

# Atomic Structure

*The nuclear atom*  
**Electron Configuration**



**Ms. Thompson - SL Chemistry**  
**Wooster High School**

# ***Topic 2.2***

## ***Electron configuration***

- Emission spectra are produced when photons are emitted from atoms as excited electrons return to a lower energy level.
- The line emission spectrum of hydrogen provides evidence for the existence of electrons in discrete energy levels, which converge at higher energies.
- The main energy level or shell is given an integer number,  $n$ , and can hold a maximum number of electrons,  $2n^2$ .
- A more detailed model of the atom describes the division of the main energy level into s, p, d and f sub-levels of successively higher energies.
- Sub-levels contain a fixed number of orbitals, regions of space where there is a high probability of finding an electron.
- Each orbital has a defined energy state for a given electronic configuration and chemical environment and can hold two electrons of opposite spin.

# ***Electron Configuration***

## **The electromagnetic spectrum**

- **Electromagnetic spectrum** consists of all wavelengths of light:
  - *gamma rays, x-rays, ultraviolet, visible light, infrared, microwaves, and radio waves*
- The energy,  $E$ , of electromagnetic radiation is inversely proportional to the wavelength,  $\lambda$ :

$$E \propto 1/\lambda$$

- Higher energy waves (gamma/x-rays) have small wavelengths where-as low energy radiation (radio/microwaves) have longer wavelengths.
  - Wavelength and frequency are related through the following relationship:
    - $c = \nu \lambda$
- $E = \text{joules}$   
 $\lambda = \text{meters (m)}$   
 $\nu = \text{Hertz (Hz)}$   
 $c = \text{speed of light } (3.00 \times 10^8 \text{ m s}^{-1})$

# Electron Configuration

## The electromagnetic spectrum

### Absorption, emission, and continuous spectra

A white hot metal object emits the full range of wavelengths, producing a **continuous spectrum** including all the colors of the rainbow from red to violet.

*i.e. white incandescent bulb*

When a pure gaseous element is subjected to an electrical discharge, it will emit radiation. This results in an **emission spectrum** which consists of a series of lines against a dark background.

*i.e. hydrogen*

If a cloud of cold gas is placed between a hot metal and a detector, an **absorption spectrum** is observed. This results in a pattern of dark lines against a colored background. The gaseous atoms absorb certain wavelengths of light from the continuous spectrum.

### Continuous Spectrum



### Emission Spectrum



### Absorption Spectrum



# ***Electron Configuration***

## **Emission spectra and Bohr's theory of the hydrogen atom**

- Each element has its own characteristic line spectrum which is used to identify the element
- Lines in emission spectrum have specific wavelengths which corresponds to a discrete amount of energy
  - **Quantization:** Discrete “packets” of electromagnetic radiation
  - A **Photon** is a quantum of radiation
- $E = h\nu = hc/\lambda$   $\longrightarrow$  *Energy of a photon*
  - $h = \text{Planck's constant} = 6.63 \times 10^{-34} \text{ J s}$
  - $\nu = \text{frequency of radiation}$
  - $c = \text{speed of light} = 3.00 \times 10^8 \text{ m s}^{-1}$

# Practice Problem

**... I Do ...**

A certain electromagnetic wave has a wavelength of 625 nm.

a.) What is the frequency of the wave?

*First, convert nm to m:*  $625 \text{ nm} \times \frac{10^{-9} \text{ m}}{1 \text{ nm}} = 6.25 \times 10^{-7} \text{ m}$

*Second, calculate frequency using the equation:*

$$\nu = c/\lambda = \frac{3.00 \times 10^8 \text{ m s}^{-1}}{6.25 \times 10^{-7} \text{ m}} = 4.80 \times 10^{14} \text{ s}^{-1}$$

$$c = \text{speed of light} = 3.00 \times 10^8 \text{ m s}^{-1}$$

# Electron Configuration

## Emission spectra and Bohr's theory of the hydrogen atom

- Neils Bohr proposed a theoretical explanation of the hydrogen line emission

- Each orbit has a definite energy associated with it:

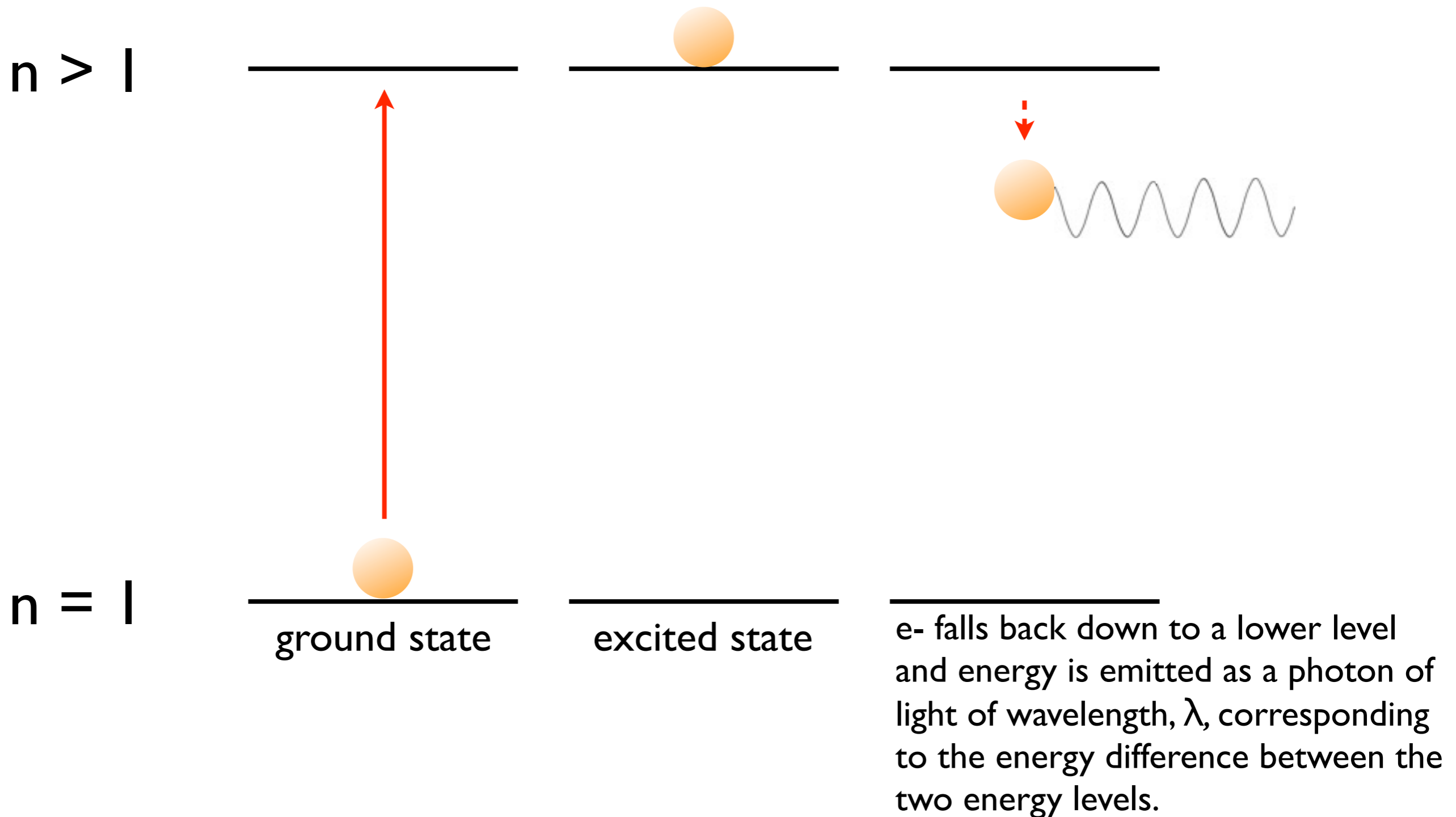
$$E = -R_H(1/n^2)$$

$R_H = \text{Rydberg constant } 2.18 \times 10^{-18} \text{ J}$

$n = \text{principal quantum number, with positive integer values } 1, 2, 3, \dots \text{ depending on the orbit or energy level the electron occupies}$

- When an electron in its **ground state** gets excited, it moves to a higher energy level and stays in this **excited-state** for a fraction of a second
- When the electron falls back from the excited-state to a lower level energy level it emits a **photon**, a discrete amount of energy. This corresponds to a particular wavelength depending on the energy difference between the two levels.

# Principles of the Bohr model of an atom when an electron is excited. $n$ is the principal quantum number





# ***Electron Configuration***

## **Emission spectra and Bohr's theory of the hydrogen atom**

- An electron can be excited to any energy level higher than its current level. They can also fall back down to any lower energy level.
- The difference in energy between the two energy levels can be expressed as follows

$$\begin{aligned}\Delta E &= E_f - E_i \\ &= h\nu = \frac{hc}{\lambda}\end{aligned}$$

# *Electron Configuration*

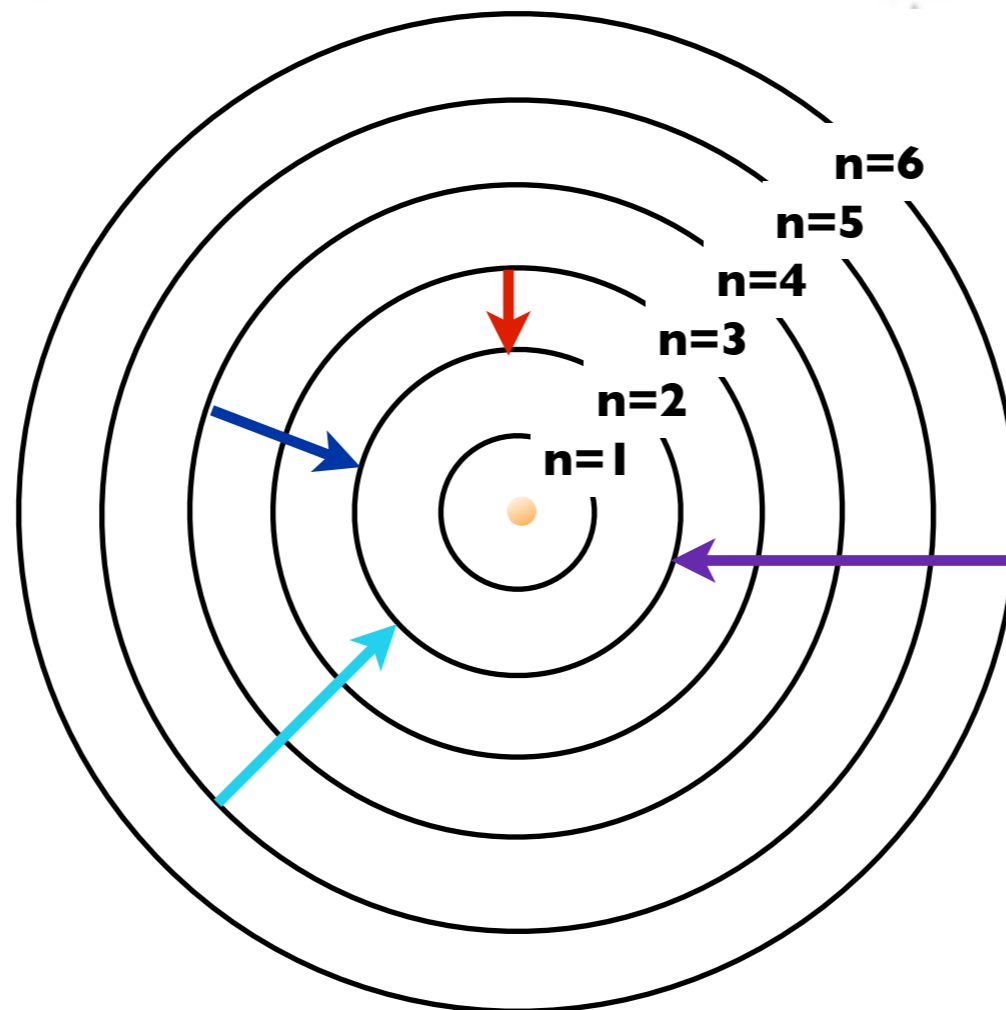
## Emission spectra and Bohr's theory of the hydrogen atom



<b>colour</b>	violet	blue	blue-green	red
<b><math>\lambda/\text{nm}</math></b>	410	434	486	656
<b>transition</b>	$n = 6$ to $n = 2$	$n = 5$ to $n = 2$	$n = 4$ to $n = 2$	$n = 3$ to $n = 2$

# ***Electron Configuration***

## **Emission spectra and Bohr's theory of the hydrogen atom**



# *Electron Configuration*

## **Emission spectra and Bohr's theory of the hydrogen atom**

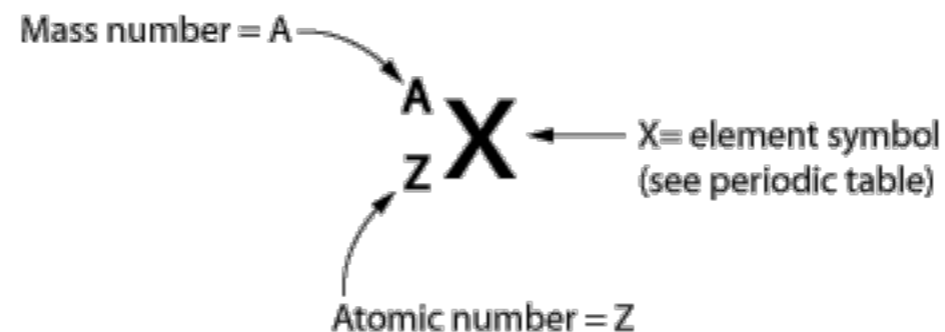
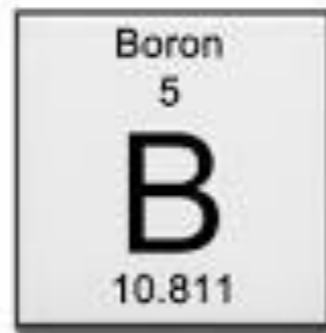


<b>Series</b>	<b><math>n_f</math></b>	<b><math>n_i</math></b>	<b>Region of EMS</b>
Lyman	1	2,3,4,5, ...	UV
Balmer	2	3,4,5,6, ...	visible and UV
Paschen	3	4,5,6,7, ...	IR

# Practice Problem

... I Do ...

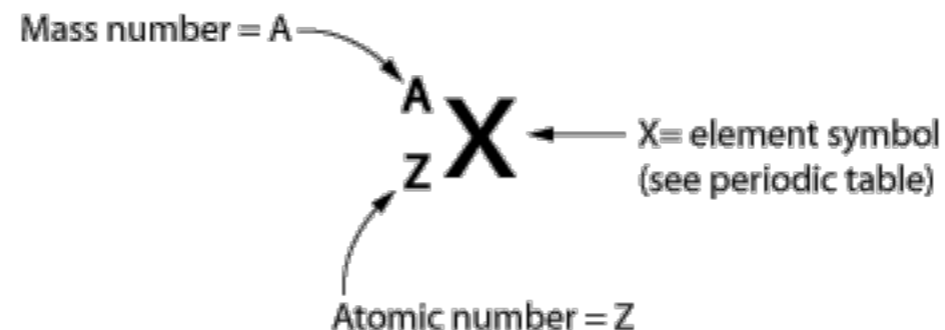
What is the number of protons, electrons, and neutrons in boron-11?



# Practice Problem

... We Do ...

What is the number of protons, electrons, and neutrons  
in  ${}_{16}^{34}\text{S}^{2-}$ ?



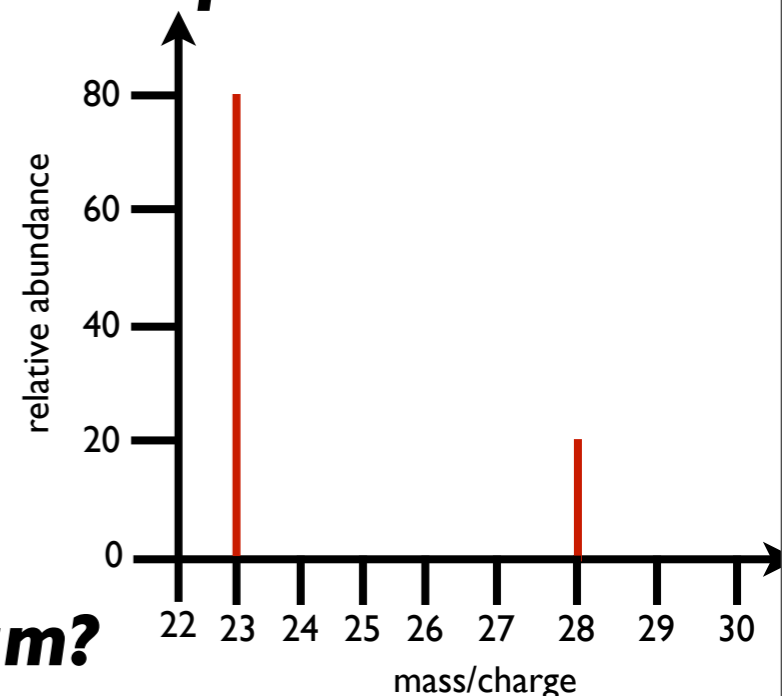
# Practice Problem

20 mins

... You Do ...

**Work with a partner and answer the following questions:**

- 1. A sample of element X contains 69% of  $^{63}\text{X}$  and 31% of  $^{65}\text{X}$ . What is the relative atomic mass of X in this sample?**
- 2. What is the relative atomic mass of an element with the mass spectrum shown to the right?**
- 3. In the emission spectrum of hydrogen, which electronic transition would produce a line in the visible region of the electromagnetic spectrum?**  
A)  $n = 2 \rightarrow n = 1$  B)  $n = 3 \rightarrow n = 2$   
C)  $n = 2 \rightarrow n = 3$  D)  $n = \infty \rightarrow n = 1$



# ***Electron Configuration***

## **History of atomic structure**





# ***Electron Configuration***

## **Heisenberg's uncertainty principle questions:**

1. The wavelength of a quantum object is determined by its:
  - a. Position
  - b. Momentum
  - c. Composition
  - d. Multiple properties
  
2. For a quantum particle to have a long wavelength, it should have:
  - a. Have a large mass and move at high speed
  - b. Have a large mass and move at low speed
  - c. Have a small mass and move at high speed
  - d. Have a small mass and move at low speed
  
3. Adding together multiple waves to make a wave packet means that:
  - a. There are multiple particles in the packet, each with a different wavelength
  - b. There are multiple particles in the packet, each with a different position
  - c. The object has a single momentum but can be found in one of many positions
  - d. The object has a single position but can be found in one of many momenta

# ***Electron Configuration***

**Heisenberg's uncertainty principle**



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# ***Electron Configuration***

## **Schrodinger's equation questions:**

1. We know quantum particles can occupy more than one state at the same time because:
  - a. We find particles confined to a small region of space
  - b. We see particles with well-defined momentum
  - c. We see electrons interfering after passing through a double slit
  - d. We see cats that are both alive and dead
  
2. We see wave behavior in electrons but not in cats because:
  - a. Electrons move much faster than cats
  - b. Electrons move much slower than cats
  - c. Electrons are easier to control than cats
  - d. Electrons have much lower mass than cats
  
3. An electron in the vicinity of two nearby atoms will be orbiting:
  - a. The heavier of the two atoms only
  - b. The lighter of the two atoms only
  - c. Only one of the two atoms, but it's impossible to know which
  - d. Both of the two atoms at the same time

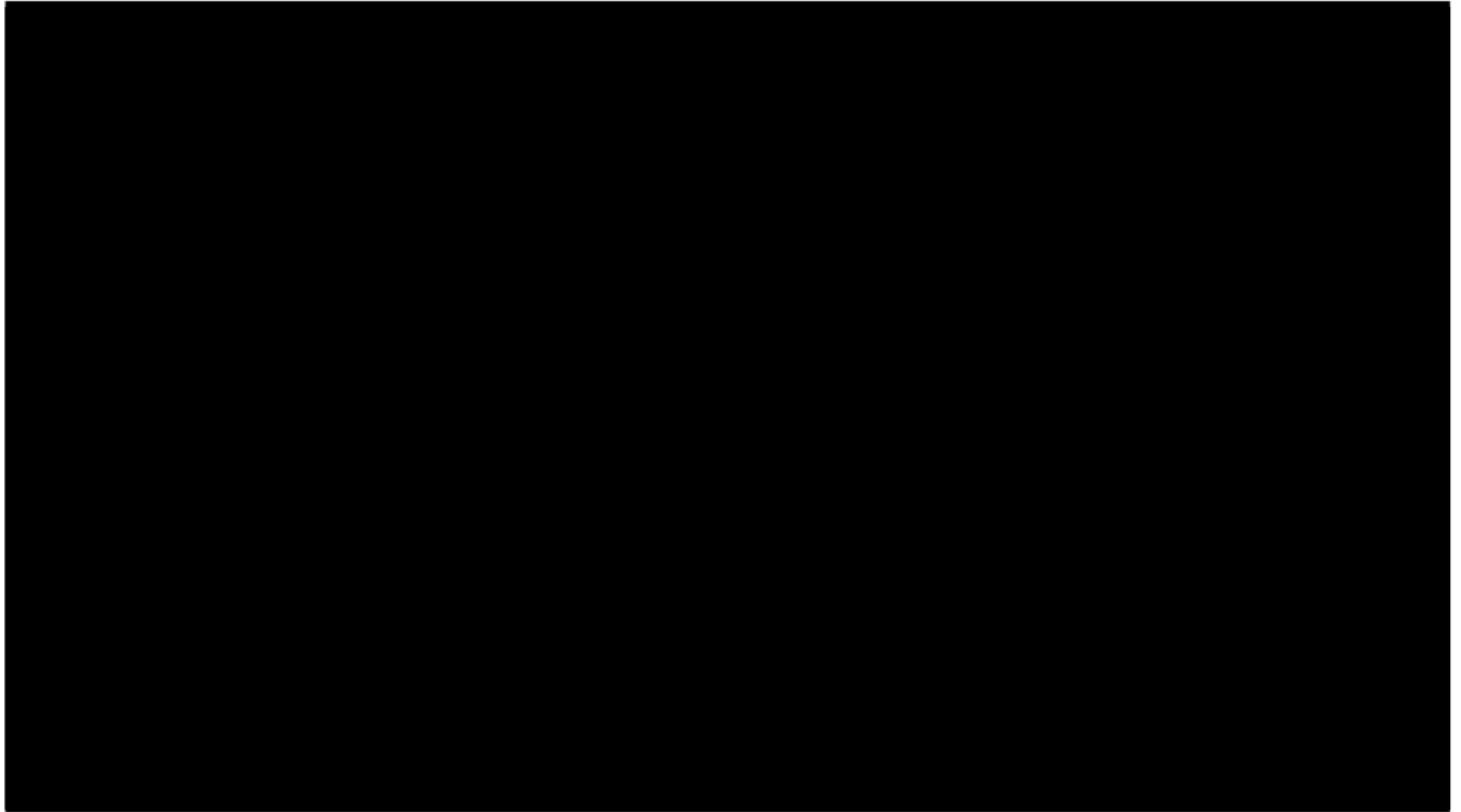
# ***Electron Configuration***

## **Schrodinger's equation questions:**

4. The behavior of electrons moving through solid objects is determined by:
  - a. Electrons bouncing off atoms in the solid like billiard balls
  - b. Electrons going around the outside of the solid like waves
  - c. Electrons being shared between all of the atoms in the solid
  - d. Electrons being pushed through the solid by other electrons
  
5. Quantum physics is important for computer technology because:
  - a. The sharing of electrons between atoms determines the properties of semiconductors
  - b. Internet messages are sent using the wave properties of electrons
  - c. The flow of electrical current can only be understood in terms of waves
  - d. The bits in a modern computer use single electrons

# ***Electron Configuration***

**Heisenberg's uncertainty principle**



# ***Electron Configuration***

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# ***Electron Configuration***

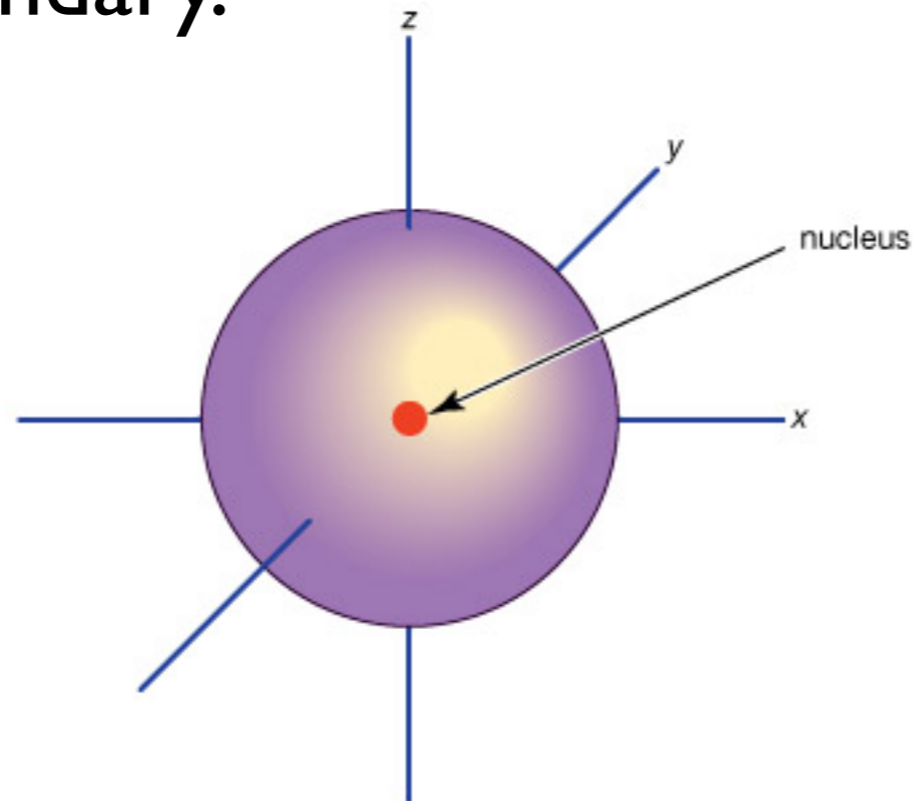
## **Schrodinger and atomic orbitals**

- Schrodinger's work led to a series of mathematical functions called **wavefunctions** describing the electron in the hydrogen atom and associated possible energy states the electron can occupy.
  - $\Psi$  represents wave function and  $\Psi^2$  represents the probability of finding an electron in a region of space at a given point in distance,  $r$ , from the nucleus. Known as **probability density**
- An **atomic orbital** is a region in space where there is a high *probability* of finding an electron.
  - Can hold a *maximum* of two electrons, with opposite spins
  - Types of orbitals:  $s, p, d, f$  (*for SL only need to know  $s$  and  $p$  orbital shape*)
  - Each has a characteristic *shape* and associated *energy*

# ***Electron Configuration***

## **The atomic orbitals**

- **The s atomic orbital**
  - Spherically symmetrical
  - 99% chance of probability of finding an electron within the spherical boundary.

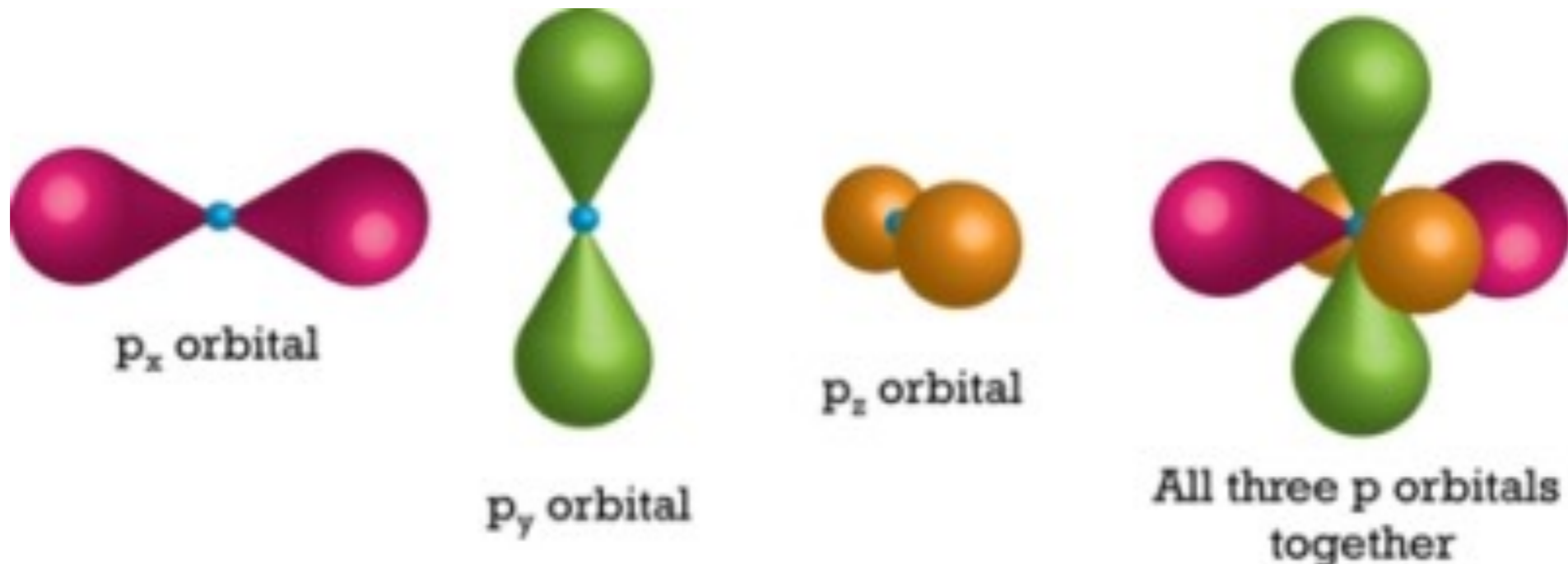


# Electron Configuration

## The atomic orbitals

- **The p atomic orbital**

- Dumbbell shaped
- Three p atomic orbitals  $p_x$ ,  $p_y$ ,  $p_z$  all with boundary surfaces conveying probable electron density pointing in different directions along the three respective Cartesian axes x, y, and z.



# ***Electron Configuration***

## **The atomic orbitals**

- **Energy levels, sublevels, orbitals, and electron spin**
  - **Principal quantum numbers:**  $n$ , energy levels (1, 2, 3, 4, ...)
  - As  $n$  increases, the mean position of the electron is further from nucleus, therefore, orbital energy increases as  $n$  increases.
  - Each can hold two a max of two electrons given by  $2n^2$
  - Electron capacity for  $n=1$  is 2,  $n=2$  is 8, for  $n=3$  is 18.
  - On periodic table:
    - Row 1 contains two elements
    - Row 2 contains eight elements
    - Row 3 contains eight elements
    - Row 4 contains 18 elements

# Electron Configuration

## The atomic orbitals

- **Energy levels, sublevels, orbitals, and electron spin**
  - Energy levels are split into **sublevels**, which there are four common types: *s*, *p*, *d*, and *f*.
  - Each sublevel contains a number of orbitals, each of which can hold a max of two electrons

Sublevel	Number of orbitals in sublevel	Maximum number of e <sup>-</sup> in sublevel
<i>s</i>	1	2
<i>p</i>	3	6
<i>d</i>	5	10
<i>f</i>	7	14

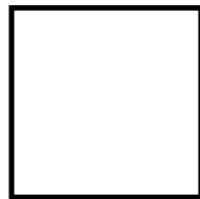


# Electron Configuration

## Orbital diagrams

Used to represent the electrons in these atomic orbitals. We will use these orbital diagrams to represent electron configurations!!!

*s* sublevel (one box represents *s* orbital)



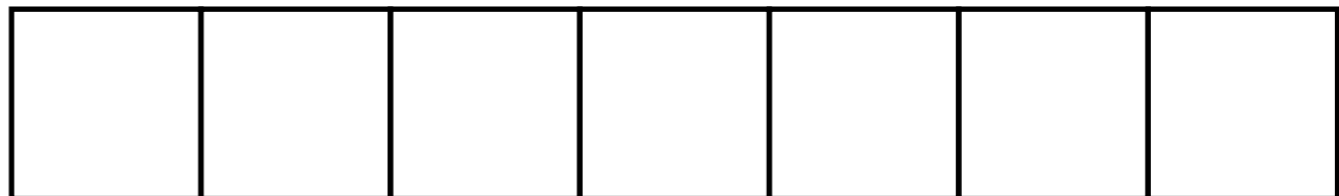
*p* sublevel (three boxes represents the three *p* orbitals)



*d* sublevel (five boxes represents the five *d* orbitals)



*f* sublevel (seven boxes represents the seven *f* orbitals)



# Electron Configuration

## Orbital diagrams

Two electrons in same orbital will have opposite spin values of the spin **magnetic quantum number,  $m_s$** .

The sign of  $m_s$  (+1/2 or -1/2) indicates the orientation of the magnetic field generated by the electron

Electrons behave like two magnets facing opposite directions and is represented by two opposite arrows in a box.

*s sublevel (one box represents s orbital)*



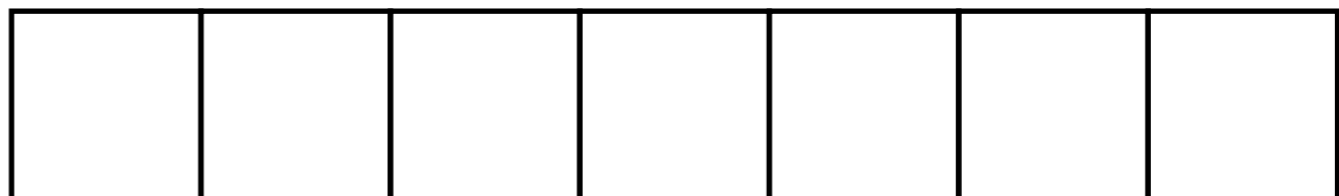
*p sublevel (three boxes represents the three p orbitals)*

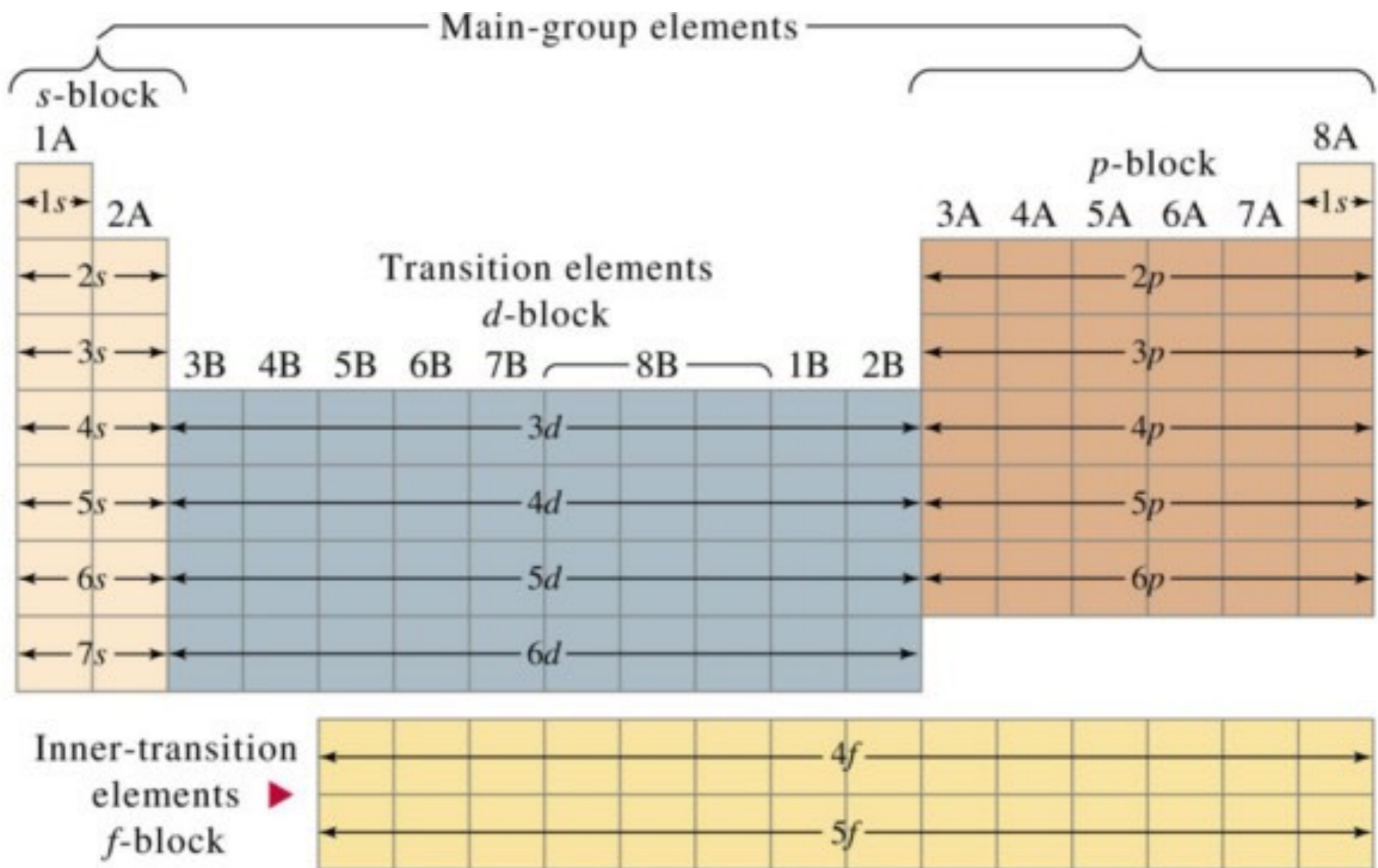


*d sublevel (five boxes represents the five d orbitals)*



*f sublevel (seven boxes represents the seven f orbitals)*





# Electron Configuration

## Aufbau principle

- Electrons will fill from the *lowest-energy orbital* that is available first.
- Works up to Ca ( $Z=20$ ) after that Sc ( $Z=21$ ), the d orbital fills before the s and p orbitals since it is lower in energy.
- **Electron configuration for Ca ( $Z=20$ )**
  - $1s^2, 2s^2, 2p^6, 3s^2, 3p^6, 4s^2$
- **Electron configuration for Sc ( $Z=21$ )**
  - $1s^2, 2s^2, 2p^6, 3s^2, 3p^6, 3d^1, 4s^2$  ← notice the d orbital is filled before the s orbital
- The **Pauli Exclusion Principle** states that any orbital can hold a maximum of two electrons, and these electrons have opposite spin.
- **Hund's rule of maximum multiplicity** states that when filling **degenerate** orbitals (orbitals of equal energy) electrons fill all the orbitals singly before occupying them in pairs.



# Electron Configuration

## Full electron configuration for first 36 elements

- Three ways to illustrate electron configuration:
  - *full electron configuration*
  - *condensed electron configuration*
  - *orbital diagram representation*
- Use “build up” method

Element	Z	Electron configuration
<i>Period 1 elements:</i>		
H	1	$1s^1$
He	2	$1s^2$
<i>Period 2 elements:</i>		
Li	3	$1s^2 2s^1$
Be	4	$1s^2 2s^2$
B	5	$1s^2 2s^2 2p^1$
C	6	$1s^2 2s^2 2p^2$
N	7	$1s^2 2s^2 2p^3$
O	8	$1s^2 2s^2 2p^4$
F	9	$1s^2 2s^2 2p^5$
Ne	10	$1s^2 2s^2 2p^6$
<i>Period 3 elements: continue with the same filling pattern, for example:</i>		
Na	11	$1s^2 2s^2 2p^6 3s^1$
Mg	12	$1s^2 2s^2 2p^6 3s^2$
Al	13	$1s^2 2s^2 2p^6 3s^2 3p^1$
Ar	18	$1s^2 2s^2 2p^6 3s^2 3p^6$
<i>Period 4 elements:</i> After Z = 30 the 4p sublevel is filled:		
K	19	$1s^2 2s^2 2p^6 3s^2 3p^6 4s^1$
Ca	20	$1s^2 2s^2 2p^6 3s^2 3p^6 4s^2$
Sc	21	$1s^2 2s^2 2p^6 3s^2 3p^6 3d^1 4s^2$
Ni	28	$1s^2 2s^2 2p^6 3s^2 3p^6 3d^8 4s^2$
Zn	30	$1s^2 2s^2 2p^6 3s^2 3p^6 3d^{10} 4s^2$
Ga	31	$1s^2 2s^2 2p^6 3s^2 3p^6 3d^{10} 4s^2 4p^1$
Br	35	$1s^2 2s^2 2p^6 3s^2 3p^6 3d^{10} 4s^2 4p^5$
Kr	36	$1s^2 2s^2 2p^6 3s^2 3p^6 3d^{10} 4s^2 4p^6$

# ***Electron Configuration***

## **Condensed electron configuration**

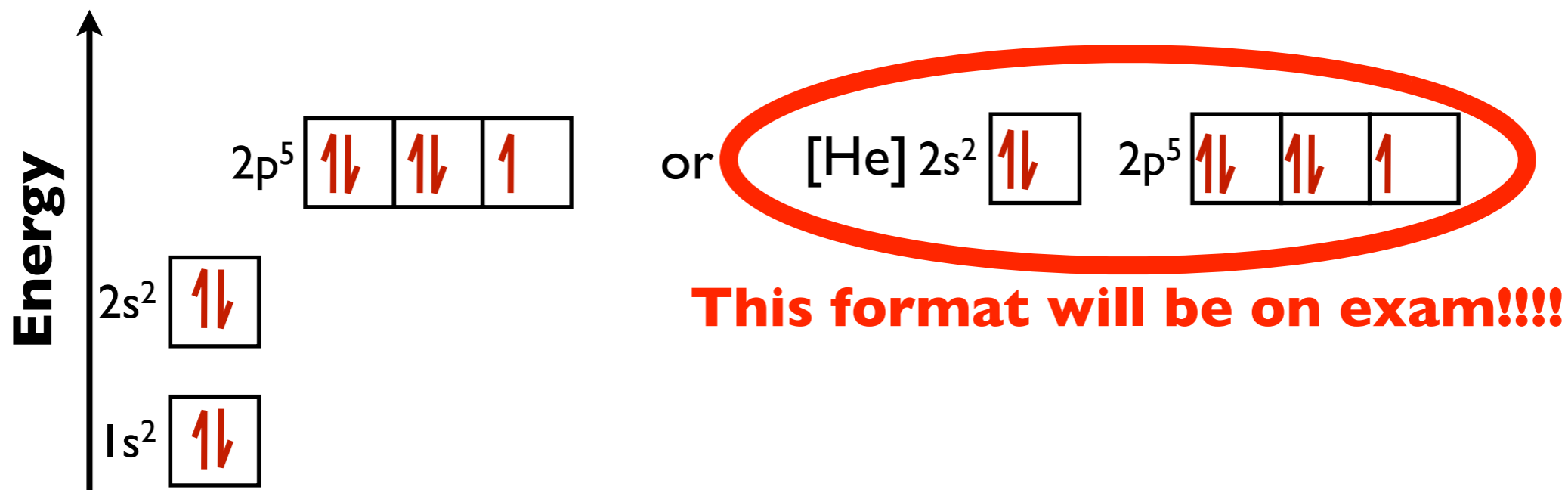
- Full electron configuration can get lengthy so scientists often write in condensed form:
  - [nearest noble gas] + valence electrons
  - Valence electrons are responsible for chemical reactions
- For example:

He	[He]
O	[He] $2s^2$ , $2p^4$
Ne	[He] $2s^2$ , $2p^6$ or simply [Ne]
P	[Ne] $3s^2$ , $3p^3$

# Electron Configuration

## Orbital diagrams

- **Orbital diagrams** uses the arrows in boxes notation
  - Electrons are represented by arrows  $\uparrow\downarrow$
  - Boxes represent orbitals
  - Degenerate orbitals are joined together to show their energy equivalence
  - Can show full configuration or just from nearest noble gas
  - Fluorine  $1s^2, 2s^2, 2p^5$



# Practice Problem

**... I Do ...**

What is the order of increasing energy of the orbitals within a single energy level?

A)  $d < s < f < p$

B)  $s < p < d < f$

C)  $p < s < f < d$

D)  $f < d < p < s$



# Practice Problem

... **We Do** ...

List the following types of electromagnetic radiation in order of **increasing** wavelength (shortest first).

I. Yellow light

II. Red light

III. Infrared radiation

IV. Ultraviolet radiation

b) Distinguish between a continuous spectrum and a line spectrum.

# Practice Problem

20 mins

**... You Do ...**

**Work with a partner and answer the following questions:**

- Deduce the full electron configuration for Mn and  $Mn^{2+}$ .*
- Deduce the condensed electron configuration for  $Cu^{2+}$ .*
- Draw orbital diagrams for  $Co^{2+}$  and As.*

# ***Topic 2.2***

## ***Electron configuration***

- ➔ Emission spectra are produced when photons are emitted from atoms as excited electrons return to a lower energy level.
- ➔ The line emission spectrum of hydrogen provides evidence for the existence of electrons in discrete energy levels, which converge at higher energies.
- ➔ The main energy level or shell is given an integer number,  $n$ , and can hold a maximum number of electrons,  $2n^2$ .
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