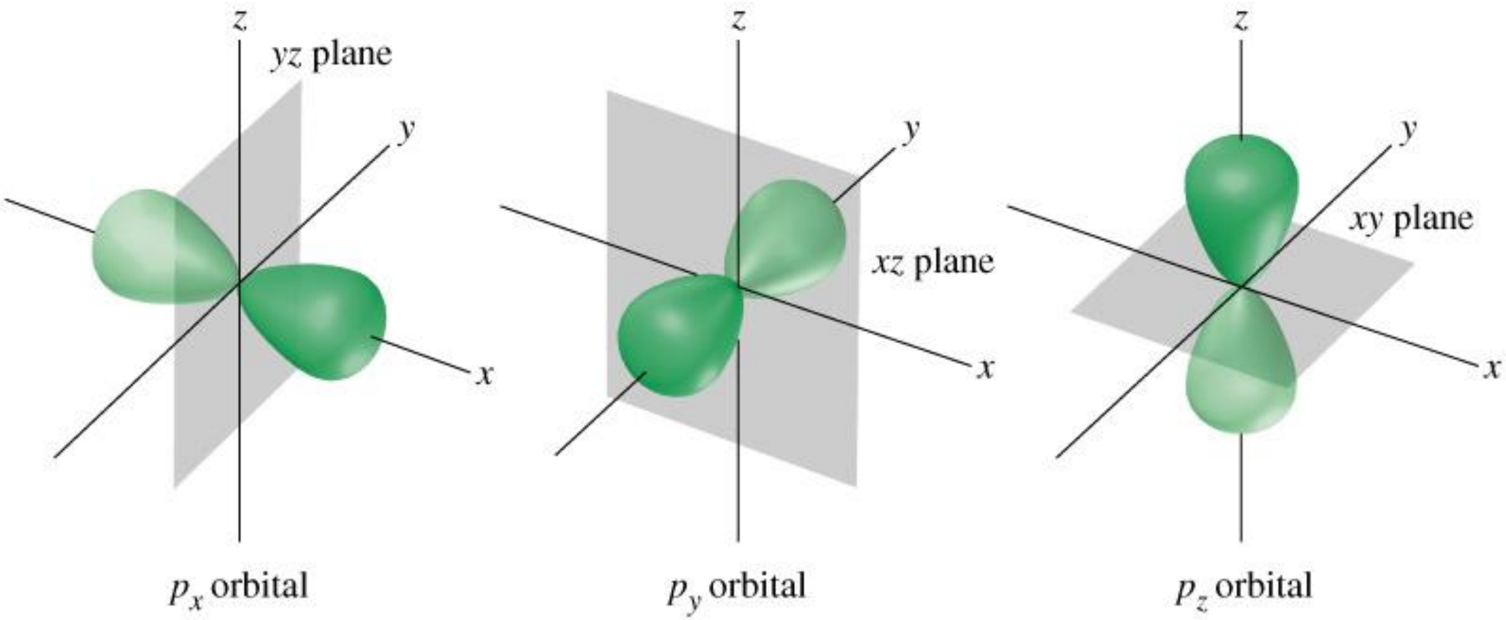
- A graphical solution to Schrödinger's Wave Equation – *s-orbitals*.

Advanced Theories of Atomic Structure



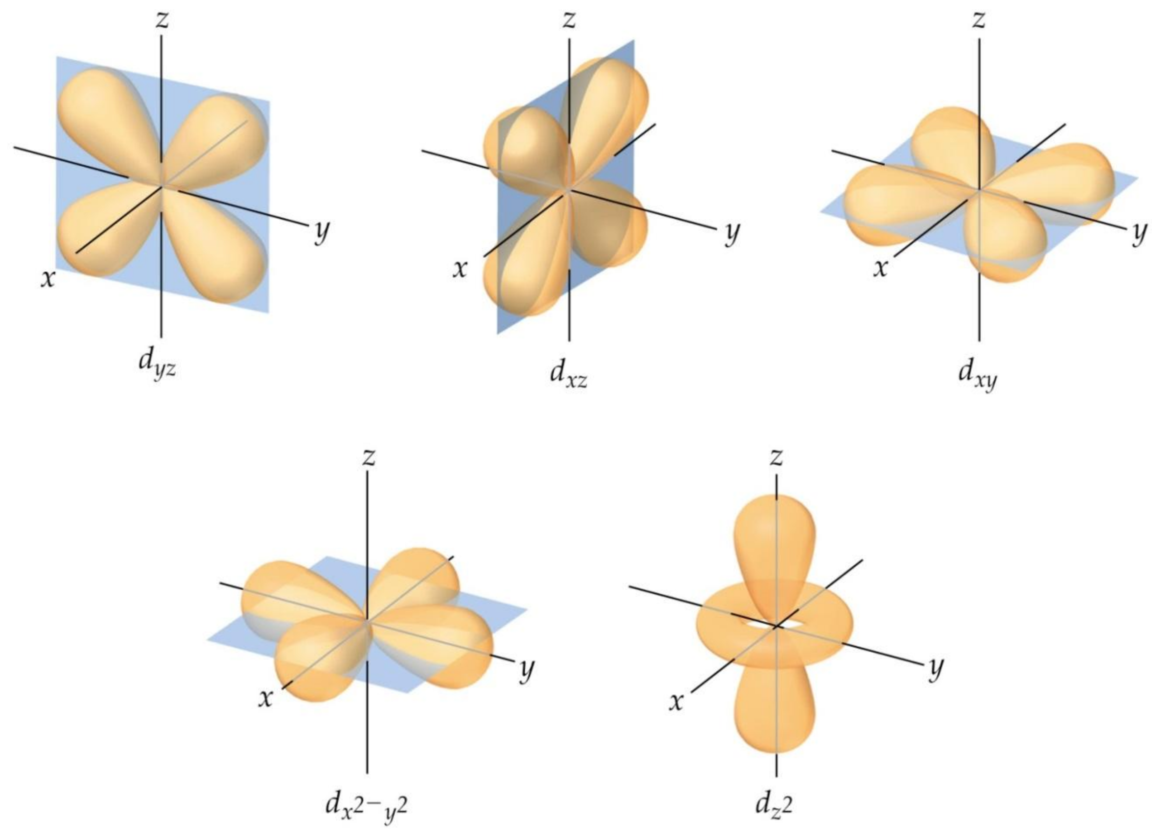
- A graphical solution to Schrödinger's Wave Equation – *s-orbitals*.

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_{xz^2}	f_{z^2}	f_{yz^2}	f_{z^2}	$f_{y(3x^2-y^2)}$
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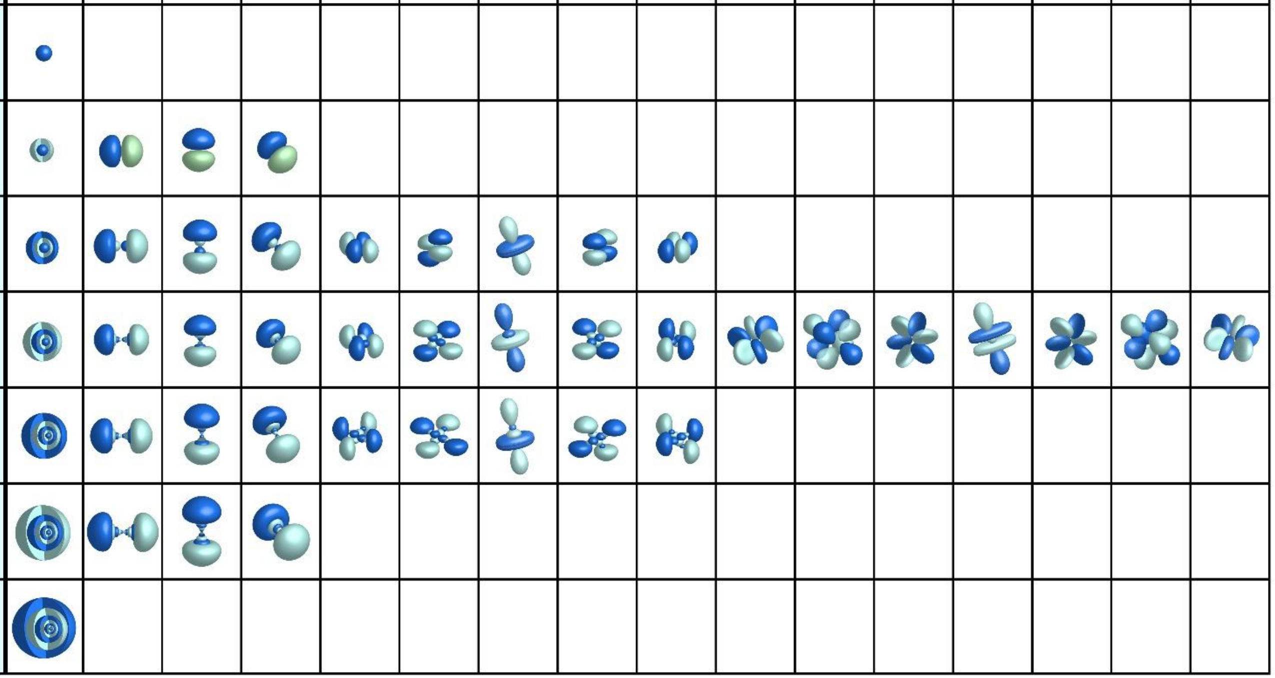
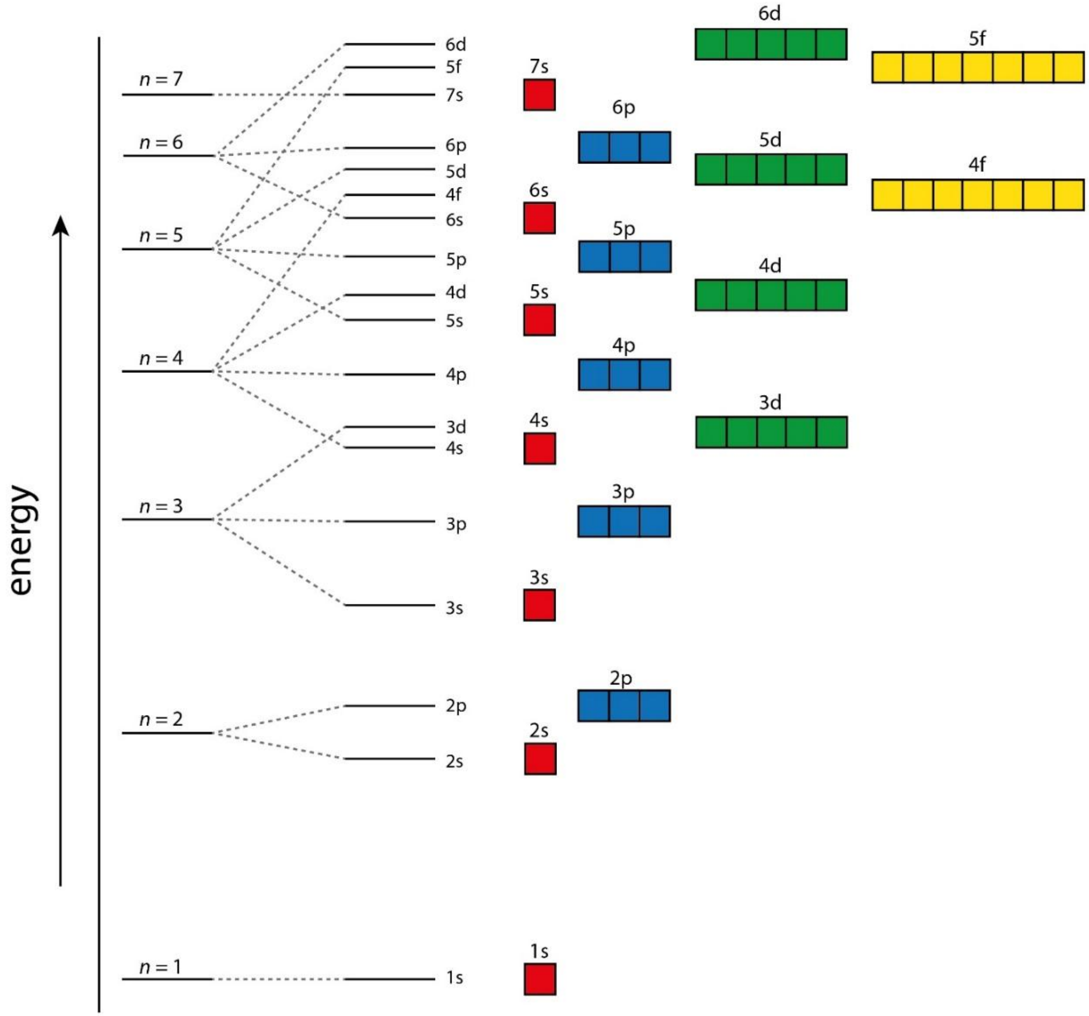


- Diagram showing the various orbitals that arise from different values of the electron quantum numbers *n*, *l* and *m*.

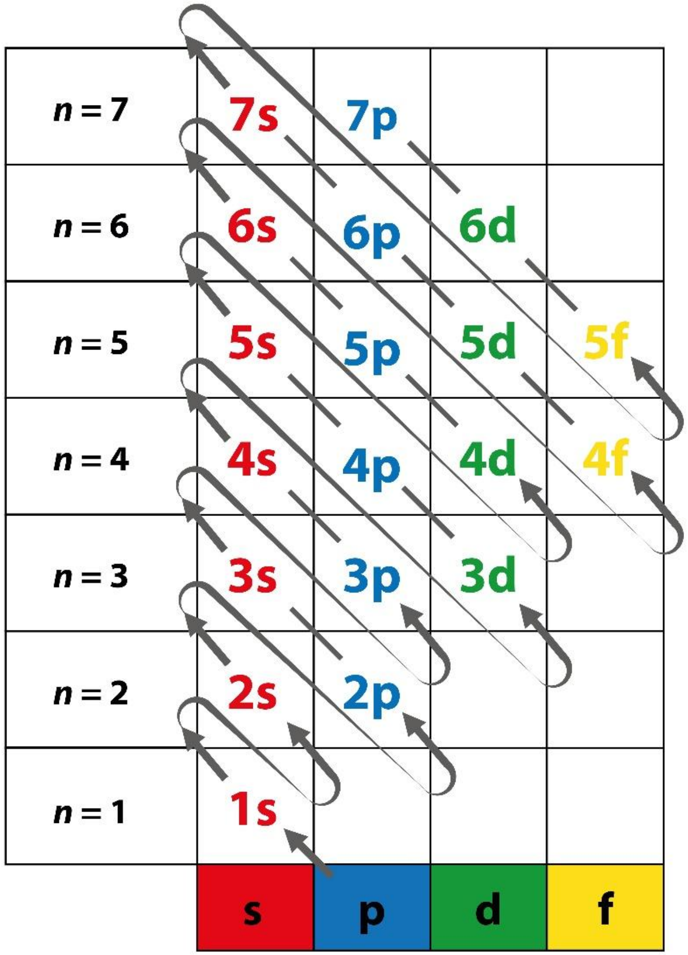
Advanced Theories of Atomic Structure



- 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

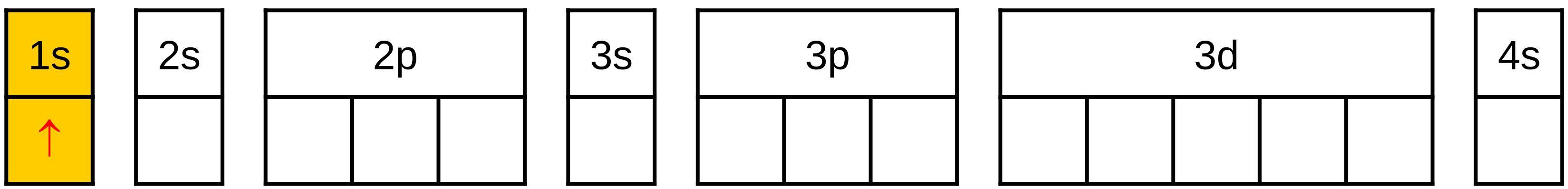
Rules for Filling Atomic Orbitals

- 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.

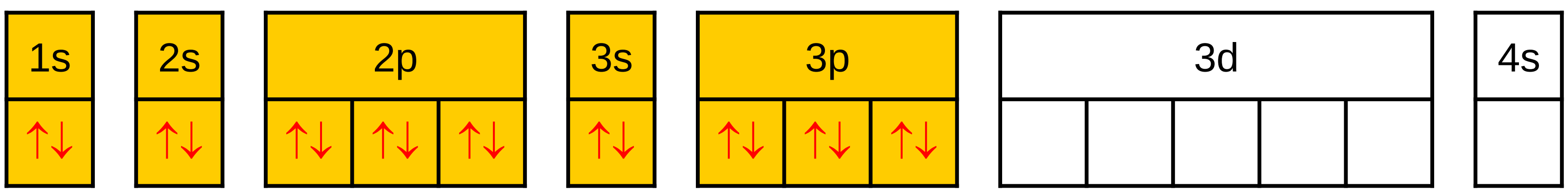
Advanced Theories of Atomic Structure

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 four quantum numbers. Every electron in the same atom must have a unique combination of quantum numbers. Electrons in the same orbital must spin in opposite directions. In atomic orbital diagrams, the spin quantum number is represented by an arrow (↑ or ↓). Two arrows pointing in opposite directions represent two electrons with opposite spin (↑ and ↓).

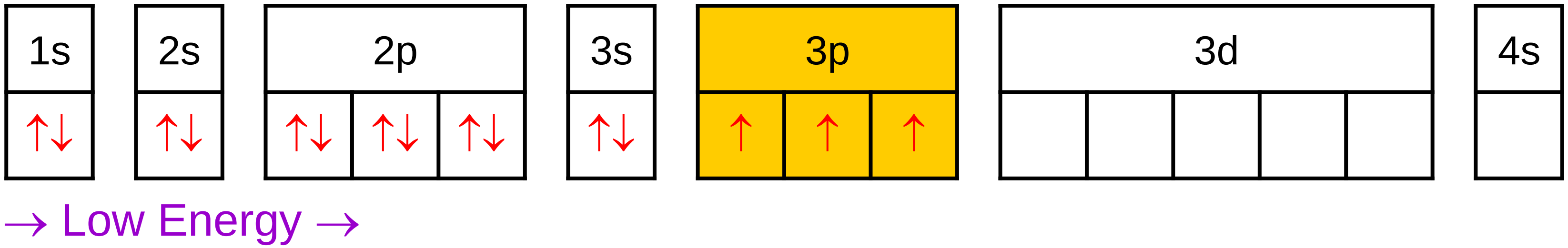
1s	2s	2p			3s	3p			3d					4s
↑↓	↑↓	↑↓	↑↓	↑↓	↑↓	↑↓	↑↓	↑↓	↑↓	↑↓	↑↓	↑↓	↑↓	↑↓

→ Low Energy →

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.



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	1*s*	2
2	2*s*, 2*p*	8
3	3*s*, 3*p*, 3*d*	18
4	4*s*, 4*p*, 4*d*, 4*f*	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	1*s*	2
2	2*s*, 2*p*	8
3	3*s*, 3*p*, 3*d*	18
4	4*s*, 4*p*, 4*d*, 4*f*	32

Advanced Theories of Atomic Structure

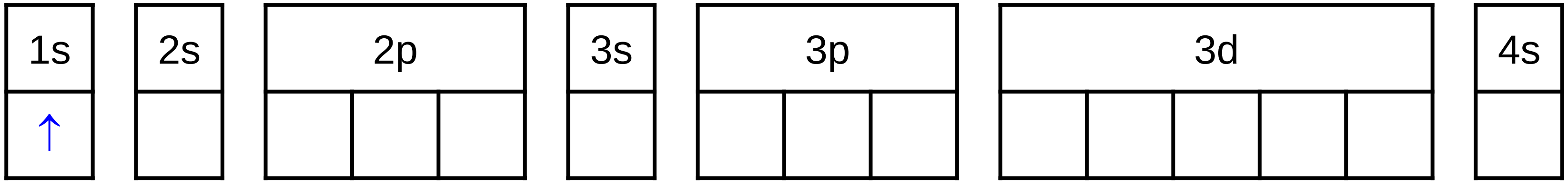
Shapes and Occurrence of Atomic Orbitals

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

Advanced Theories of Atomic Structure

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

- Lower energy than 3d

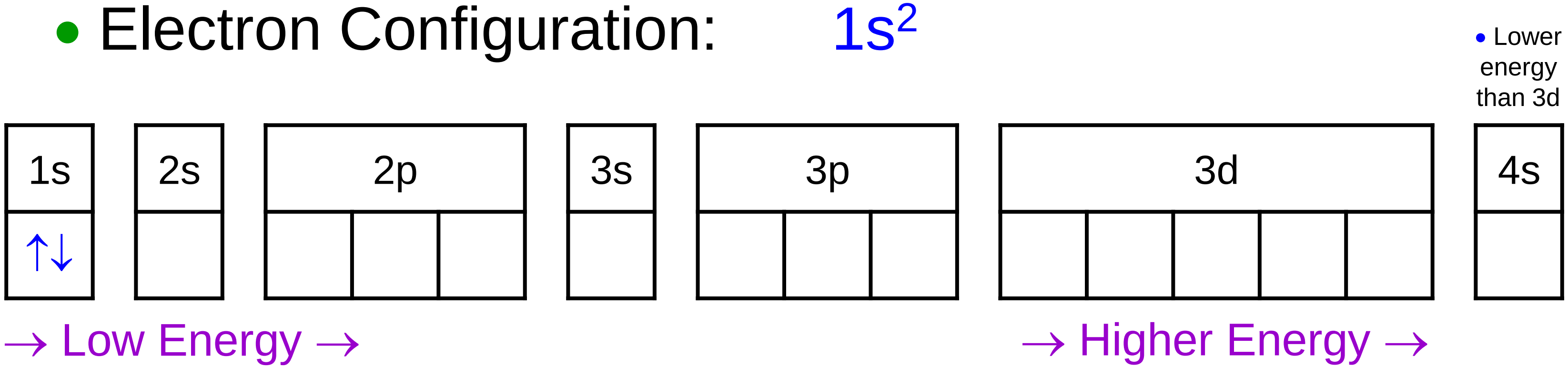


→ Low Energy → → Higher Energy →

- Electrons are represented by *arrows* (↑ and ↓) which fill atomic orbitals that are represented by *boxes*.

Advanced Theories of Atomic Structure

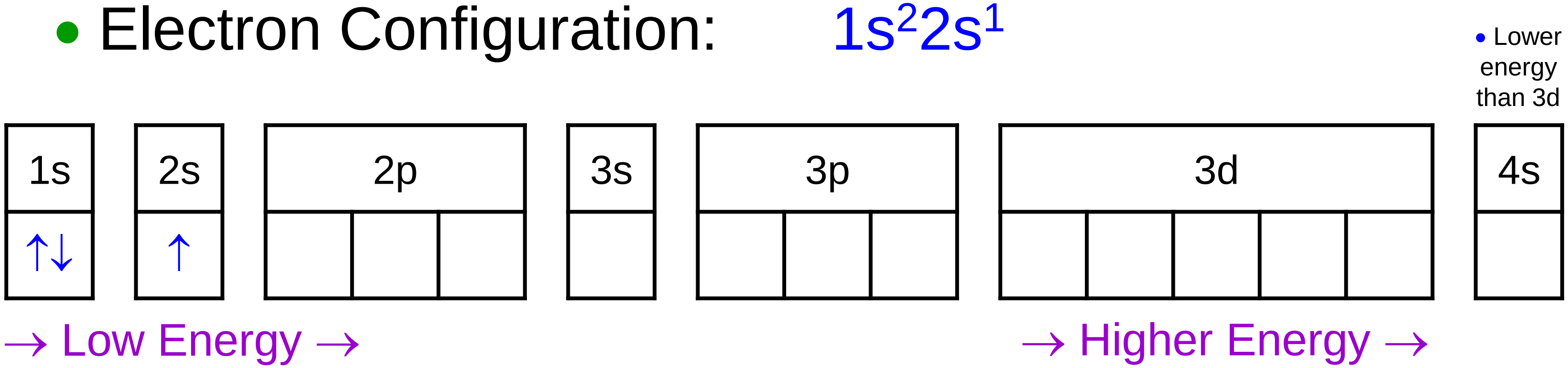
- Atomic Number: 2
- Name: Helium
- Symbol: He
- Electron Configuration: $1s^2$



- Electrons are represented by *arrows* (↑ and ↓) which fill atomic orbitals that are represented by *boxes*.

Advanced Theories of Atomic Structure

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

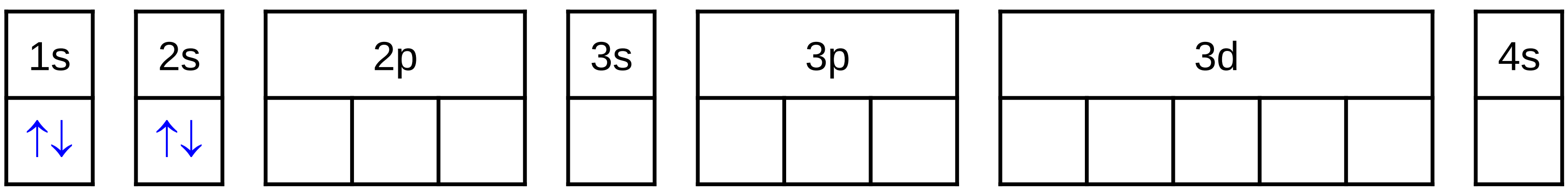


- Electrons are represented by *arrows* (↑ and ↓) which fill atomic orbitals that are represented by *boxes*.
- Electrons occupy atomic orbitals from the lowest energy to the highest energy. This is known as the *Aufbau Principle*.

Advanced Theories of Atomic Structure

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

• Lower energy than 3d

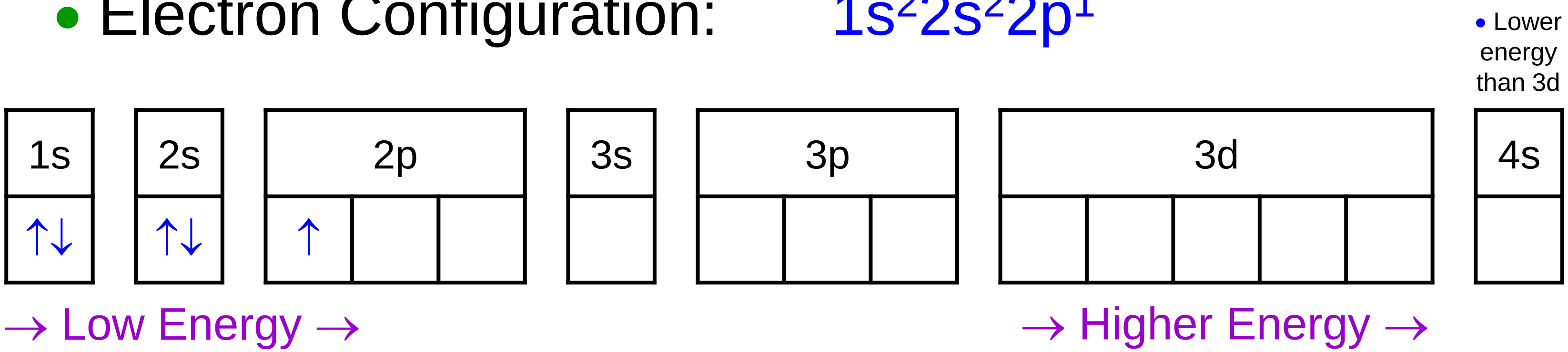


→ Low Energy → → Higher Energy →

- 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: 5
- Name: Boron
- Symbol: B
- Electron Configuration: $1s^2 2s^2 2p^1$



- 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
- Electron Configuration: $1s^2 2s^2 2p^2$

• Lower energy than 3d

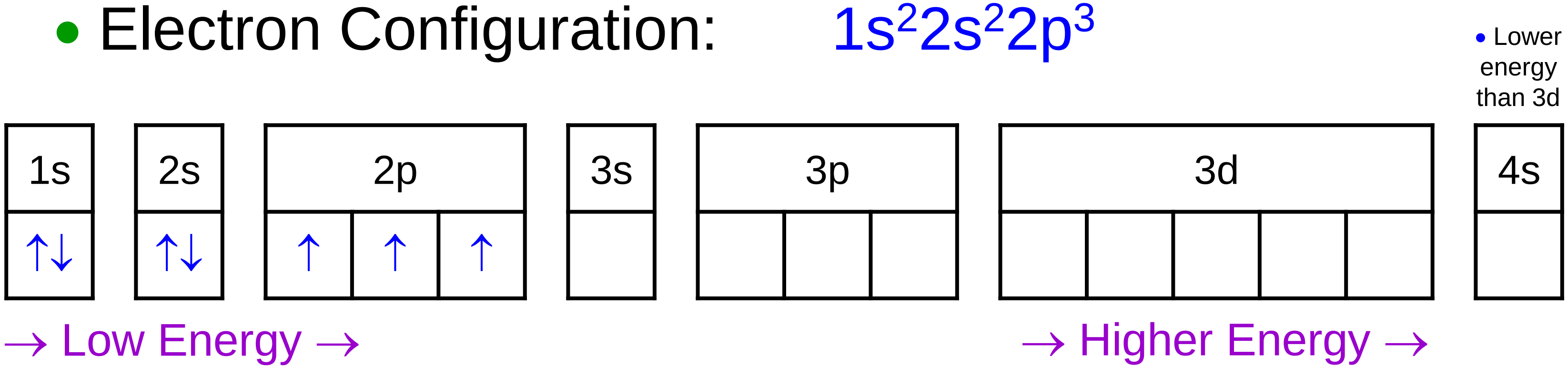
1s	2s	2p	3s	3p	3d	4s
↑↓	↑↓	↑ ↑				

→ Low Energy → → Higher Energy →

- 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.

Advanced Theories of Atomic Structure

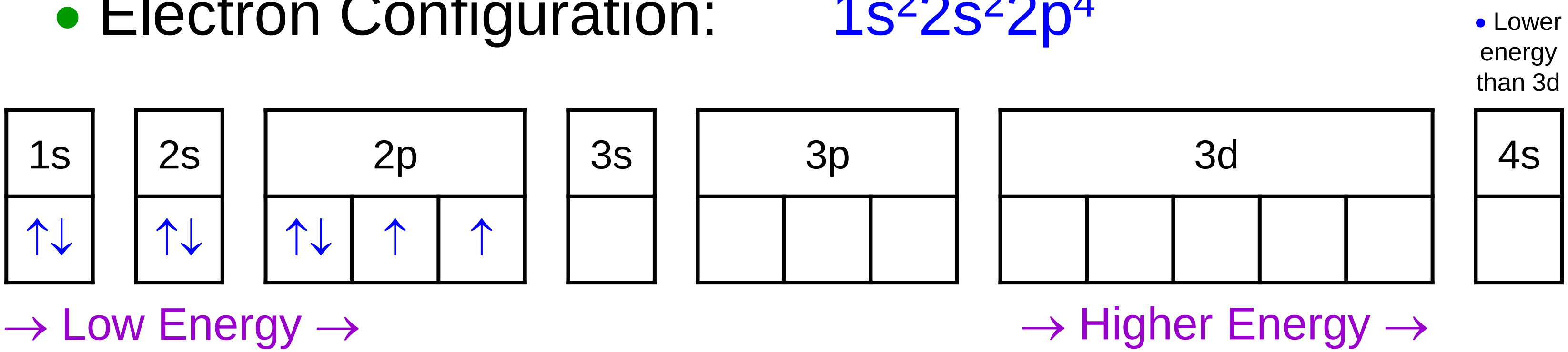
- Atomic Number: 7
- Name: Nitrogen
- Symbol: N
- Electron Configuration: $1s^2 2s^2 2p^3$



- 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.

Advanced Theories of Atomic Structure

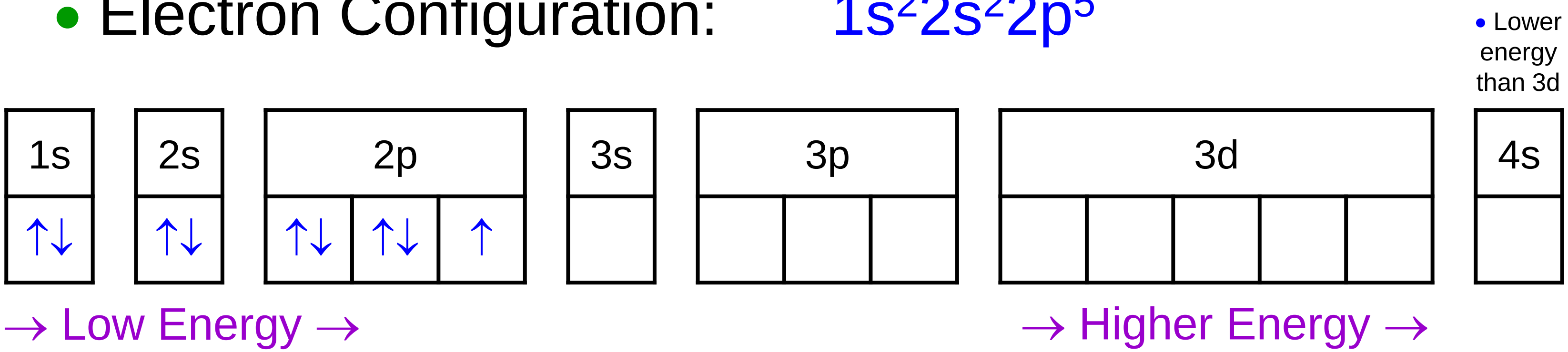
- Atomic Number: 8
- Name: Oxygen
- Symbol: O
- Electron Configuration: $1s^2 2s^2 2p^4$



- 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.

Advanced Theories of Atomic Structure

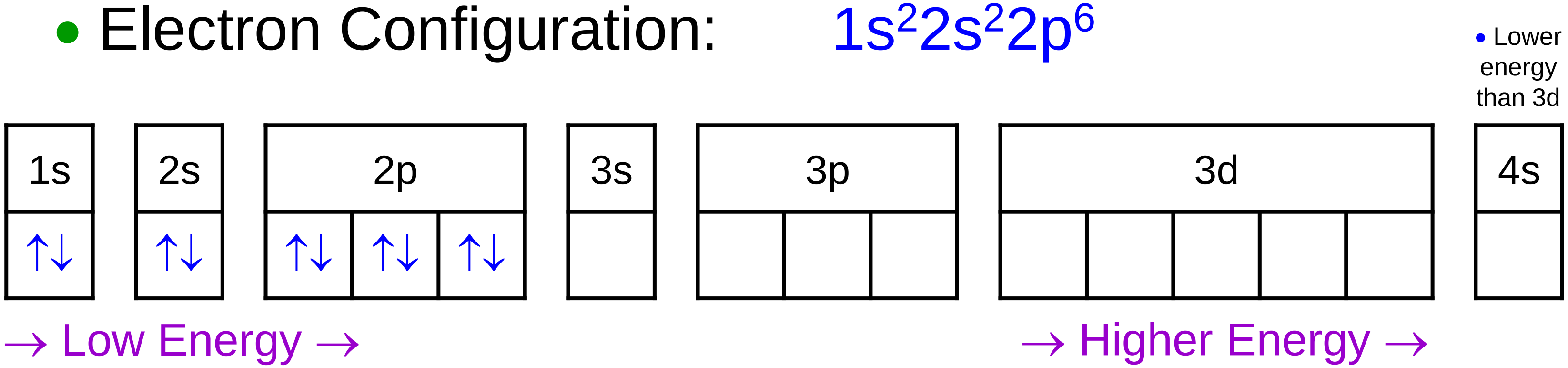
- Atomic Number: 9
- Name: Fluorine
- Symbol: F
- Electron Configuration: $1s^2 2s^2 2p^5$



- 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.

Advanced Theories of Atomic Structure

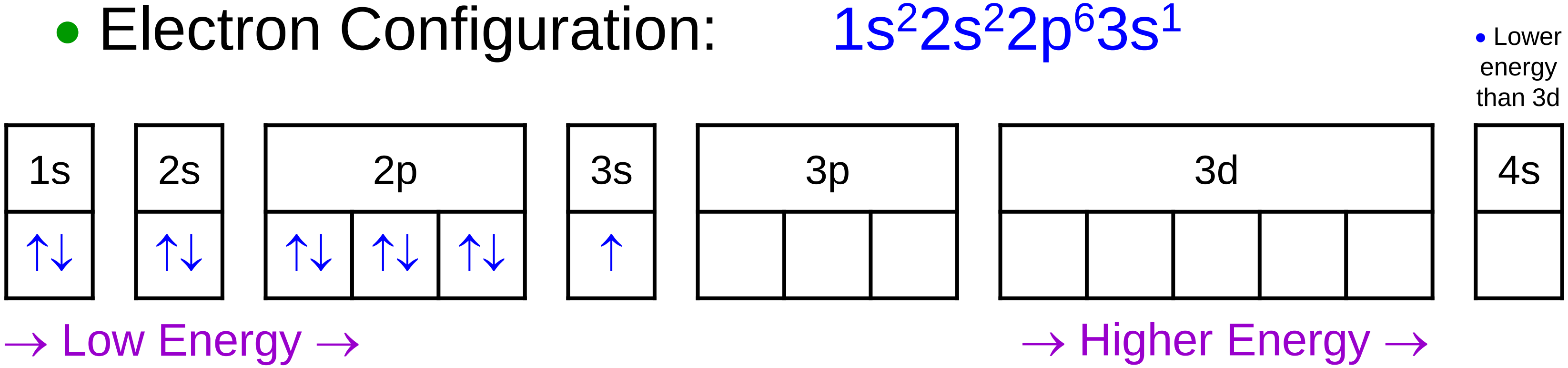
- Atomic Number: 10
- Name: Neon
- Symbol: Ne
- Electron Configuration: $1s^2 2s^2 2p^6$



- 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.

Advanced Theories of Atomic Structure

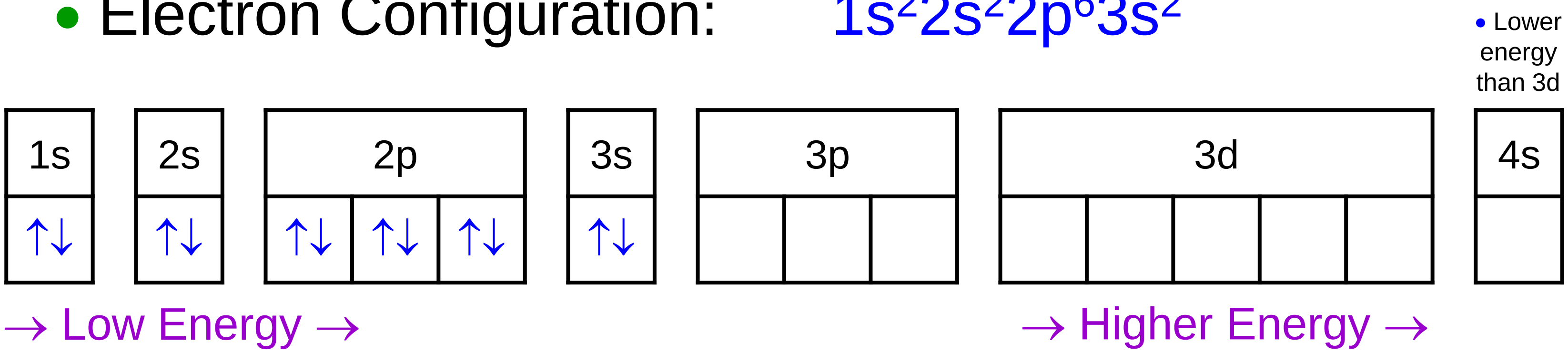
- Atomic Number: 11
- Name: Sodium
- Symbol: Na
- Electron Configuration: $1s^2 2s^2 2p^6 3s^1$



- 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.

Advanced Theories of Atomic Structure

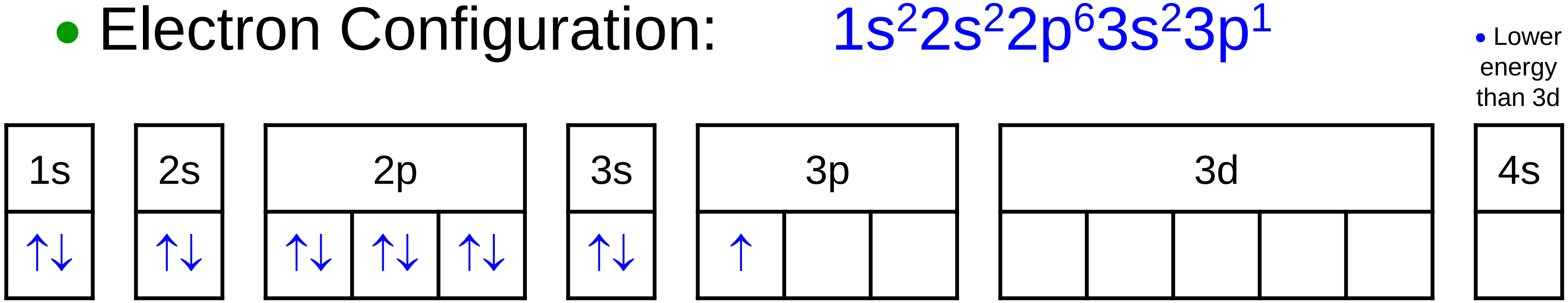
- Atomic Number: 12
- Name: Magnesium
- Symbol: Mg
- Electron Configuration: $1s^2 2s^2 2p^6 3s^2$



- 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.

Advanced Theories of Atomic Structure

- Atomic Number: 13
- Name: Aluminium
- Symbol: A*l*
- Electron Configuration: $1s^2 2s^2 2p^6 3s^2 3p^1$

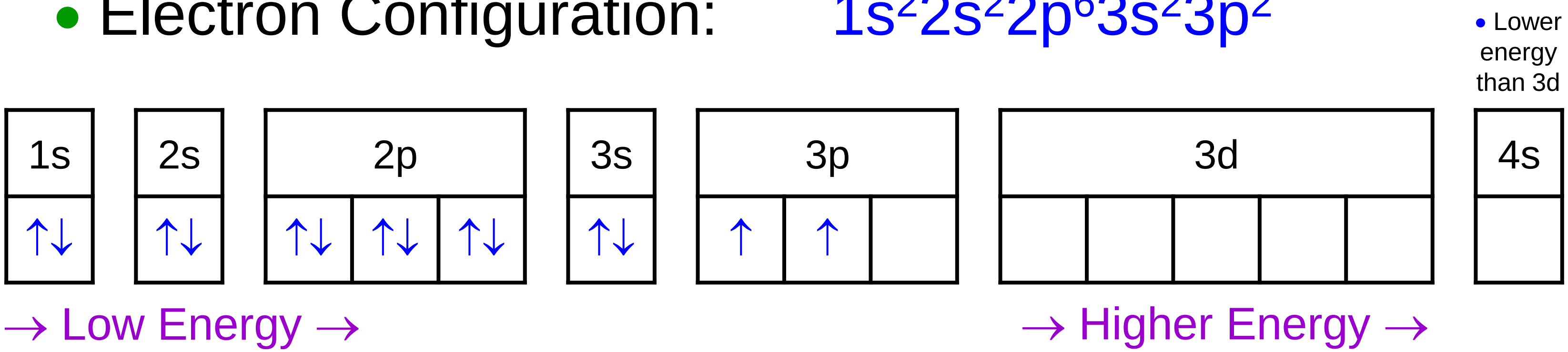


→ Low Energy → → Higher Energy →

- 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.

Advanced Theories of Atomic Structure

- Atomic Number: 14
- Name: Silicon
- Symbol: Si
- Electron Configuration: $1s^2 2s^2 2p^6 3s^2 3p^2$



- 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.

Advanced Theories of Atomic Structure

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

• Lower energy than 3d

1s	2s	2p			3s	3p			3d					4s
↑↓	↑↓	↑↓	↑↓	↑↓	↑↓	↑	↑	↑						

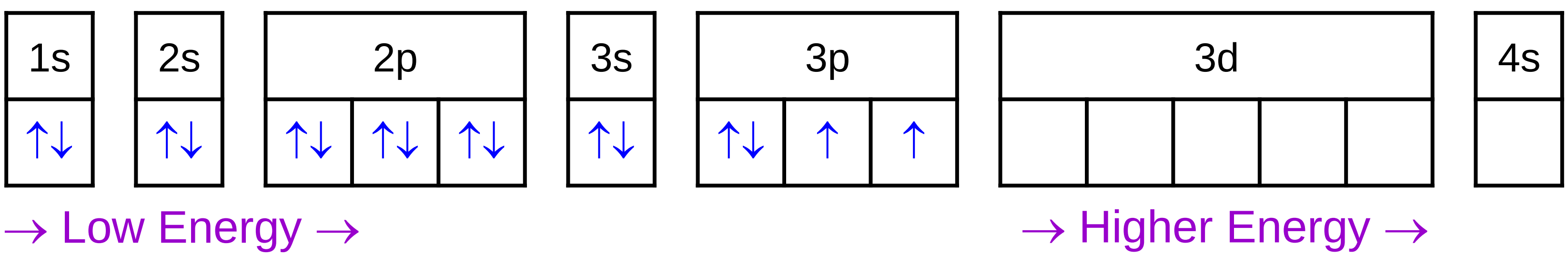
→ Low Energy → → Higher Energy →

- 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.

Advanced Theories of Atomic Structure

- Atomic Number: 16
- Name: Sulfur
- Symbol: S
- Electron Configuration: $1s^2 2s^2 2p^6 3s^2 3p^4$

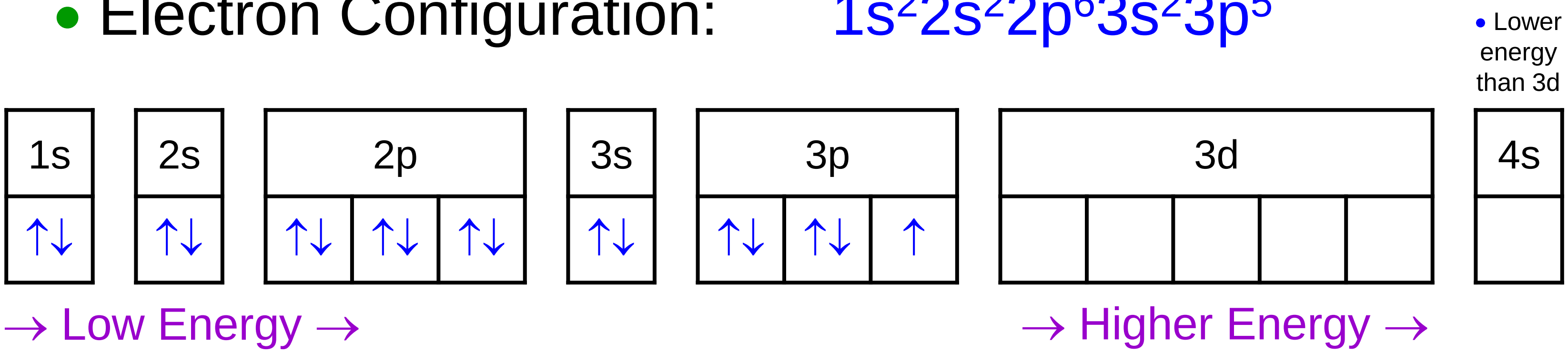
- Lower energy than 3d



- 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.

Advanced Theories of Atomic Structure

- Atomic Number: 17
- Name: Chlorine
- Symbol: C*l*
- Electron Configuration: $1s^2 2s^2 2p^6 3s^2 3p^5$

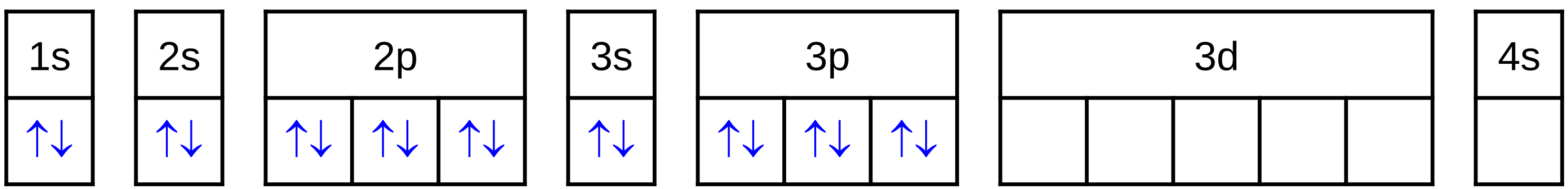


- Electrons are represented by *arrows* (↑ and ↓) which fill atomic orbitals that are represented by *boxes*.
- Electrons occupy atomic orbitals from the lowest energy to the highest energy. This is known as the *Aufbau Principle*.
- A single electron will occupy a single atomic orbital before two electrons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

Advanced Theories of Atomic Structure

- Atomic Number: 18
- Name: Argon
- Symbol: Ar
- Electron Configuration: $1s^2 2s^2 2p^6 3s^2 3p^6$

- Lower energy than 3d

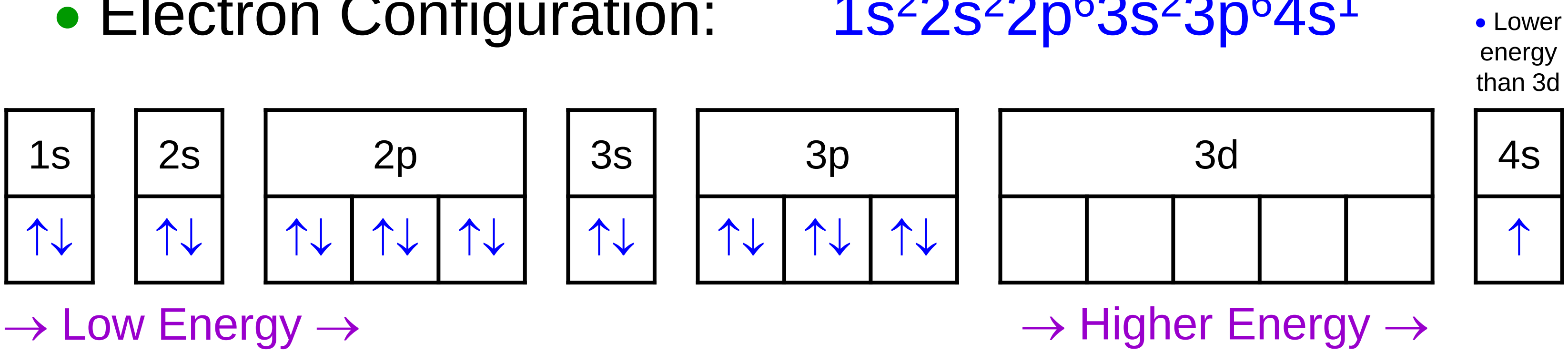


→ Low Energy → → Higher Energy →

- 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.

Advanced Theories of Atomic Structure

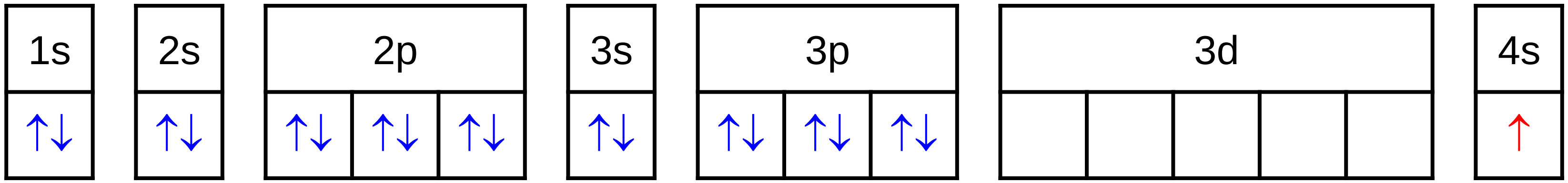
- Atomic Number: 19
- Name: Potassium
- Symbol: K
- Electron Configuration: $1s^2 2s^2 2p^6 3s^2 3p^6 4s^1$



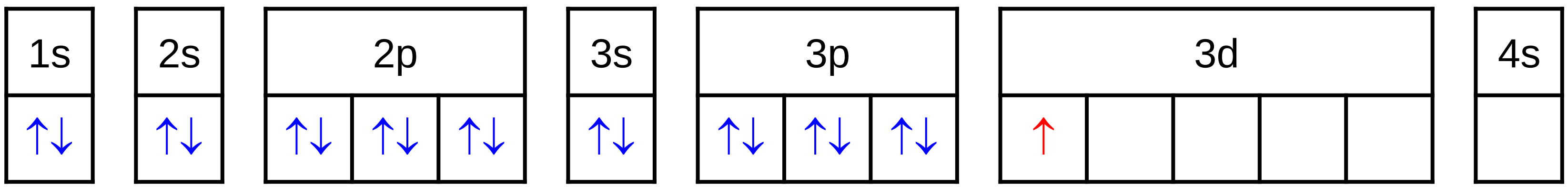
- Electrons are represented by *arrows* (↑ and ↓) which fill atomic orbitals that are represented by *boxes*.
- Electrons occupy atomic orbitals from the lowest energy to the highest energy. This is known as the *Aufbau Principle*.
- A single electron will occupy a single atomic orbital before two electrons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

Advanced Theories of Atomic Structure

- Why is the electron configuration of *potassium*…



…instead of…

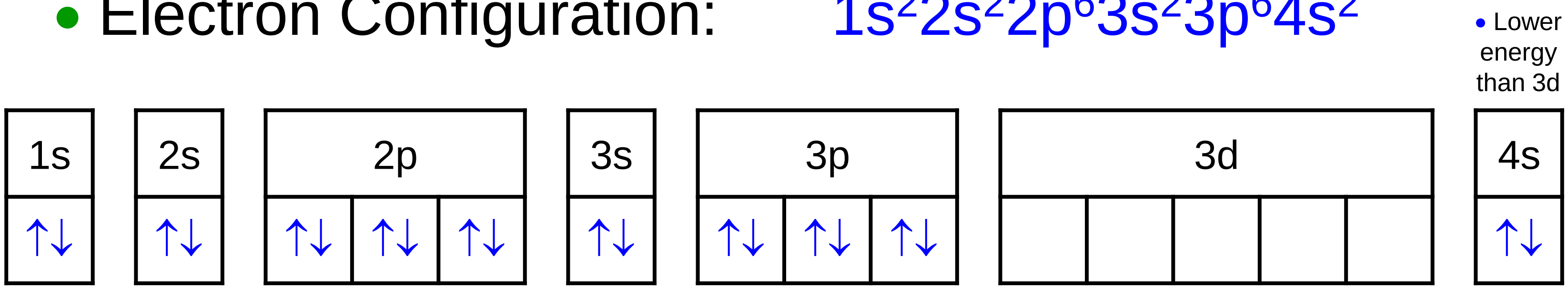


- Although the 4s sub-shell is further from the nucleus than the 3d sub-shell, the 4s sub-shell is *lower in energy* than the 3d sub-shell.

- 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: 20
- Name: Calcium
- Symbol: Ca
- Electron Configuration: $1s^2 2s^2 2p^6 3s^2 3p^6 4s^2$



→ Low Energy → → Higher Energy →

- 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.

Advanced Theories of Atomic Structure

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

1s	2s	2p			3s	3p			3d					4s
↑↓	↑↓	↑↓	↑↓	↑↓	↑↓	↑↓	↑↓	↑↓	↑					↑↓

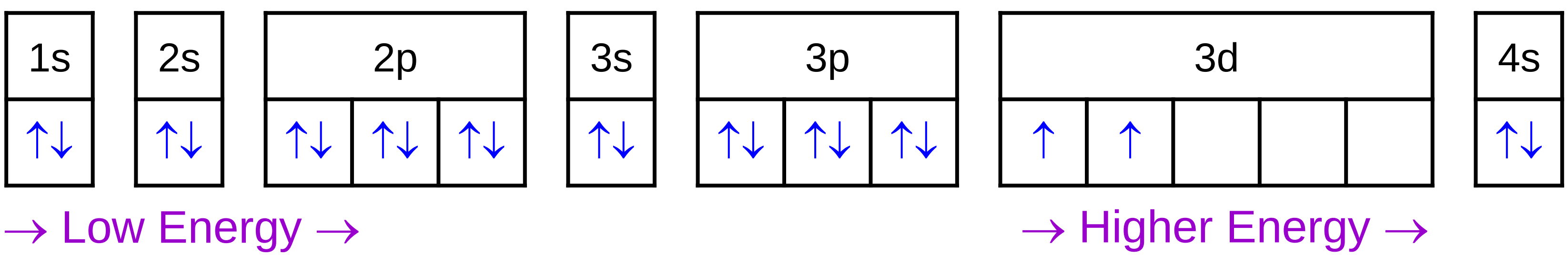
→ Low Energy →   → Higher Energy →

- 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.

Advanced Theories of Atomic Structure

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

  - Lower energy than 3d

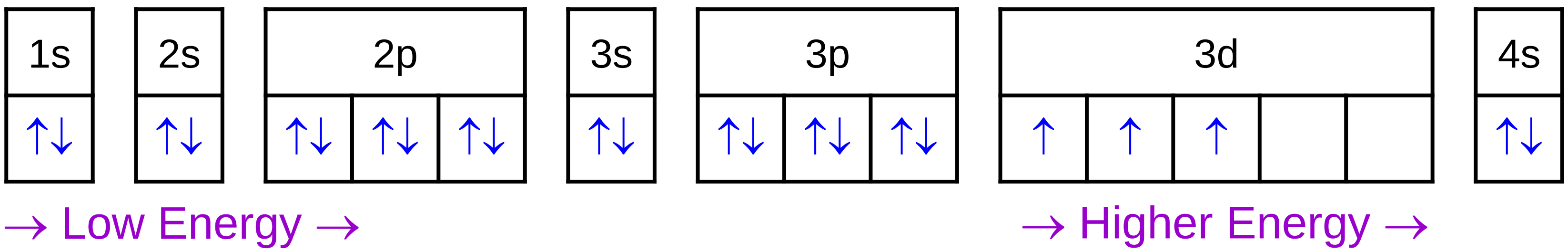


- Electrons are represented by *arrows* (↑ and ↓) which fill atomic orbitals that are represented by *boxes*.
- Electrons occupy atomic orbitals from the lowest energy to the highest energy. This is known as the *Aufbau Principle*.
- A single electron will occupy a single atomic orbital before two electrons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

Advanced Theories of Atomic Structure

- Atomic Number: 23
- Name: Vanadium
- Symbol: V
- Electron 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.

Advanced Theories of Atomic Structure

- Atomic Number: 24
- Name: Chromium
- Symbol: Cr
- Electron Configuration: $1s^22s^22p^63s^23p^63d^54s^1$

• Lower energy than 3d

1s	2s	2p			3s	3p			3d					4s
↑↓	↑↓	↑↓	↑↓	↑↓	↑↓	↑↓	↑↓	↑↓	↑	↑	↑	↑	↑	↑

→ Low Energy → → Higher Energy →

- 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.

Advanced Theories of Atomic Structure

- Why is the electron configuration of *chromium*…

1s	2s	2p			3s	3p			3d					4s
↑↓	↑↓	↑↓	↑↓	↑↓	↑↓	↑↓	↑↓	↑↓	↑	↑	↑	↑	↑	↑

…instead of…

1s	2s	2p			3s	3p			3d					4s
↑↓	↑↓	↑↓	↑↓	↑↓	↑↓	↑↓	↑↓	↑↓	↑	↑	↑	↑		↑↓

- 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
- Electron Configuration: $1s^2 2s^2 2p^6 3s^2 3p^6 3d^5 4s^2$

• Lower energy than 3d

1s	2s	2p	3s	3p	3d	4s
↑↓	↑↓	↑↓ ↑↓ ↑↓	↑↓	↑↓ ↑↓ ↑↓	↑ ↑ ↑ ↑ ↑	↑↓

→ Low Energy → → Higher Energy →

- 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.

Advanced Theories of Atomic Structure

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

1s	2s	2p			3s	3p			3d					4s
↑↓	↑↓	↑↓	↑↓	↑↓	↑↓	↑↓	↑↓	↑↓	↑↓	↑	↑	↑	↑	↑↓

→ Low Energy → → Higher Energy →

- 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.

Advanced Theories of Atomic Structure

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

• Lower energy than 3d

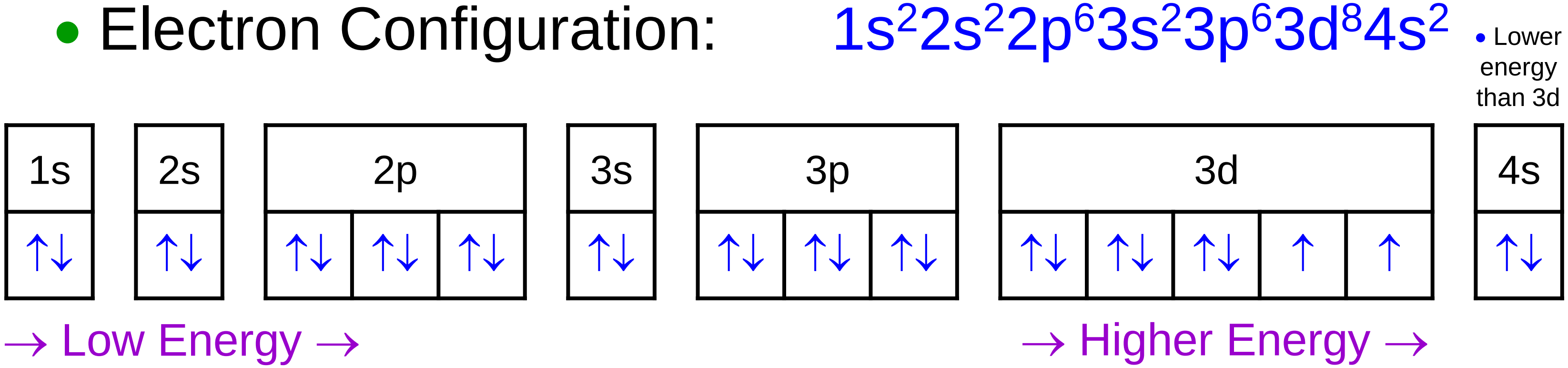
1s	2s	2p			3s	3p			3d					4s
↑↓	↑↓	↑↓	↑↓	↑↓	↑↓	↑↓	↑↓	↑↓	↑↓	↑↓	↑	↑	↑	↑↓

→ Low Energy → → Higher Energy →

- 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.

Advanced Theories of Atomic Structure

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



- 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.

Advanced Theories of Atomic Structure

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

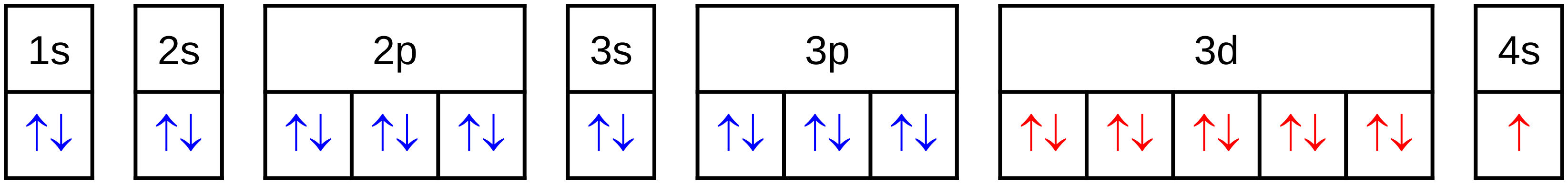
1s	2s	2p			3s	3p			3d					4s
↑↓	↑↓	↑↓	↑↓	↑↓	↑↓	↑↓	↑↓	↑↓	↑↓	↑↓	↑↓	↑↓	↑↓	↑

→ Low Energy → → Higher Energy →

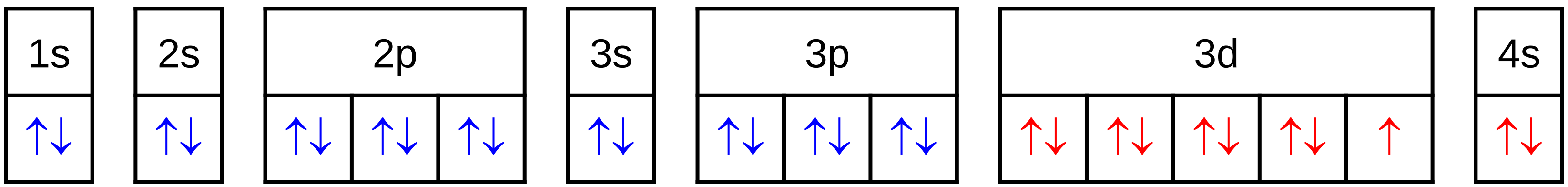
- Electrons are represented by *arrows* (↑ and ↓) which fill atomic orbitals that are represented by *boxes*.
- Electrons occupy atomic orbitals from the lowest energy to the highest energy. This is known as the *Aufbau Principle*.
- A single electron will occupy a single atomic orbital before two electrons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

Advanced Theories of Atomic Structure

- Why is the electron configuration of *copper*…



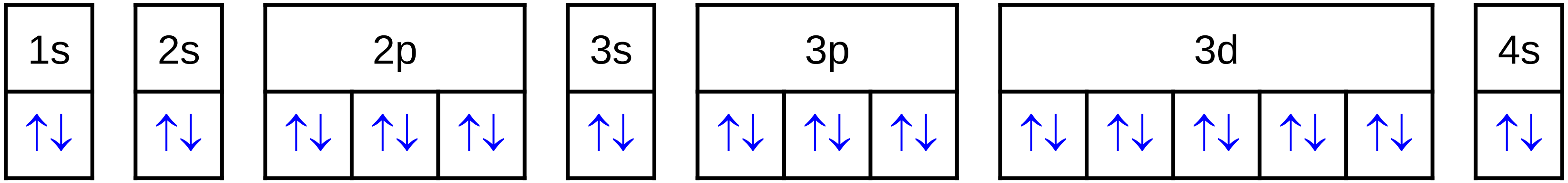
…instead of…



- 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 ($4s^{2}$).

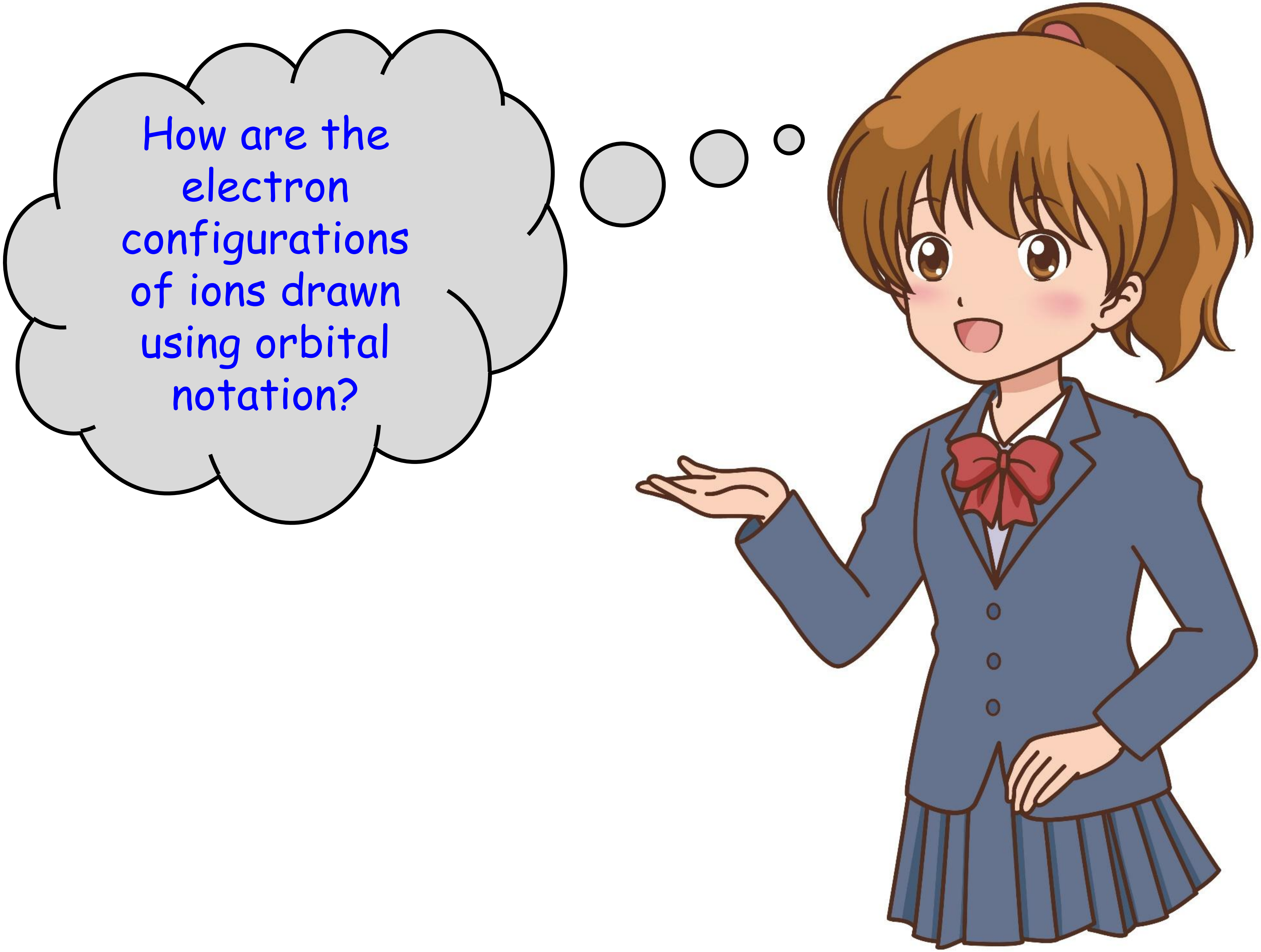
Advanced Theories of Atomic Structure

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



→ Low Energy → → Higher Energy →

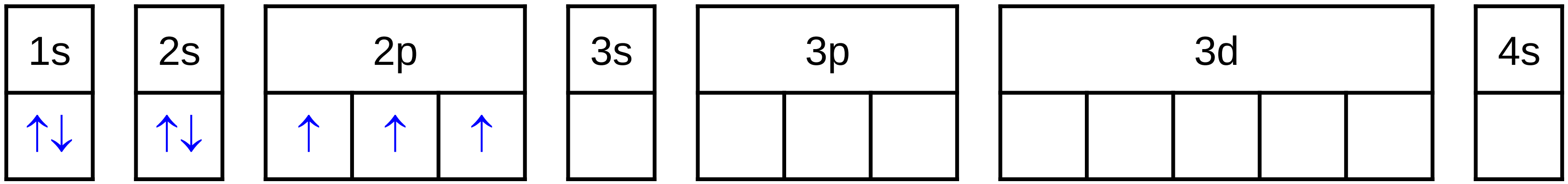
- Electrons are represented by *arrows* (↑ and ↓) which fill atomic orbitals that are represented by *boxes*.
- Electrons occupy atomic orbitals from the lowest energy to the highest energy. This is known as the *Aufbau Principle*.
- A single electron will occupy a single atomic orbital before two electrons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


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

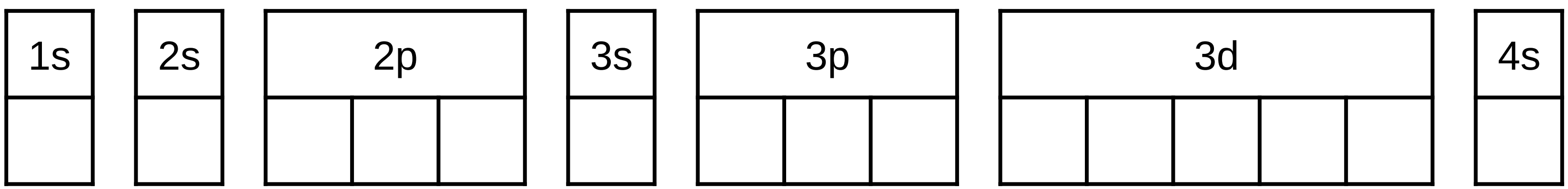
Advanced Theories of Atomic Structure

Electron Configurations of Ions

Electron Configuration of a Nitrogen Atom



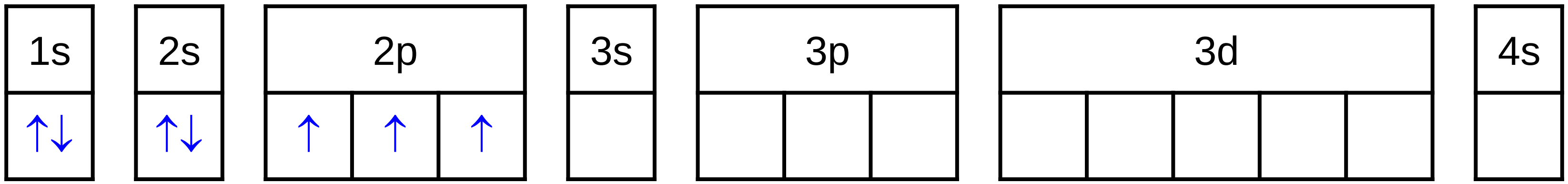
Electron Configuration of a Nitride Ion – N^{3-}



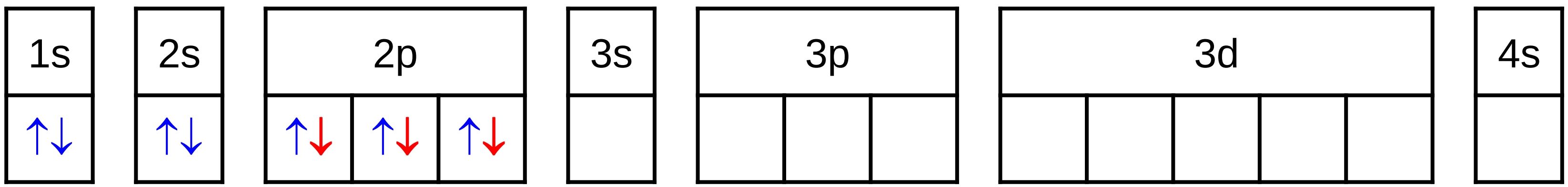
Advanced Theories of Atomic Structure

Electron Configurations of Ions

Electron Configuration of a Nitrogen Atom



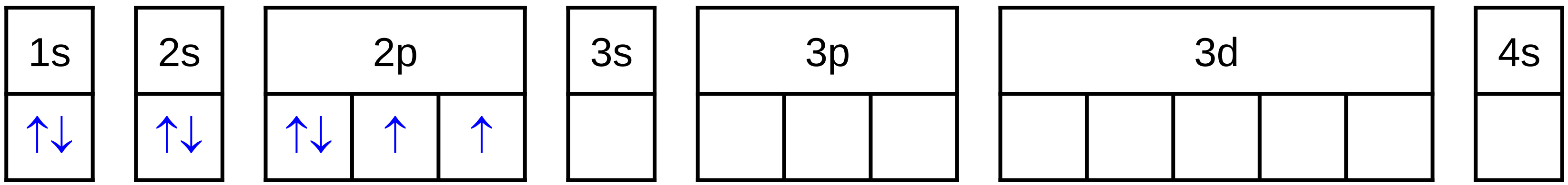
Electron Configuration of a Nitride Ion – N^{3-}



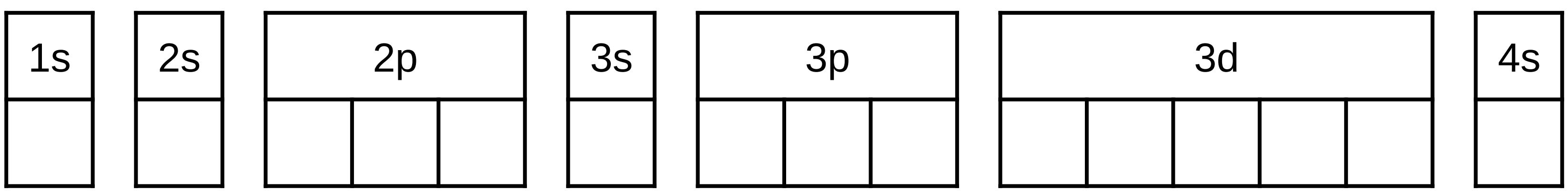
Advanced Theories of Atomic Structure

Electron Configurations of Ions

Electron Configuration of an Oxygen Atom



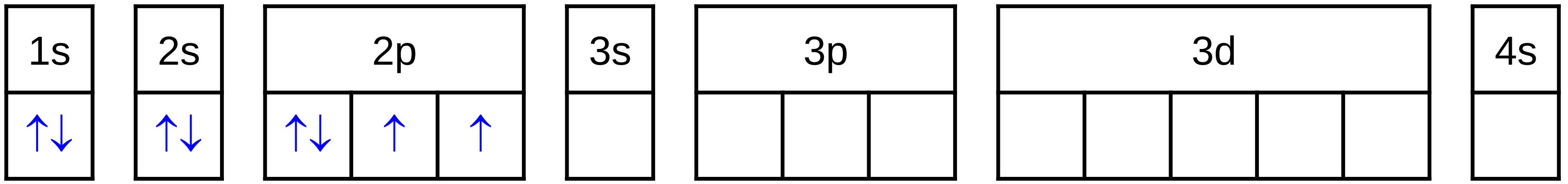
Electron Configuration of a Oxide Ion – O^{2-}



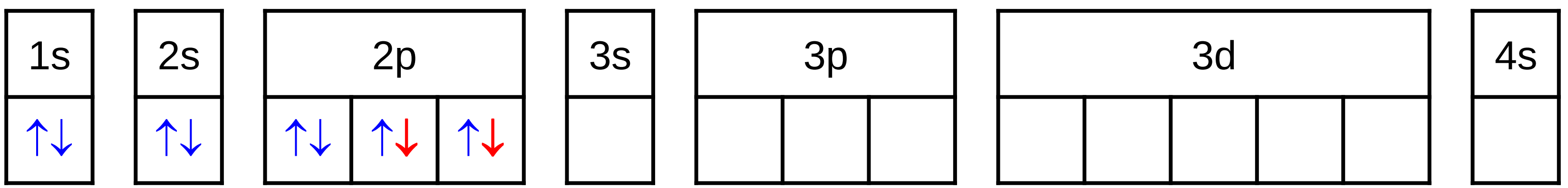
Advanced Theories of Atomic Structure

Electron Configurations of Ions

Electron Configuration of an Oxygen Atom



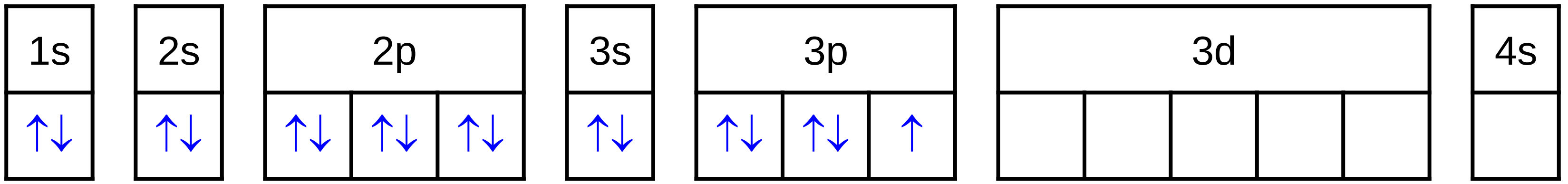
Electron Configuration of a Oxide Ion – O^{2-}



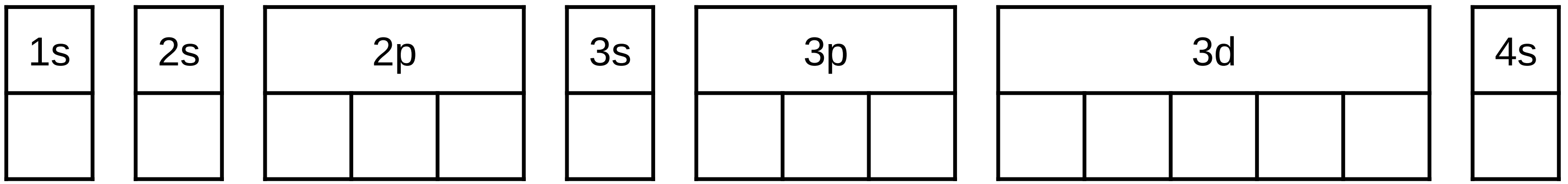
Advanced Theories of Atomic Structure

Electron Configurations of Ions

Electron Configuration of a Chlorine Atom



Electron Configuration of a Chloride Ion – Cl^{-}



Advanced Theories of Atomic Structure

Electron Configurations of Ions

Electron Configuration of a Chlorine Atom

1s	2s	2p			3s	3p			3d					4s
↑↓	↑↓	↑↓	↑↓	↑↓	↑↓	↑↓	↑↓	↑						

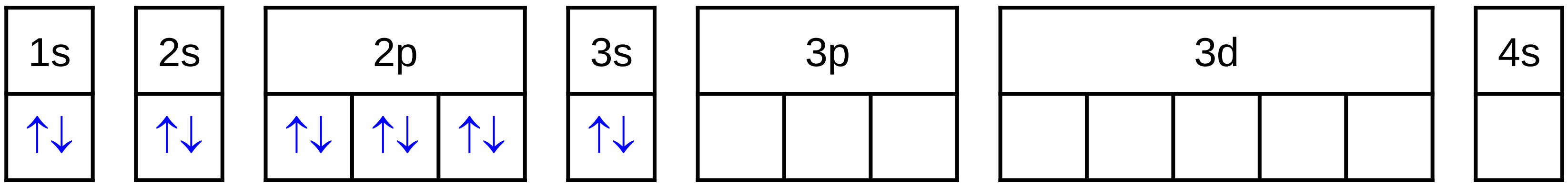
Electron Configuration of a Chloride Ion – Cl^-

1s	2s	2p			3s	3p			3d					4s
↑↓	↑↓	↑↓	↑↓	↑↓	↑↓	↑↓	↑↓	↑↓						

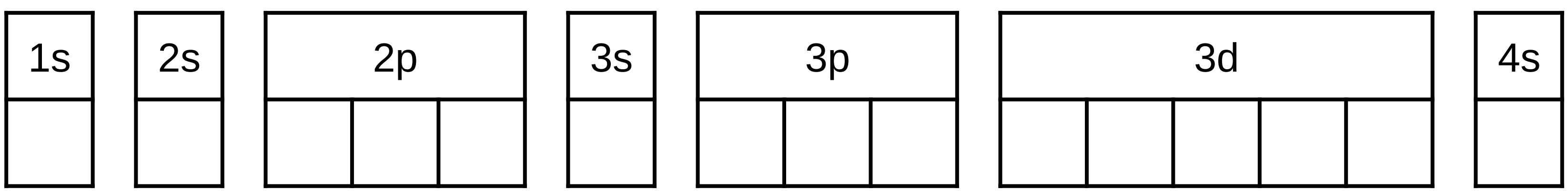
Advanced Theories of Atomic Structure

Electron Configurations of Ions

Electron Configuration of a Magnesium Atom



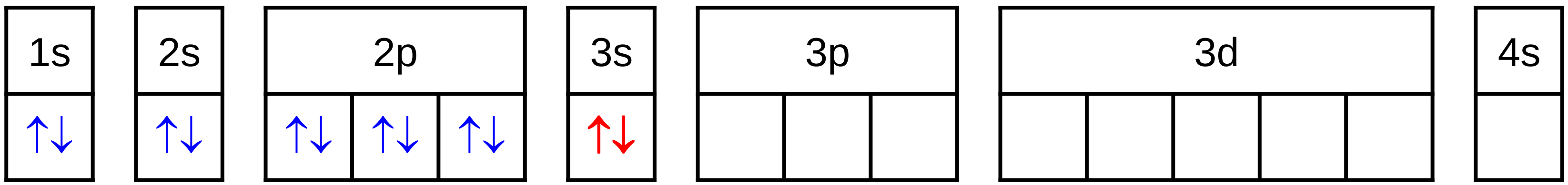
Electron Configuration of a Magnesium Ion – Mg^{2+}



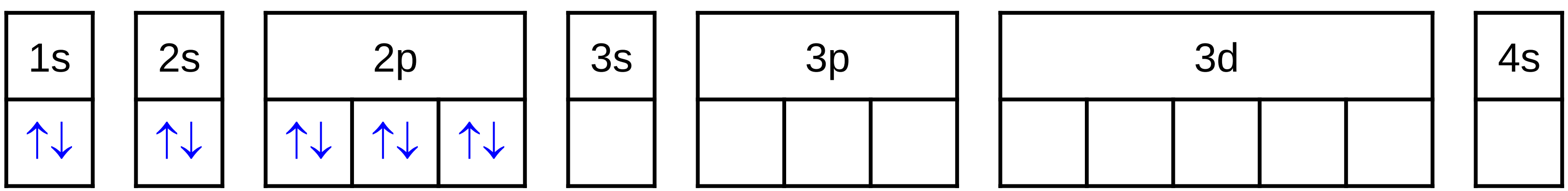
Advanced Theories of Atomic Structure

Electron Configurations of Ions

Electron Configuration of a Magnesium Atom



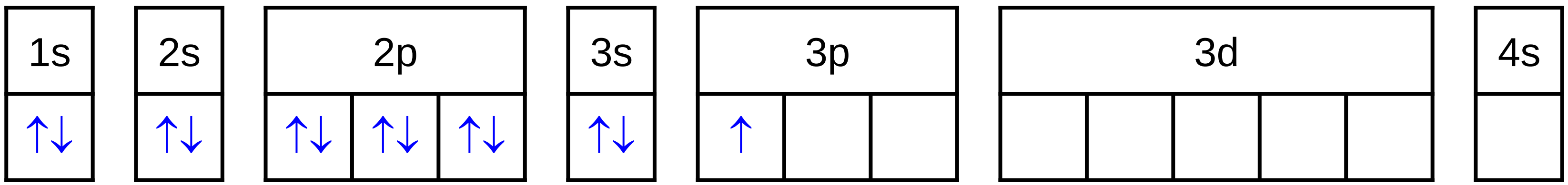
Electron Configuration of a Magnesium Ion – Mg^{2+}



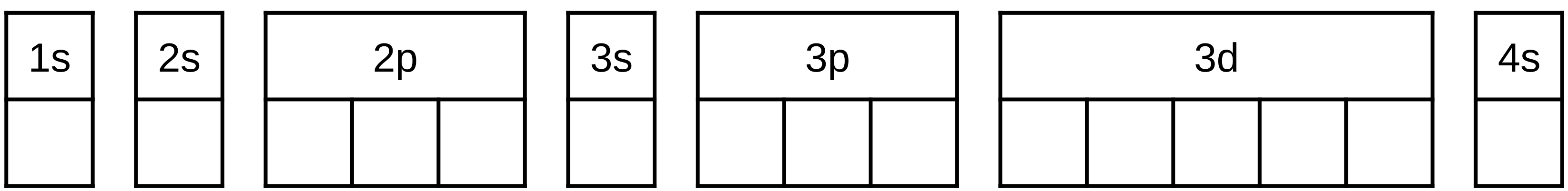
Advanced Theories of Atomic Structure

Electron Configurations of Ions

Electron Configuration of an Aluminium Atom



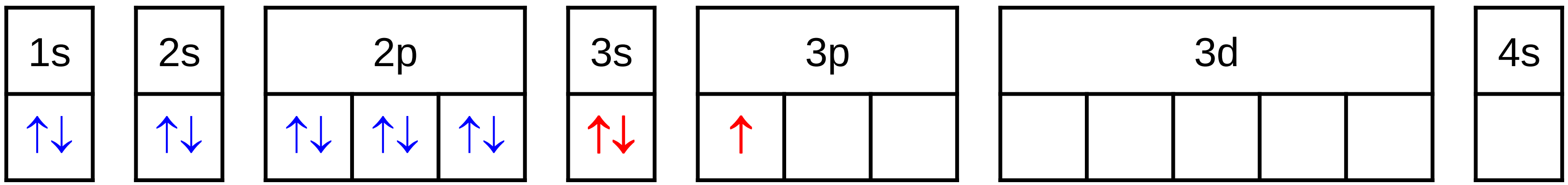
Electron Configuration of a Aluminium Ion – Al^{3+}



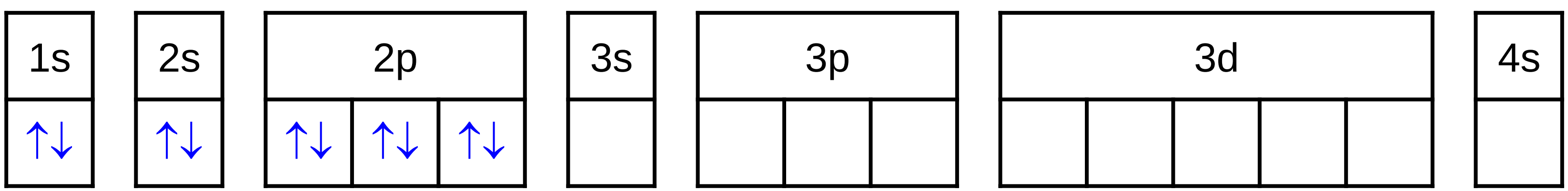
Advanced Theories of Atomic Structure

Electron Configurations of Ions

Electron Configuration of an Aluminium Atom



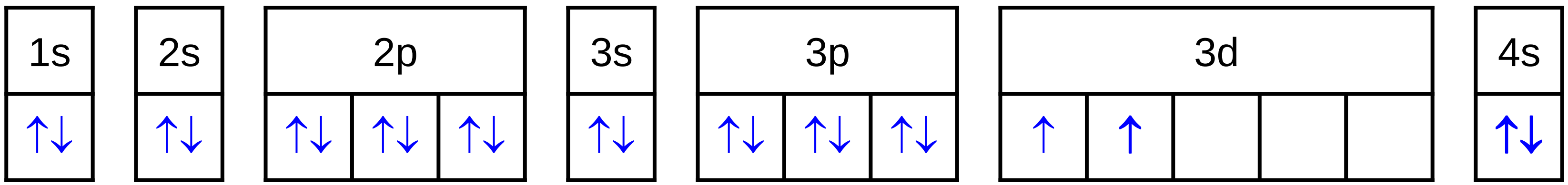
Electron Configuration of a Aluminium Ion – Al^{3+}



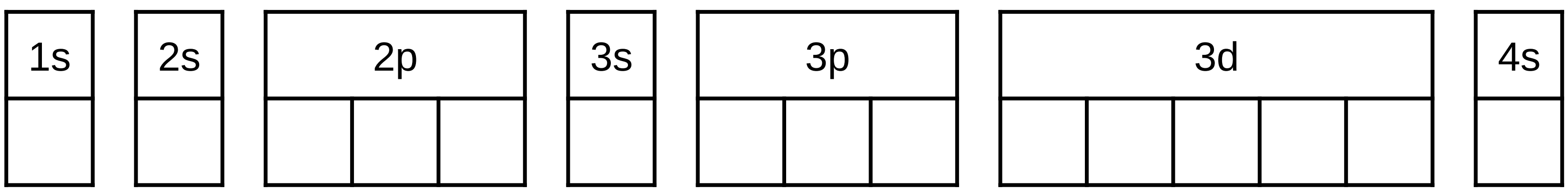
Advanced Theories of Atomic Structure

Electron Configurations of Ions

Electron Configuration of a Titanium Atom



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



Advanced Theories of Atomic Structure

Electron Configurations of Ions

Electron Configuration of a Titanium Atom

1s	2s	2p			3s	3p			3d					4s
↑↓	↑↓	↑↓	↑↓	↑↓	↑↓	↑↓	↑↓	↑↓	↑	↑				↑↓

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

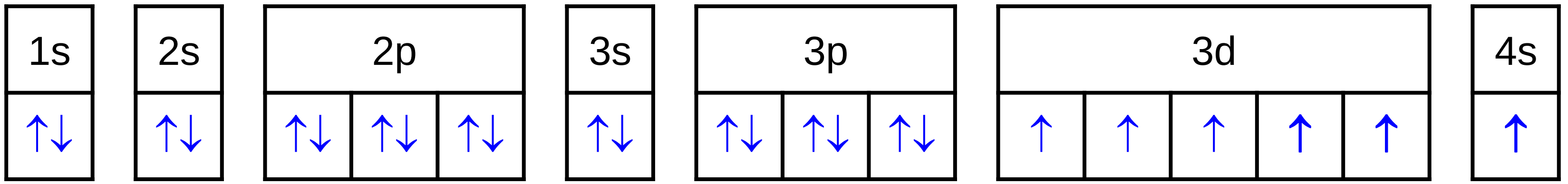
1s	2s	2p			3s	3p			3d					4s
↑↓	↑↓	↑↓	↑↓	↑↓	↑↓	↑↓	↑↓	↑↓	↑					

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

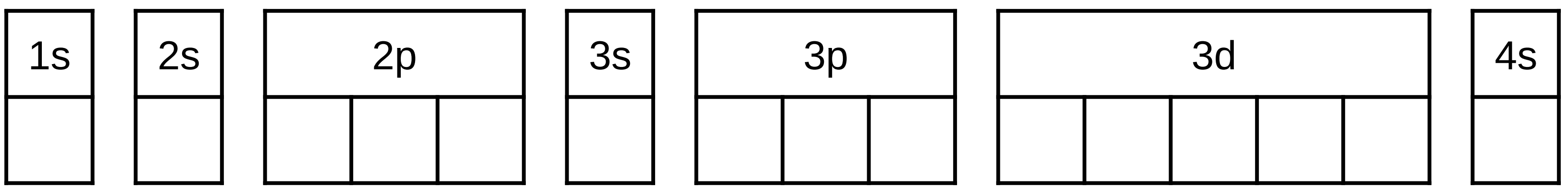
Advanced Theories of Atomic Structure

Electron Configurations of Ions

Electron Configuration of a Chromium Atom



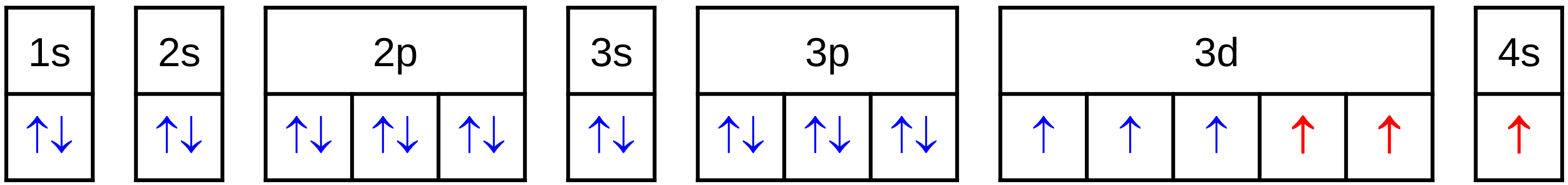
Electron Configuration of a Chromium(III) Ion – Cr^{3+}



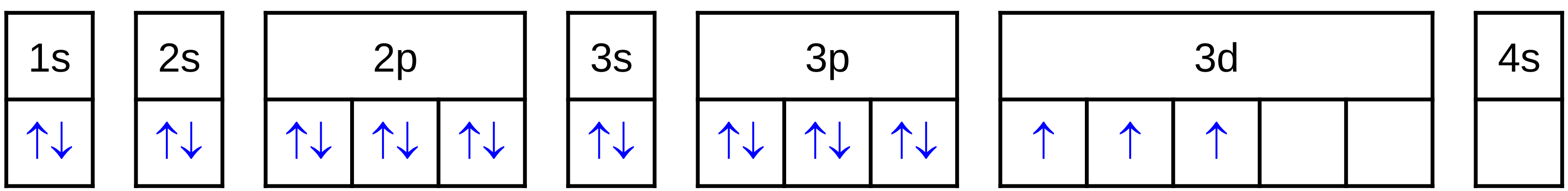
Advanced Theories of Atomic Structure

Electron Configurations of Ions

Electron Configuration of a Chromium Atom



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

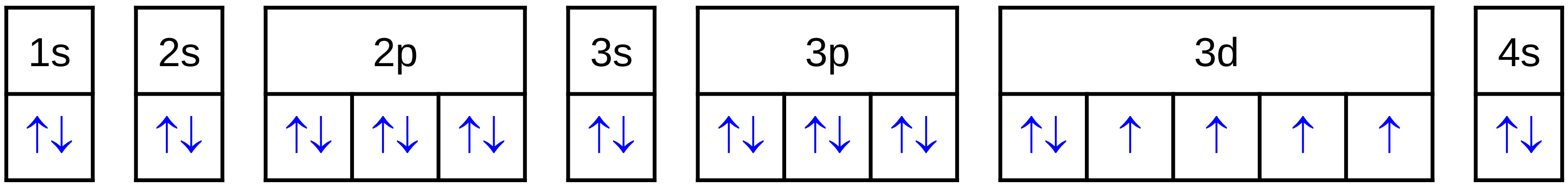


Note: When atoms of the transition metals react to form ions, electrons from the 4s orbital are the first to be removed.

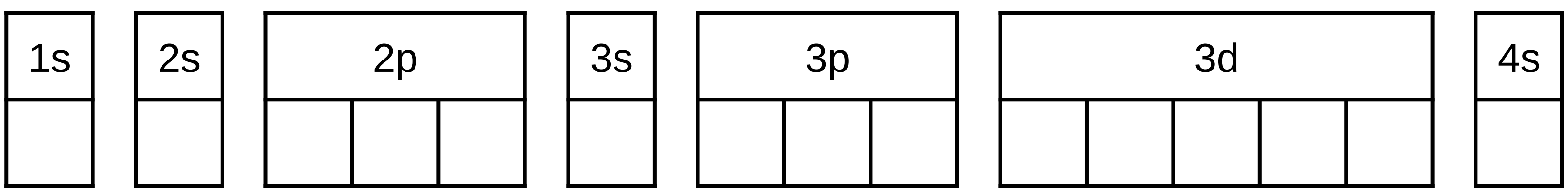
Advanced Theories of Atomic Structure

Electron Configurations of Ions

Electron Configuration of an Iron Atom



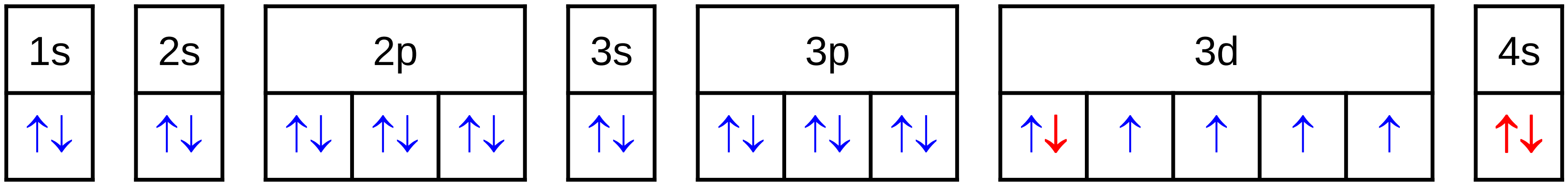
Electron Configuration of a Iron(III) Ion – Fe^{3+}



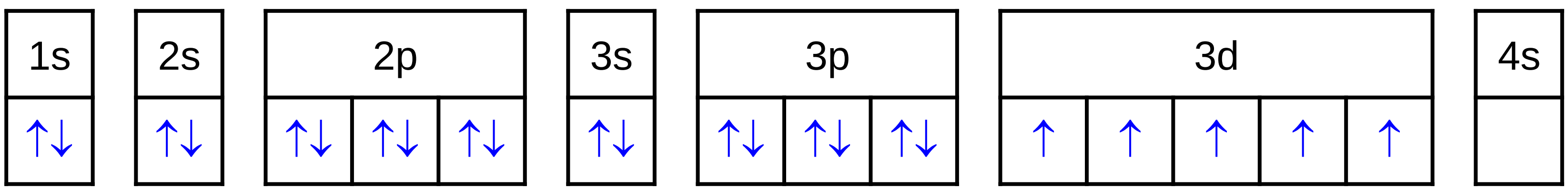
Advanced Theories of Atomic Structure

Electron Configurations of Ions

Electron Configuration of an Iron Atom



Electron Configuration of a Iron(III) Ion – Fe^{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

Electron Configurations of Ions

Electron Configuration of a Copper Atom

1s	2s	2p			3s	3p			3d					4s
↑↓	↑↓	↑↓	↑↓	↑↓	↑↓	↑↓	↑↓	↑↓	↑↓	↑↓	↑↓	↑↓	↑↓	↑

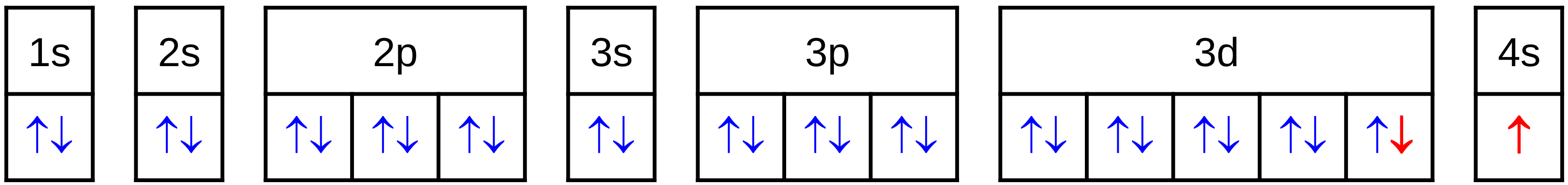
Electron Configuration of a Copper(II) Ion – Cu^{2+}

1s	2s	2p			3s	3p			3d					4s

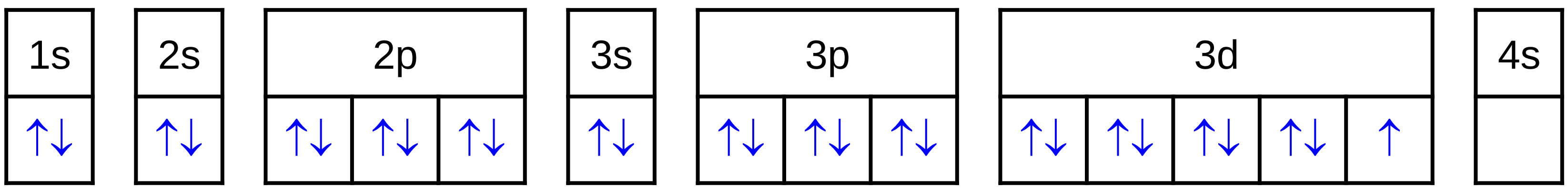
Advanced Theories of Atomic Structure

Electron Configurations of Ions

Electron Configuration of a Copper Atom



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

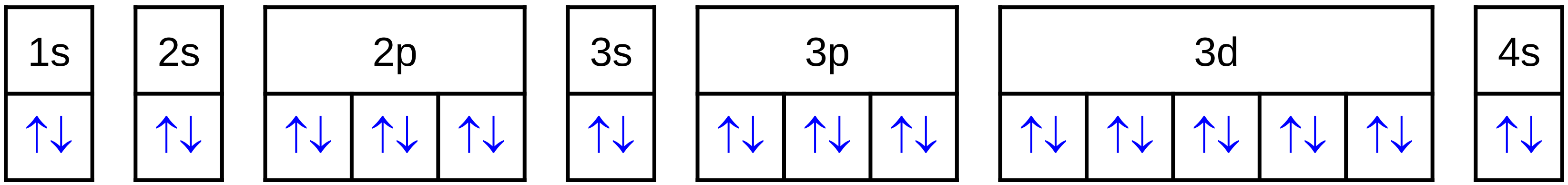


Note: When atoms of the transition metals react to form ions, electrons from the 4s orbital are the first to be removed.

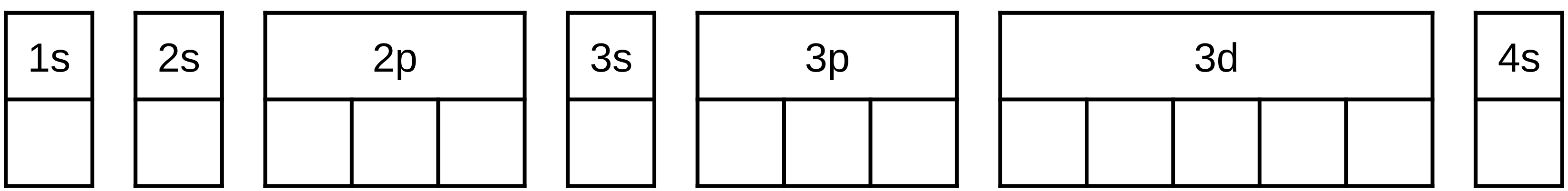
Advanced Theories of Atomic Structure

Electron Configurations of Ions

Electron Configuration of a Zinc Atom



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



Advanced Theories of Atomic Structure

Electron Configurations of Ions

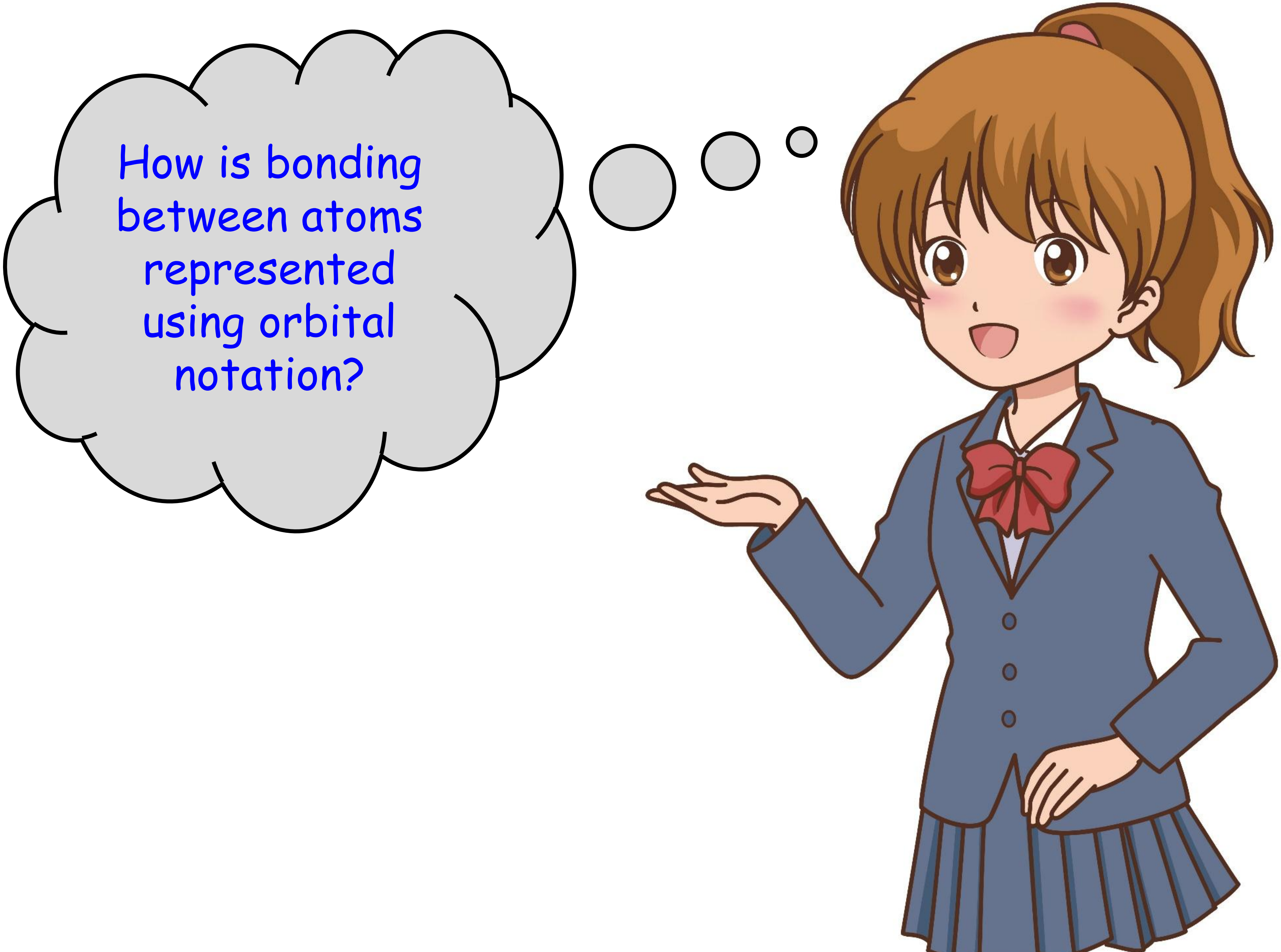
Electron Configuration of a Zinc Atom

1s	2s	2p			3s	3p			3d					4s
↑↓	↑↓	↑↓	↑↓	↑↓	↑↓	↑↓	↑↓	↑↓	↑↓	↑↓	↑↓	↑↓	↑↓	↑↓

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

1s	2s	2p			3s	3p			3d					4s
↑↓	↑↓	↑↓	↑↓	↑↓	↑↓	↑↓	↑↓	↑↓	↑↓	↑↓	↑↓	↑↓	↑↓	

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

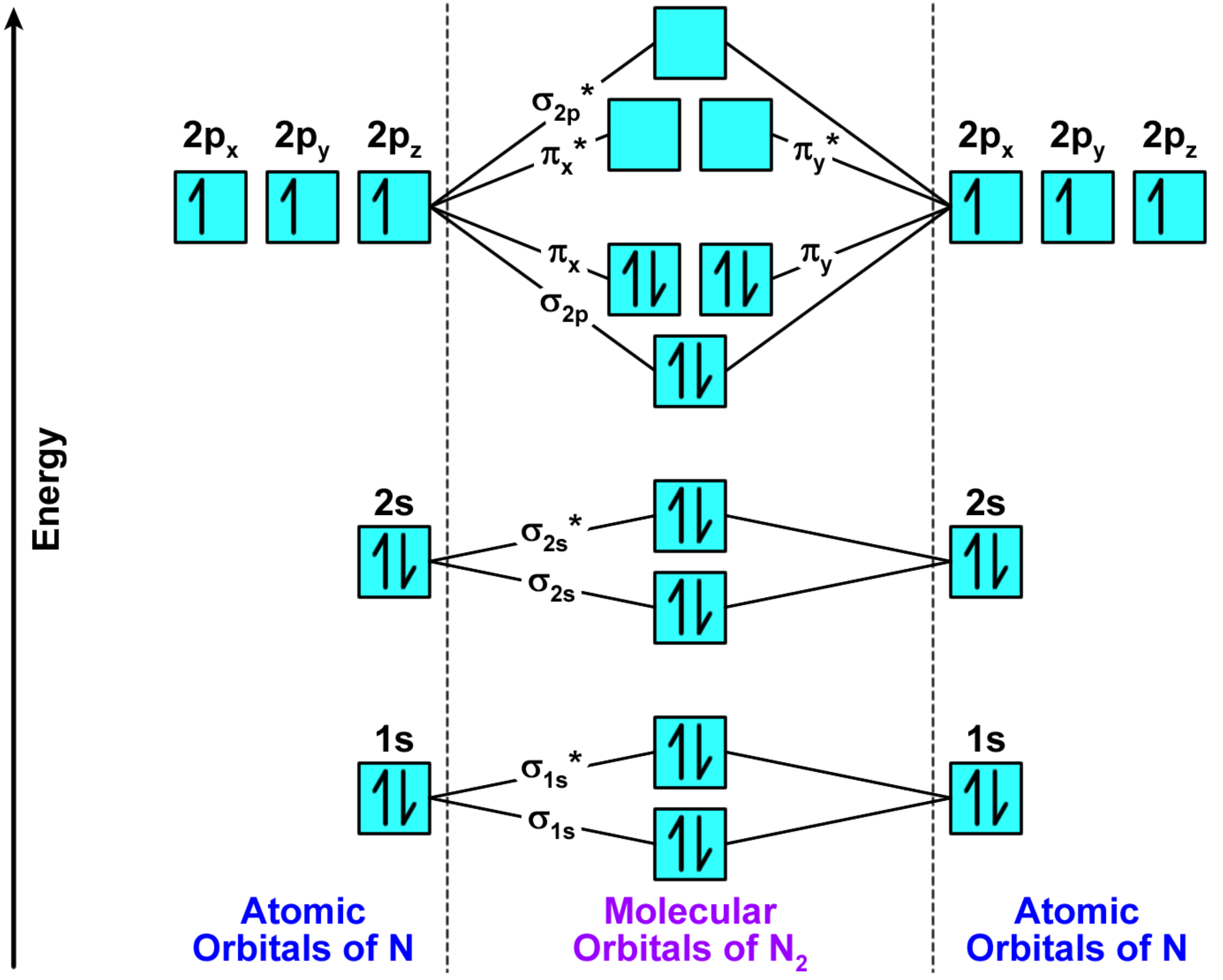
- The *atomic orbitals* of two or more atoms can combine together to form *molecular orbitals*.

- 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*.

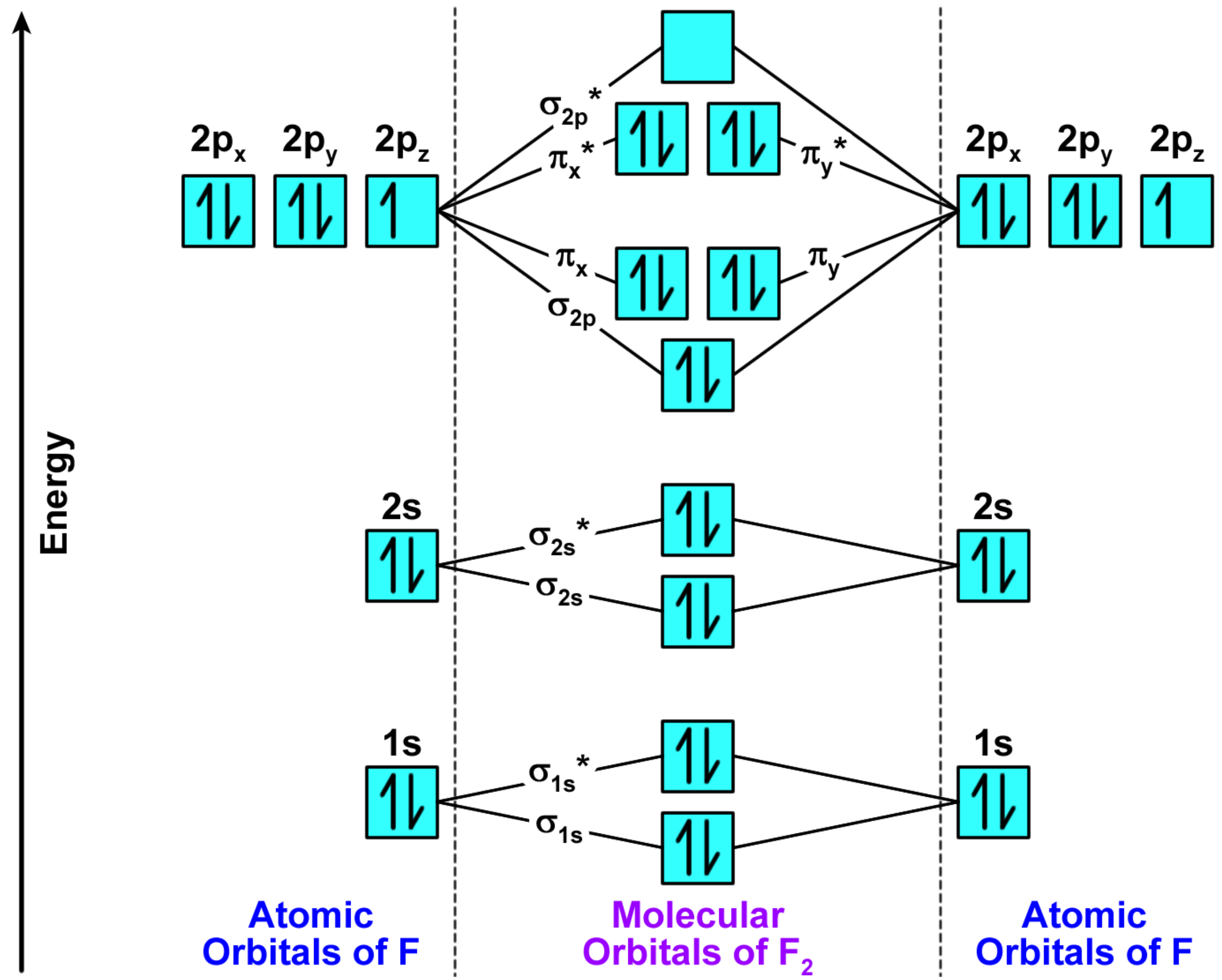
Advanced Theories of Atomic Structure

The Molecular Orbitals in Diatomic Nitrogen – N_2



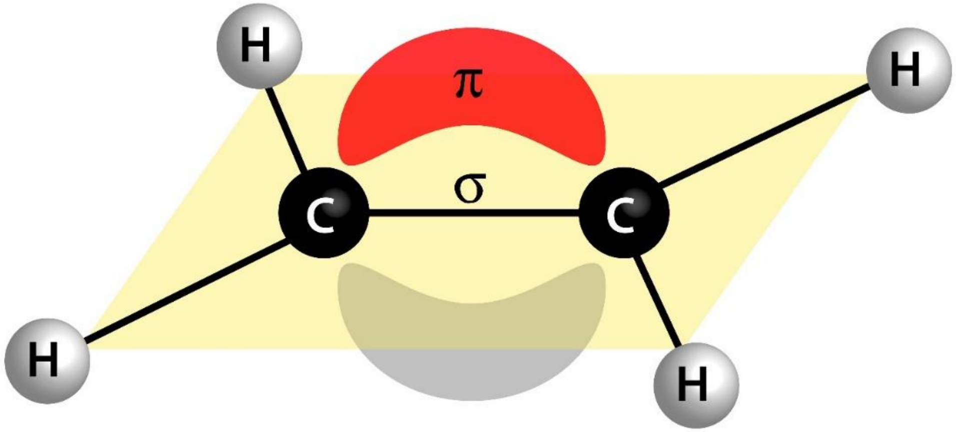
Advanced Theories of Atomic Structure

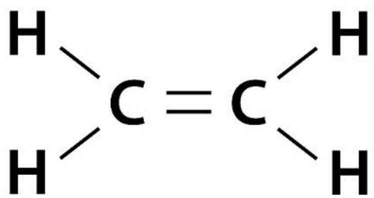
The Molecular Orbitals in Diatomic Fluorine – F_2



Advanced Theories of Atomic Structure

Pi-Bonds and Sigma-Bonds

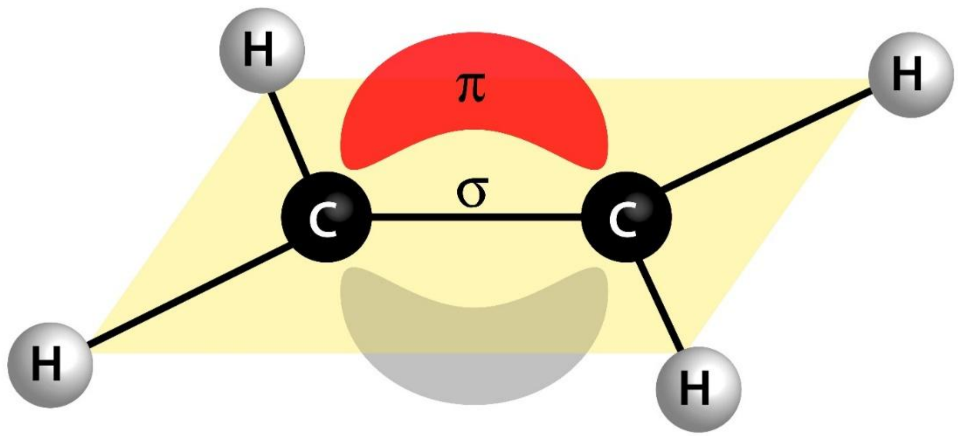


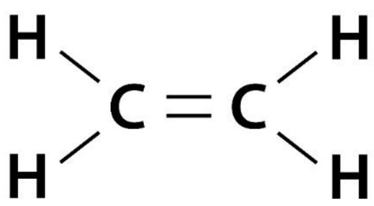


- Examples of *σ-bonds* and *π-bonds* in a molecule of ethene, C_2H_4.

Advanced Theories of Atomic Structure

Pi-Bonds and Sigma-Bonds

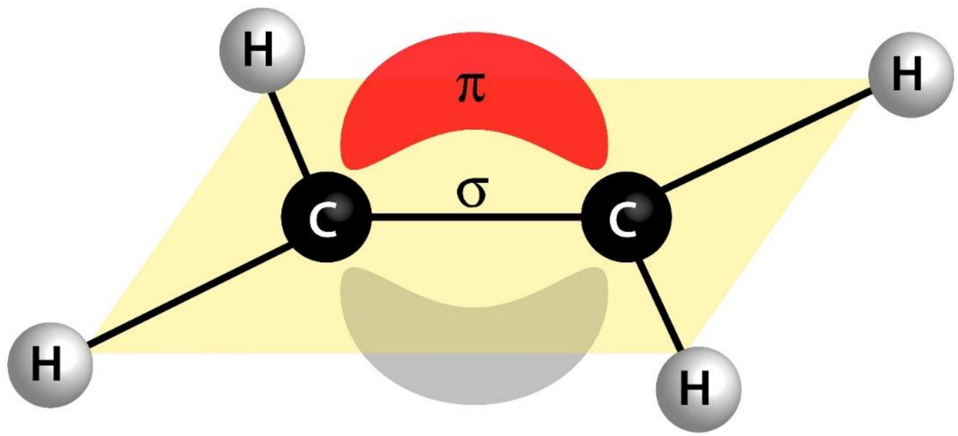


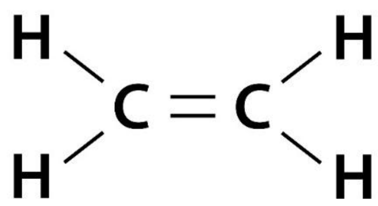


- A *σ-bond* is formed when two atomic orbitals overlap, and the region of overlap lies on an imaginary straight line that connects the nuclei of the two bonding atoms.

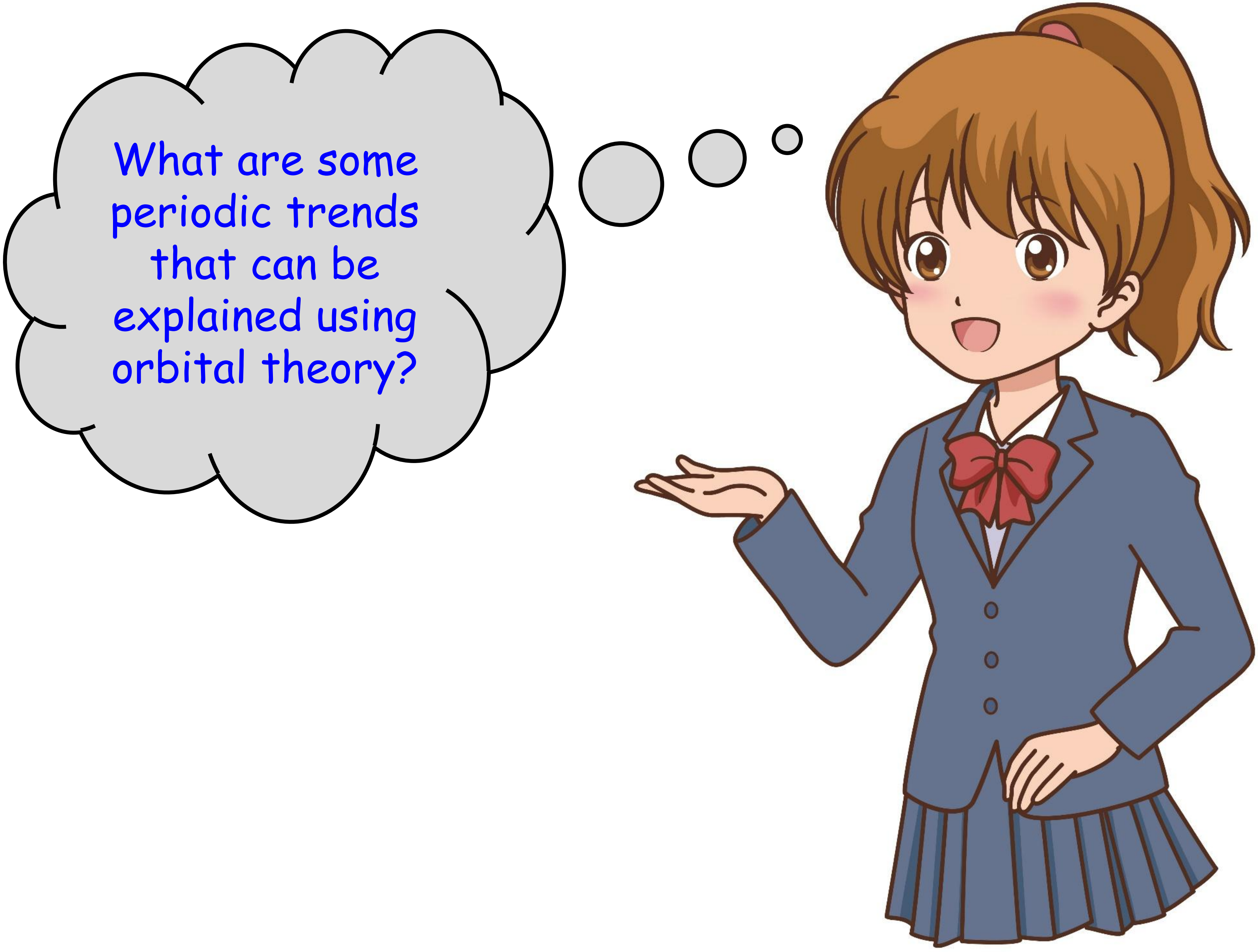
Advanced Theories of Atomic Structure

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:

$$F = \frac{1}{4 \times \pi \times \varepsilon_0} \times \frac{q_1 \times q_2}{r^2}$$

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

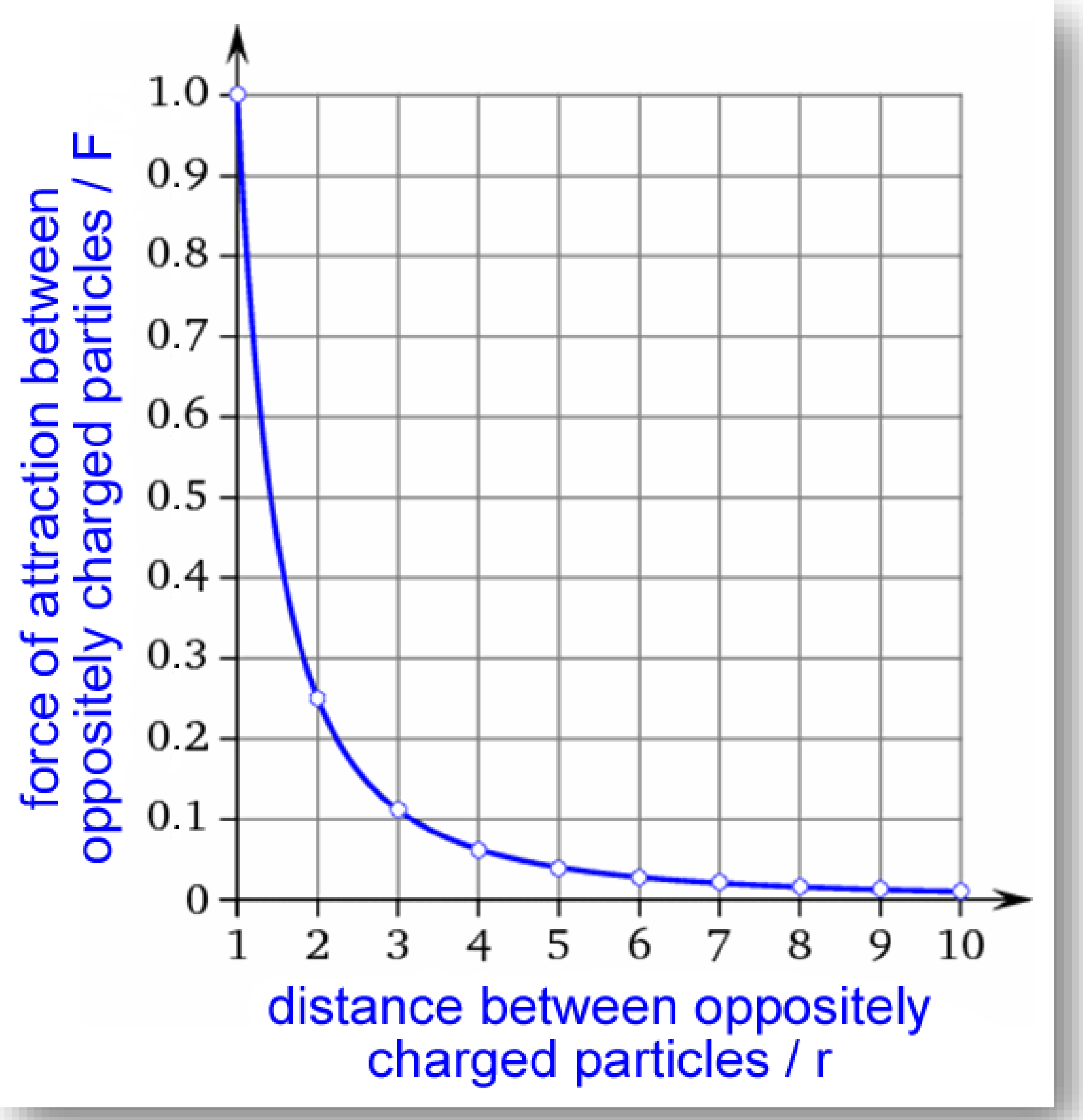
Periodic Trends

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

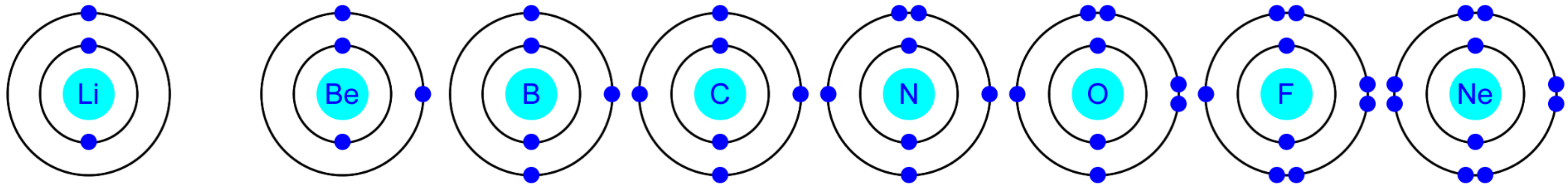
$$F = \frac{1}{4 \times \pi \times \varepsilon_0} \times \frac{q_1 \times q_2}{r^2}$$

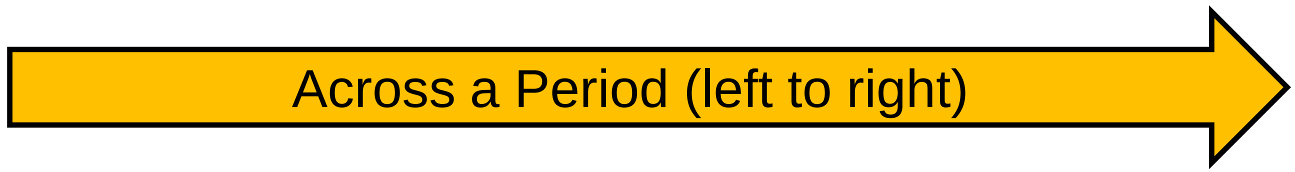
- The *force of attraction* (F) between a proton and an electron in an atom is related to their *charge* (q_1 and q_2), and it *decreases rapidly* as the *distance* between the particles (r) *increases* (inverse square law).

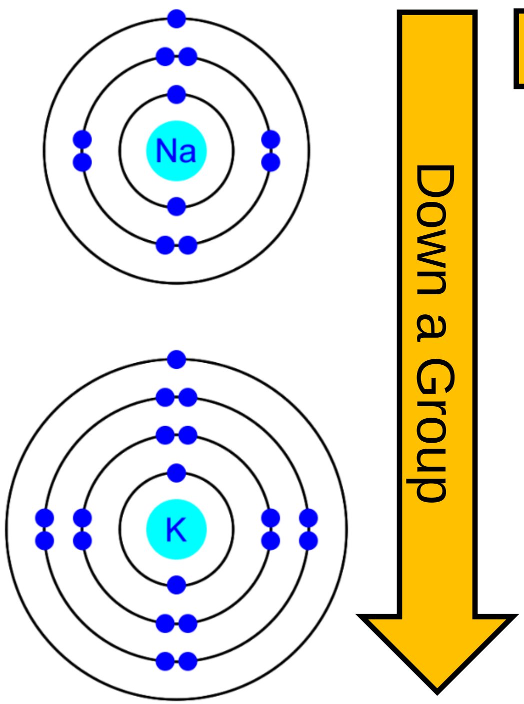
Periodic Trends



Periodic Trends



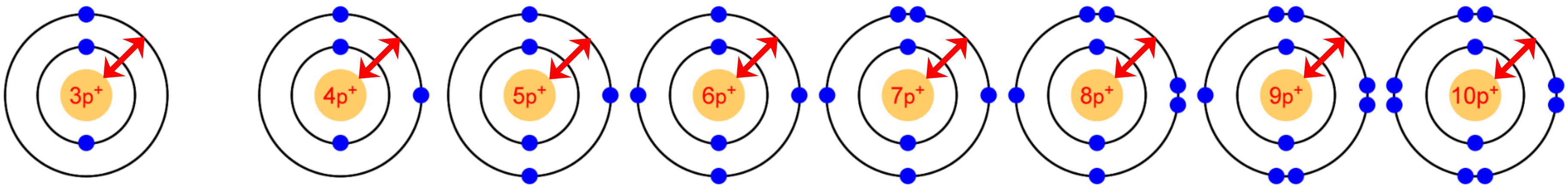


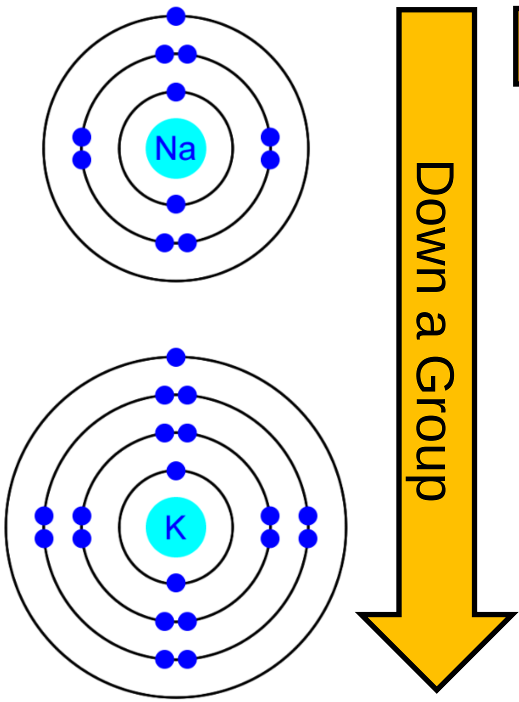


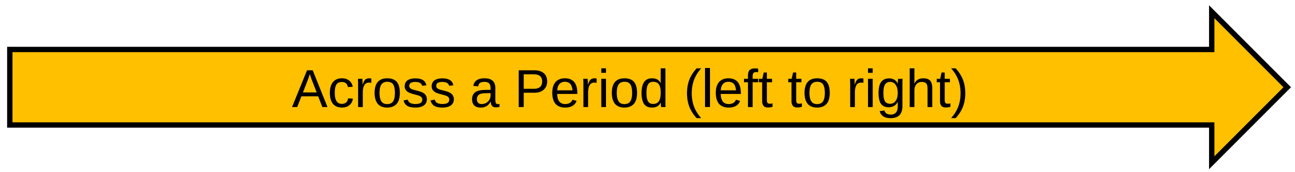
- Moving across a Period, there is an increase in the number of protons in the nucleus (increase in nuclear charge) but the number of electron shells remains constant.

- Due to the increasing nuclear charge, the *electrostatic force of attraction* between the nucleus and electrons in the valence shell *increases across a Period* from left to right.

Periodic Trends







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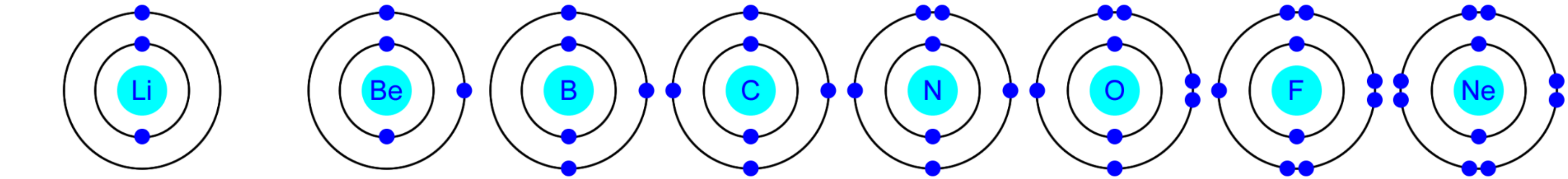
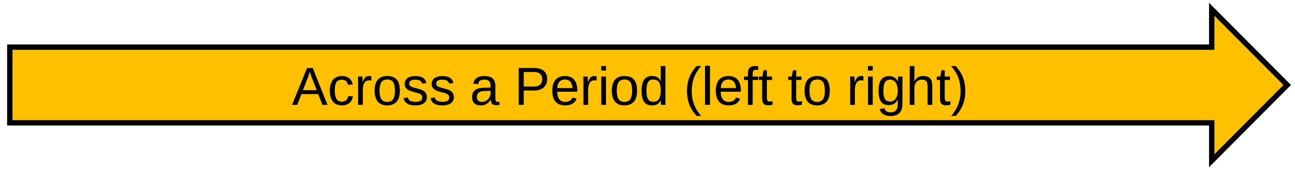
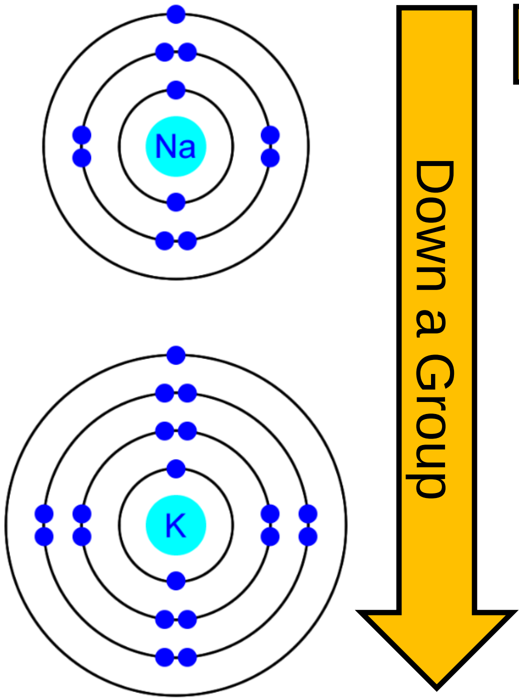
Periodic Trends



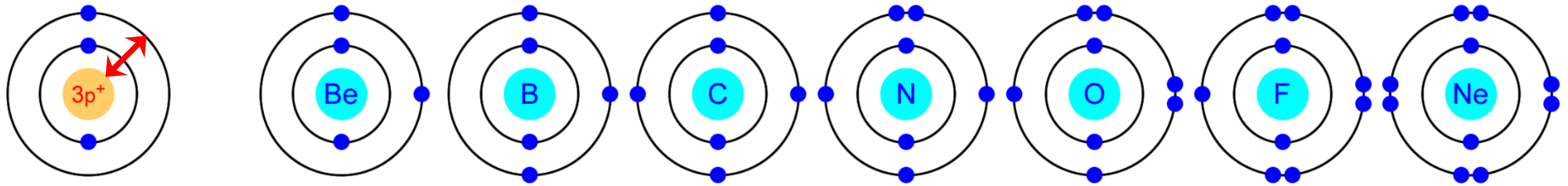


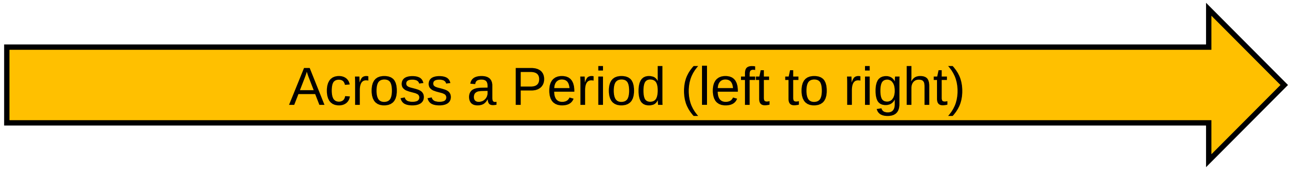


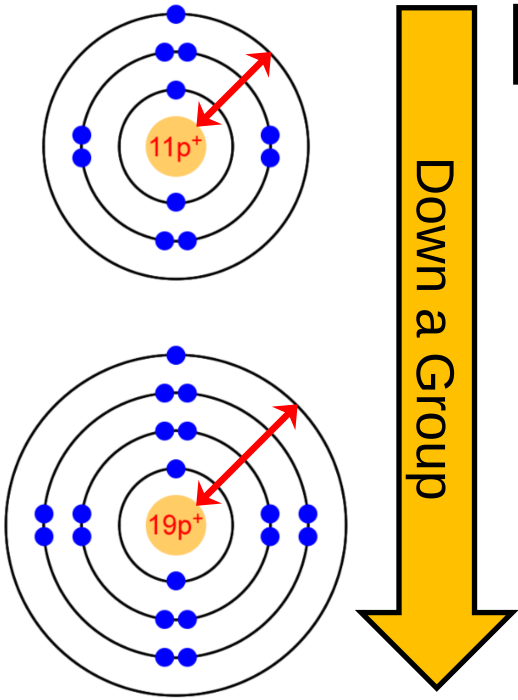
- Moving down a Group, there is an increase in the number of protons in the nucleus (increase in nuclear charge) and also an increase in the number of electron shells.

- 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 shell *decreases down a Group*.

Periodic Trends







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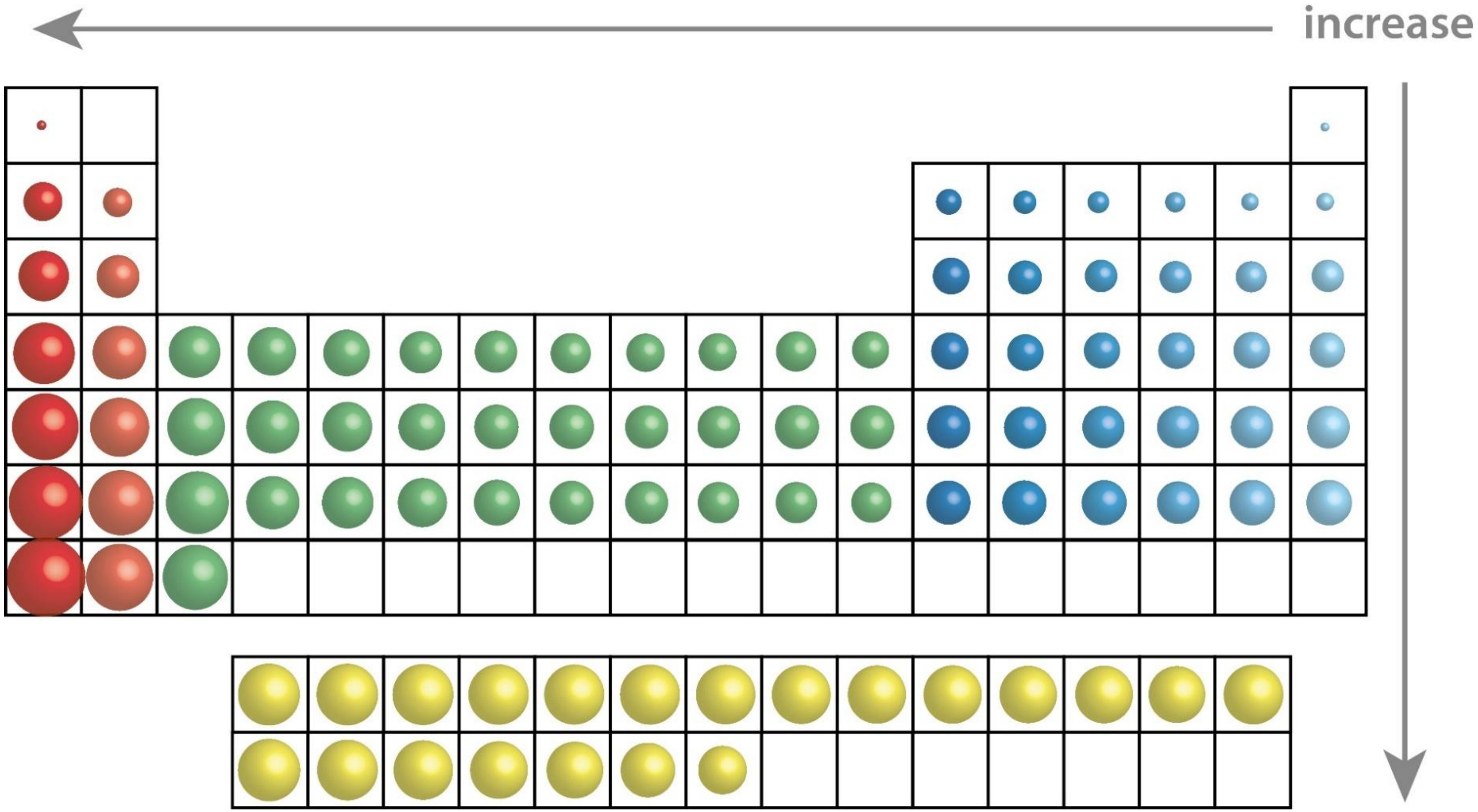
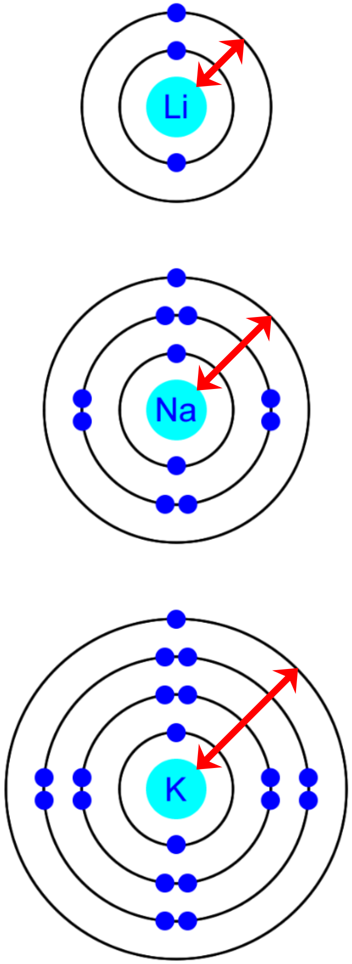
Periodic Trends



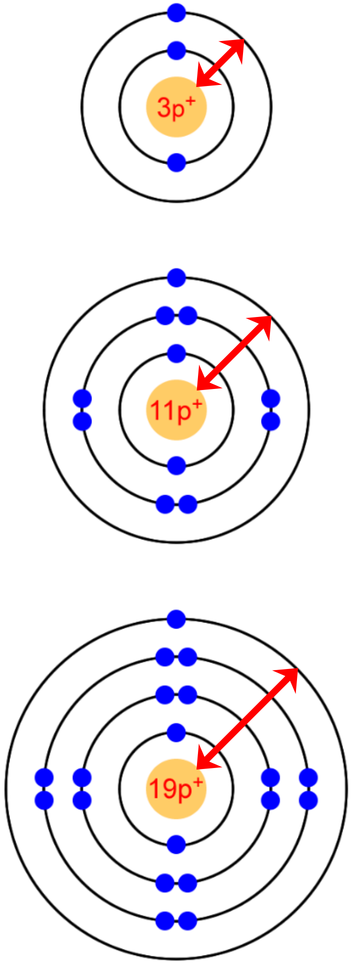
Periodic Trends



- 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



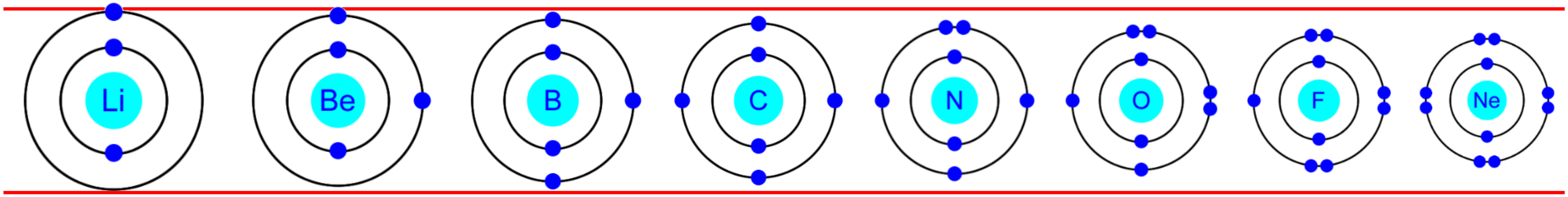
- Atomic radii *increase* from the top to the bottom of any Group of the Periodic Table.
- The number of protons in the nucleus of an atom (nuclear charge) and the number of electron shells around the nucleus of the atom *both increase* down a Group.

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 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*.

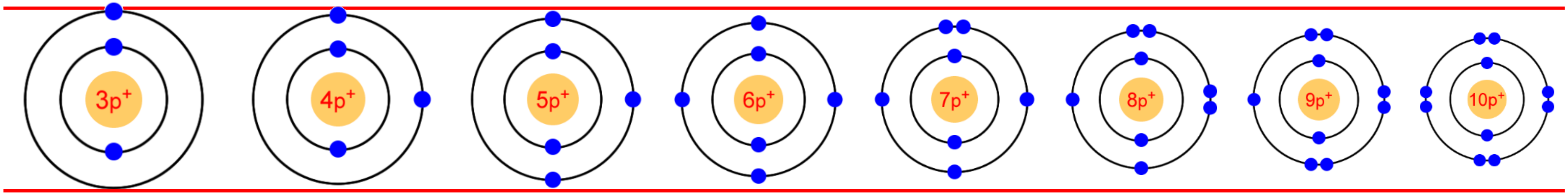
Periodic Trends



- 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



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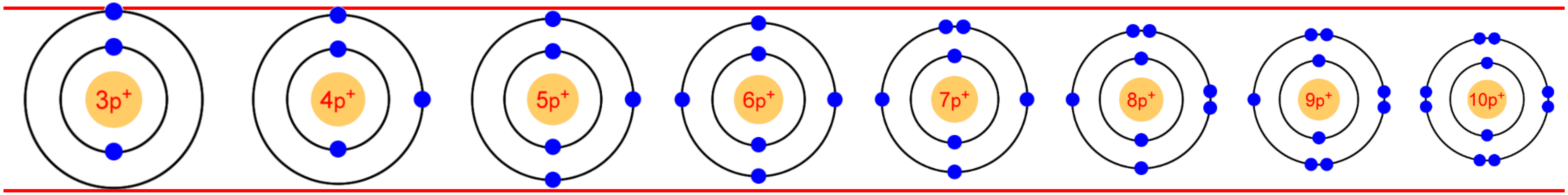
Periodic Trends



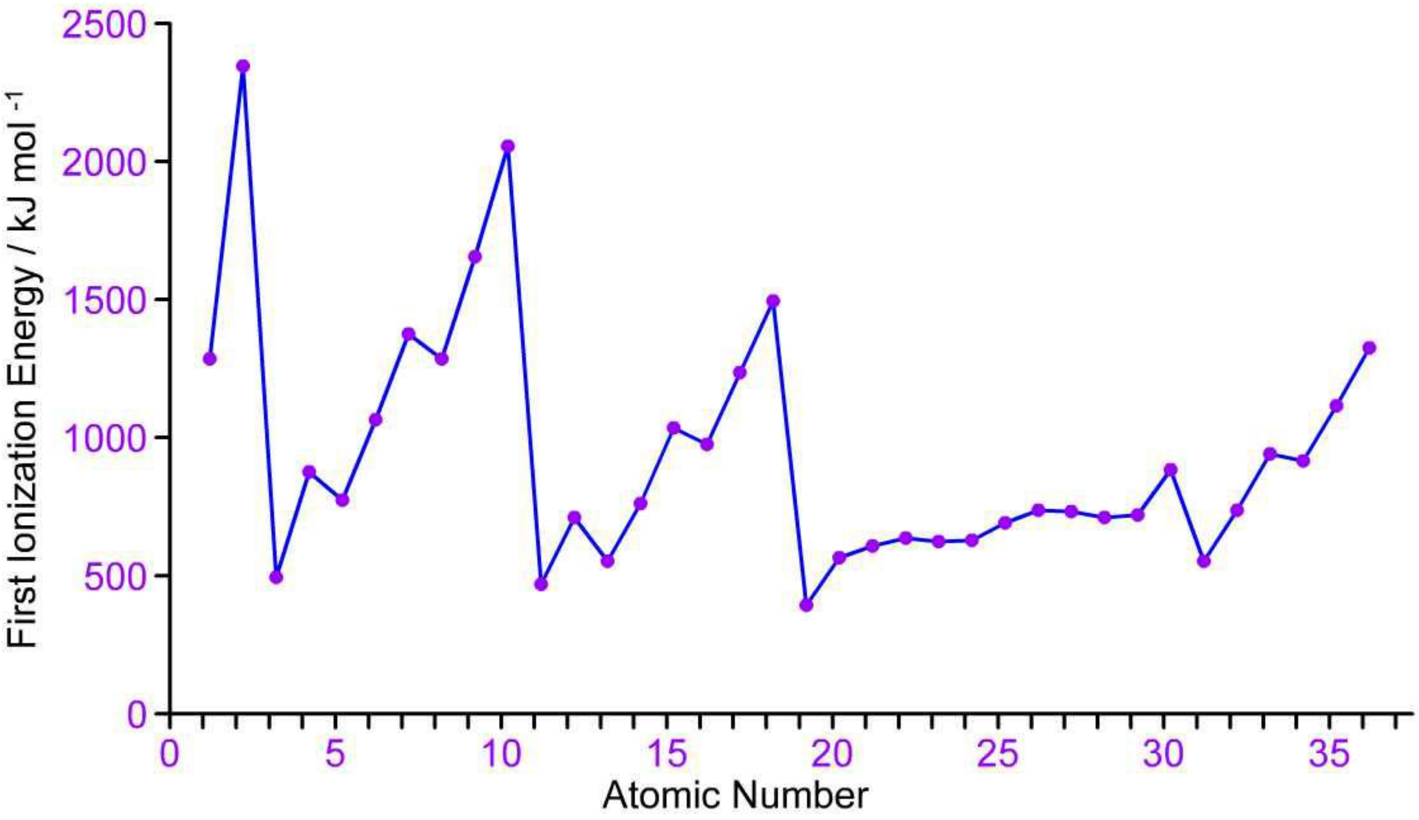
- An increase in nuclear charge means that there is an *increase in the electrostatic force of attraction* between the protons in the nucleus and electrons orbiting the nucleus.

- 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

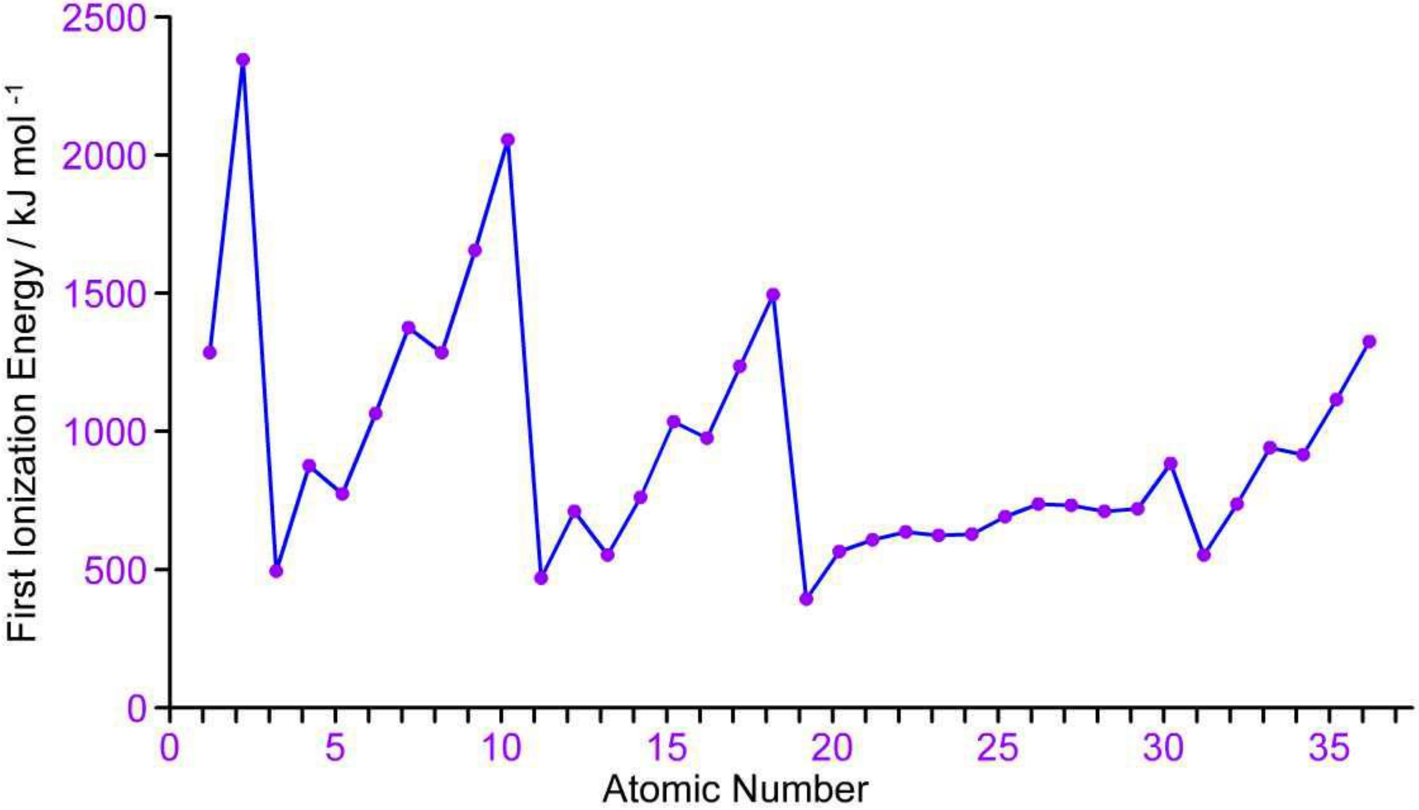
- First ionization energy is the energy required to convert 1 mole (6×10^{23}) of gaseous atoms into one mole (6×10^{23}) of unipositive (1+) gaseous ions.

$$E(g) \rightarrow E^{+}(g) + e^{-}$$



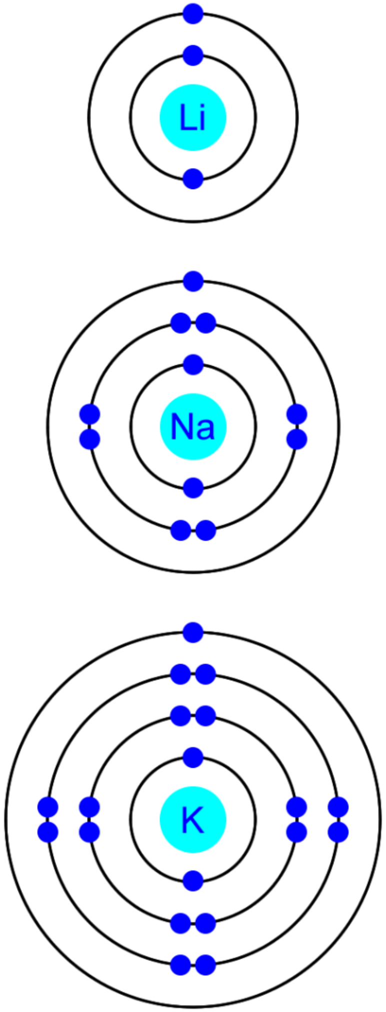
Periodic Trends

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

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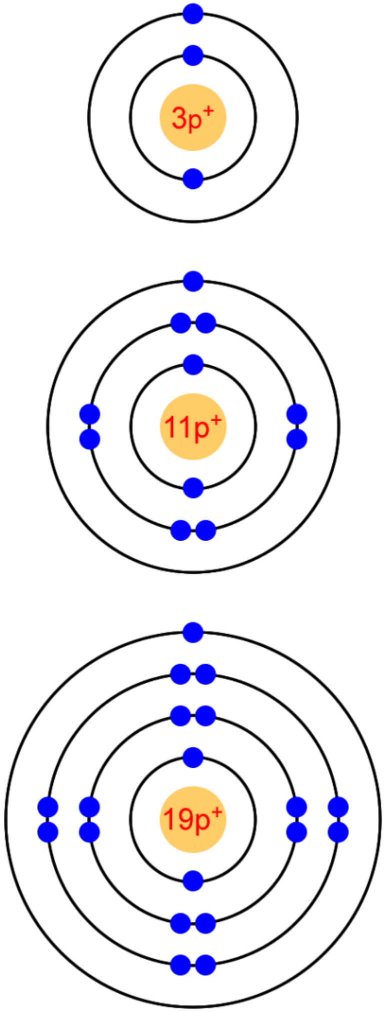

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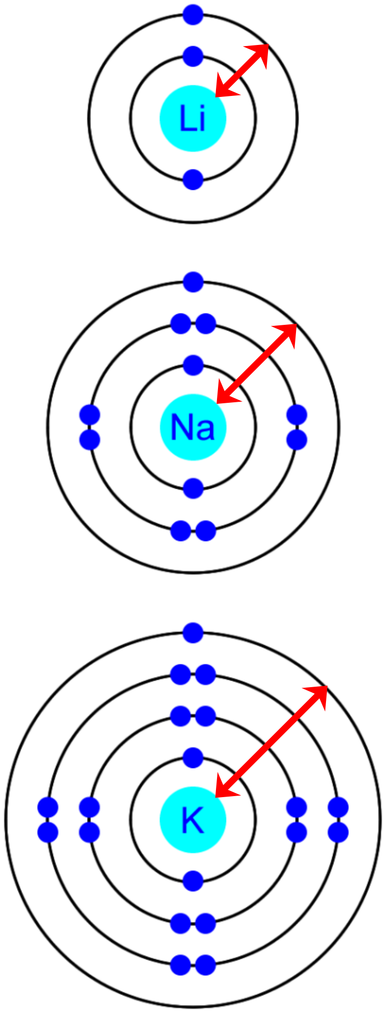
Periodic Trends

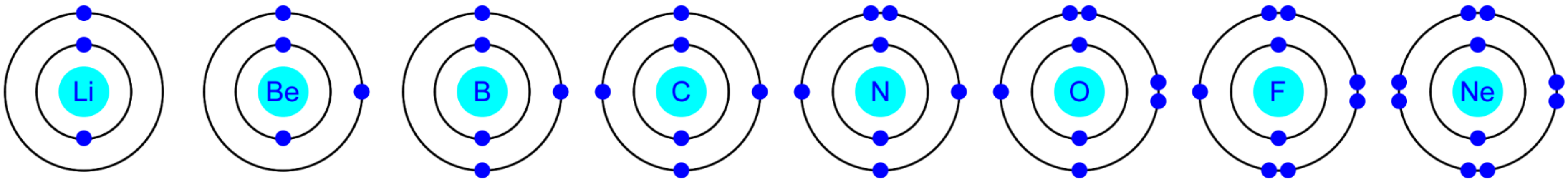
- First ionization energy *decreases down a Group*.



- 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 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 shell. *Less energy* is required to remove an electron from the valence shell, therefore *first ionization energy decreases down a Group*.

Periodic Trends

- First ionization energy *increases across a Period*.

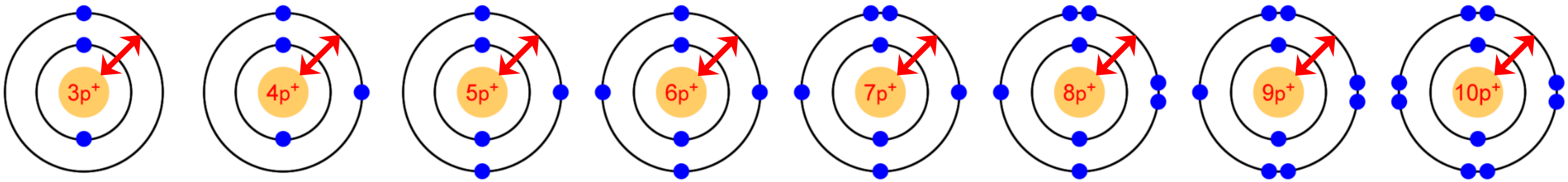


- The number of protons in the nucleus of an atom (nuclear charge) *increases* across a Period, while the number of electron shells is *constant* (remains the same).

- 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*.

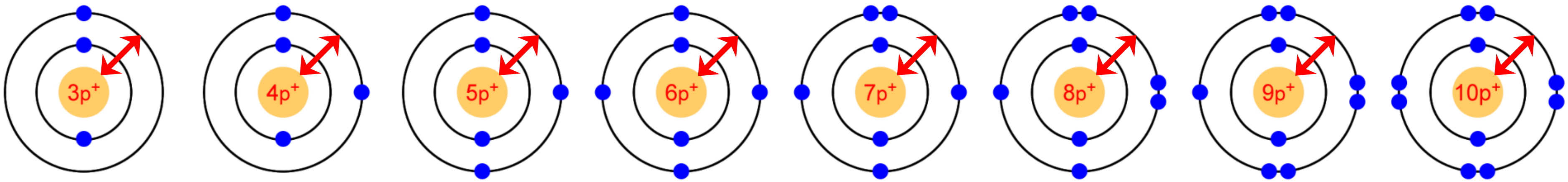


- The number of protons in the nucleus of an atom (nuclear charge) *increases* across a Period, while the number of electron shells is *constant* (remains the same).

- 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*.

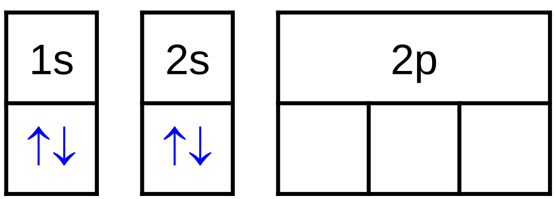


- This increasing nuclear charge exerts a *stronger electrostatic force of attraction* on electrons in the valence shell of the atom (the electrons that are lost during ionization). Therefore, *more energy* is required to remove an electron from the valence shell of the atom, resulting in an *increase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

Periodic Trends

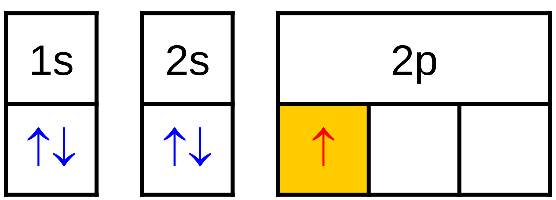
- First ionization energy *decreases* slightly between *Group 2 and Group 13* elements.
- Moving from Group 2 to Group 13, the additional electron enters a *p*-orbital of the same principle quantum shell.
- An electron in a *p*-orbital is *higher in energy* than an electron in the *s*-orbital of the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

Beryllium
Group 2



→ Higher Energy →

Boron
Group 13



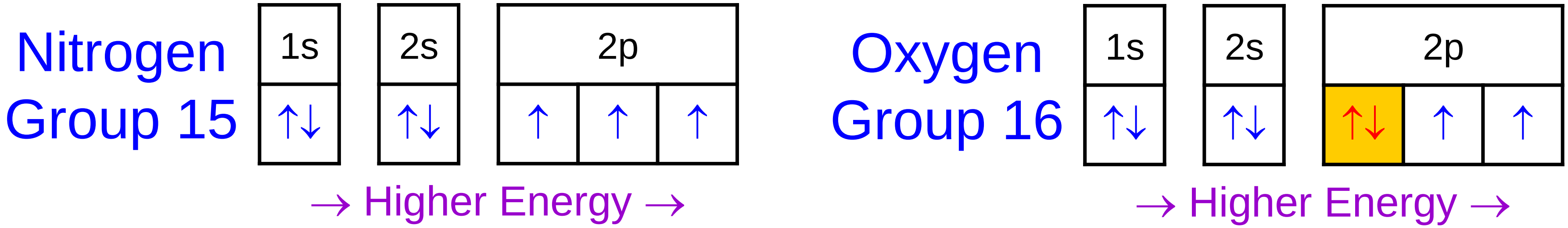
→ Higher Energy →

Periodic Trends

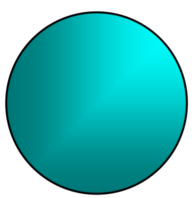
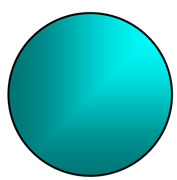
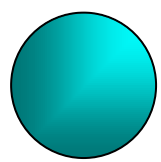
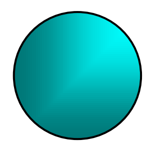
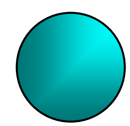
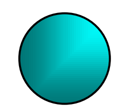
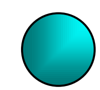
- First ionization energy *decreases* slightly between *Group 15* and *Group 16* elements.

- 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.

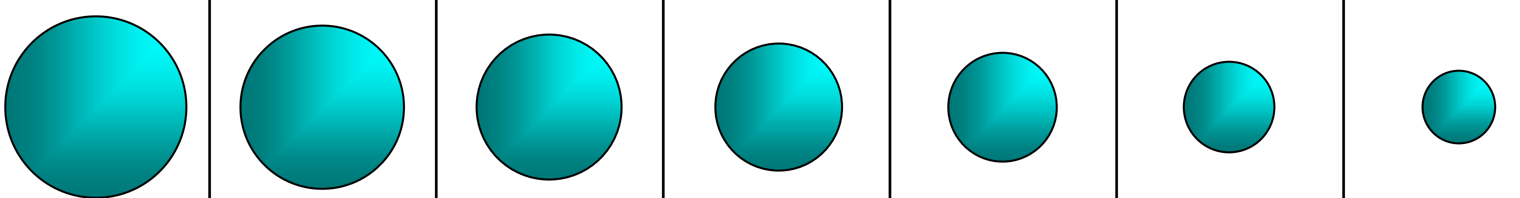


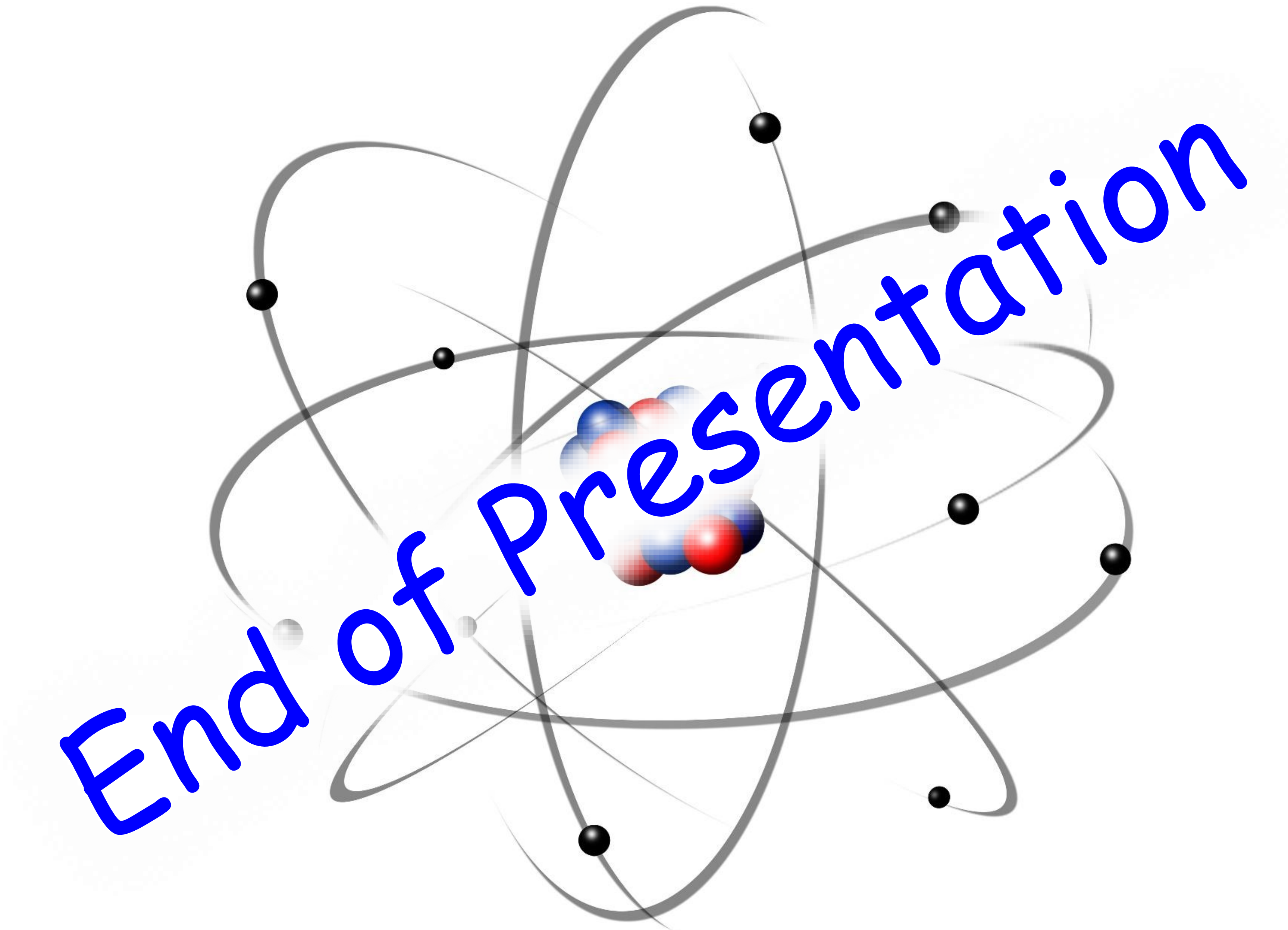
Periodic Trends – Summary

Elements of Third Period	Na Group 1	Mg Group 2	A*l* Group 13	Si Group 14	P Group 15	S Group 16	C*l* 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	A*l* Group 13	Si Group 14	P Group 15	S Group 16	C*l* Group 17
Atomic Number	• 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
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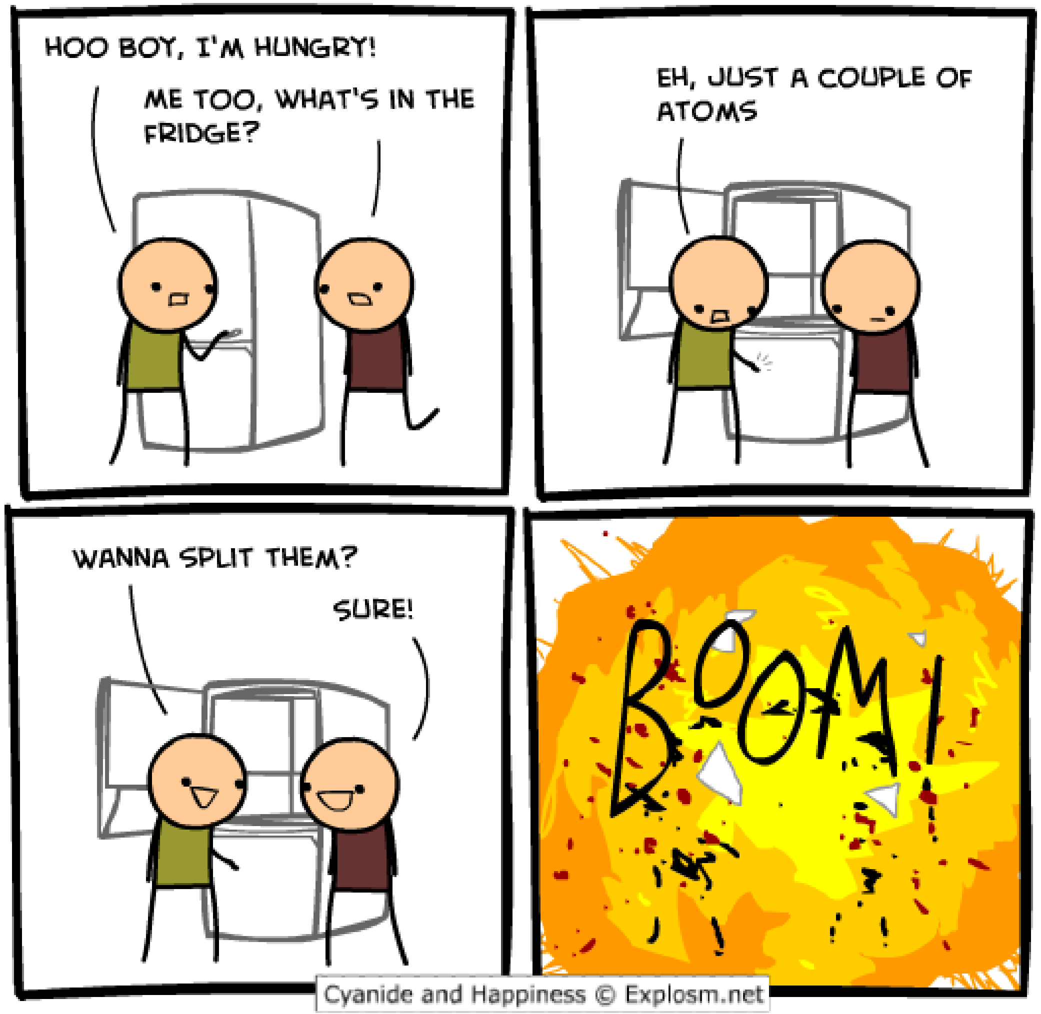
Updated for alignment with the 2017
A' Level Chemistry Periodic Table
8th January 2017

Atomic Structure



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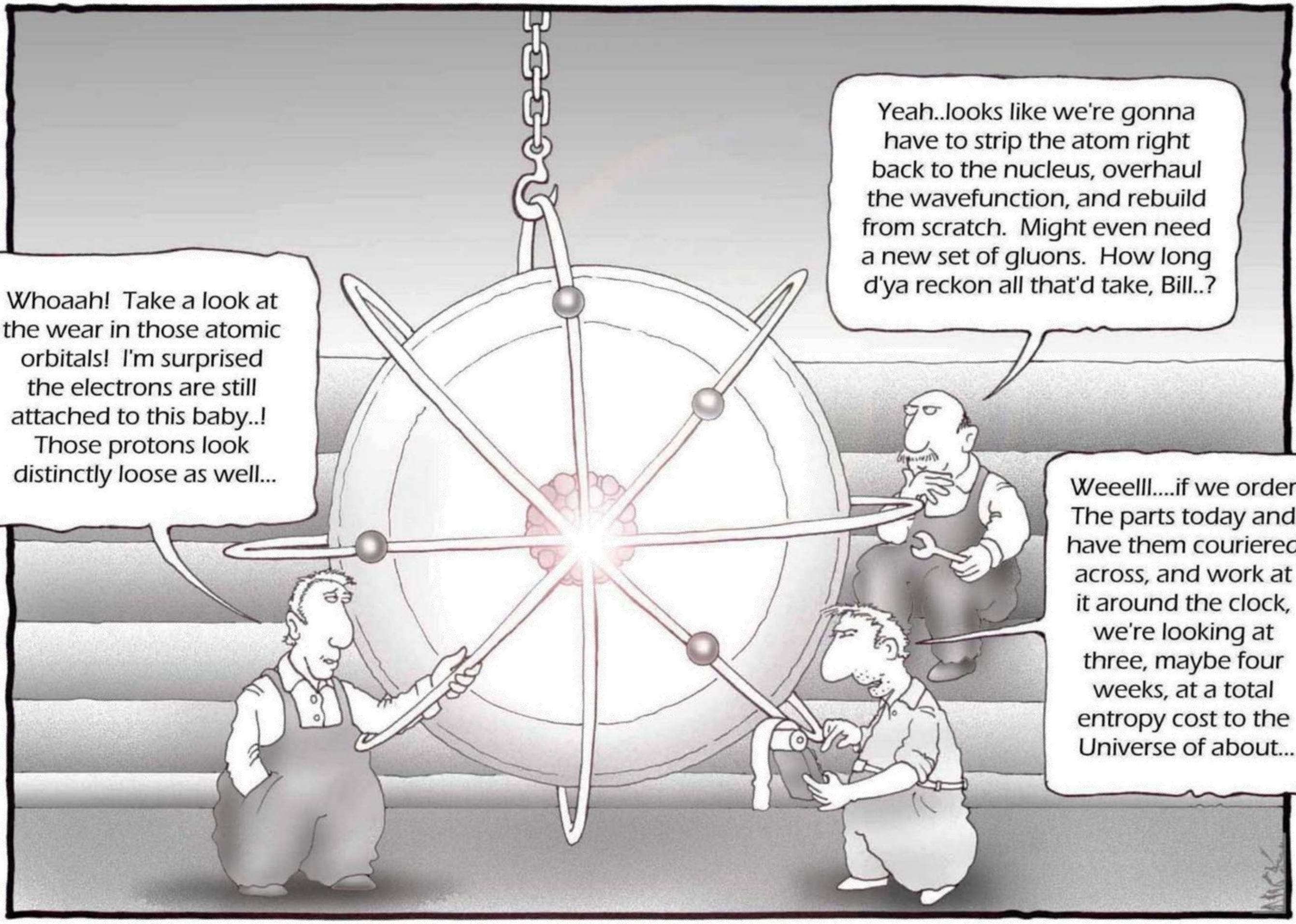

Quantum mechanics.