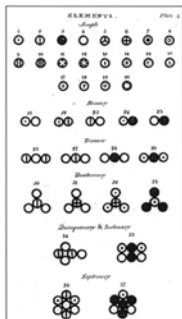


Atoms, Molecules, the Chemical Bond and Gas Laws



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1

Course Overview

During the first part of the course we will be doing some introductory 'physical' chemistry, revision for some of you, new to others:

- *Atomic structure (2.1, 2.2, 2.3) and atomic mass (3.1, 3.2, 3.3); quantum numbers and atomic orbitals (7.6, 7.7).*
- *Electron configuration and the periodic table (7.8, 7.9)*
- *Ionic (9.2) and covalent bonding (9.4, 9.5); Lewis dot structures (9.6, 9.7, 9.8).*
- *Redox reactions (4.4). Oxidation numbers (4.4). Balancing redox equations (19.1).*

During the second part of the course we will introduce physical chemistry and the gas laws (5.1-5.6):

- *Observations, laws, theories, models and predictions.*
- *Gas Laws; ideal gas equation; Dalton's law of partial pressures*

Recommended Reading:

'Chemistry'; R. Chang, McGraw-Hill, 9th Ed., 2007, **sections as listed.**

2

Introduction

Atoms, Molecules and the Chemical Bond

Atoms: These are the smallest particle of an element that retain the chemical properties of the element: *átomos* (Greek) = 'uncuttable'.

Atoms are made up of smaller sub-atomic particles – there are many of these but most of chemistry is concerned with just 3 sub-atomic particles:

How big are atoms?

Atoms are extremely small and most measure between 30-150 pm (30-150 $\times 10^{-12}$ m in diameter) and weigh between 1.67×10^{-27} and 4.52×10^{-25} kg.

Scale:

3

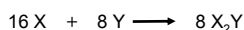
Dalton's Atomic Theory (1808):

Elements are composed of extremely small particles called **atoms**.

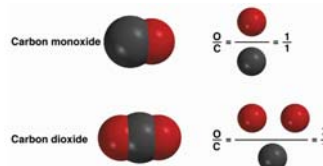


John Dalton (1766-1844):
Father of Atomic Theory

Compounds are composed of atoms of more than one element. In any compound, the ratio of the numbers of atoms of any two of the elements present is either an integer or a simple fraction.



Law of Conservation of Mass



Law of Multiple Proportions

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Figs. 2.1 & 2.2, Chang

The Structure of the Atom:

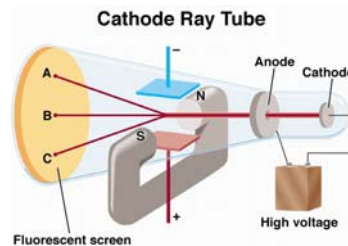
The Electron – J.J. Thompson found that atoms contain negatively charged particles (electrons) using a cathode ray tube. He also determined the charge/mass ratio.



J.J. Thompson (1856-1940):
Nobel Prize Physics (1906)

- A: Magnetic field on
- B: No field
- C: Electric field on

Fig. 2.3, Chang



Later R.A. Millikan (Physics Nobel 1923) measured the charge of the electron (-1.6033×10^{-19} C) using an 'oil-drop' experiment.

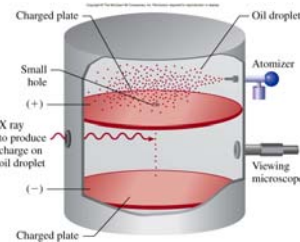


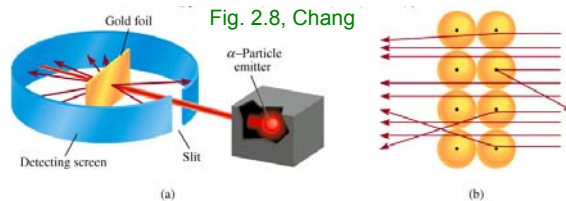
Fig. 2.5, Chang

5

The Proton – Discovered by Ernest Rutherford, Canterbury graduate. He bombarded a piece of thin gold foil with α -particles (these are +ve charged He nuclei) and found that some of the particles bounced back!



Lord Rutherford (1871-1937):
Nobel Prize Chemistry (1908)



If the size of the atom was expanded to that of a sports stadium, the size of the nucleus would be that of a marble.

6

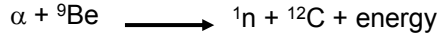
The Neutron – Discovered by James Chadwick (1932). Rutherford's model of the atom did not account for the following anomaly:



James Chadwick (1891-1974):
Nobel Prize Physics (1935)



Hydrogen: 1 proton; Helium: 2 protons



Chadwick showed that this disintegration was accompanied by release of : neutral particles of comparable mass to the proton, thus explained anomaly.

Electron cloud:

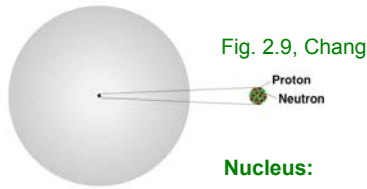


Fig. 2.9, Chang

Particle	Mass (g)	Charge	
		Coulomb	Charge Unit
Electron*	9.10939×10^{-28}	-1.6022×10^{-19}	-1
Proton	1.67262×10^{-24}	$+1.6022 \times 10^{-19}$	+1
Neutron	1.67493×10^{-24}	0	0

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Atomic Number, Mass Number and Isotopes:

All neutral atoms of the same element have the **same number** of **protons** and **electrons**.

The **Atomic Number (Z)** =
Mass Number (A) =

Isotopes are atoms of the same element (**X**) with different numbers of in their nuclei.

Example: Carbon (C) is the 6th element in the periodic table.

Therefore, C has 6 protons (and 6 electrons): **Z =**

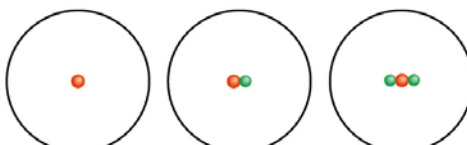
The mass number, **A =** , is the total of protons + neutrons, therefore there are neutrons.

8

The following symbolism is used to denote an elements make-up:



The Isotopes of Hydrogen



${}^1_1\text{H}$ ${}^2_1\text{H}$ ${}^3_1\text{H}$
 H atom Deuterium D Tritium T
 (1p, 1e⁻, 0n) (1p, 1e⁻, 1n) (1p, 1e⁻, 2n)

Relative atomic mass:

Naturally occurring Magnesium

78.7% ${}^{24}_{12}\text{Mg}$ (12 p, 12 n)

10.1% ${}^{25}_{12}\text{Mg}$ (12 p, 13 n)

11.2% ${}^{26}_{12}\text{Mg}$ (12 p, 14 n)

Relative atomic mass:

≈

=

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Avogadro's Number and the Molar Mass of an Element:

Macroscopic samples of matter contain an enormous number of atoms. A special unit has been devised to describe this large number:

A **Mole (mol)**:

1 mol =

: Avogadro's number (N_A)



Amadeo Avogadro
(1776-1856):



One Mole of:



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Do You Understand Molar Mass?

How many atoms are in 0.551 g of potassium (K)?

Electronic Structure of Atoms (Chang, Chapter 7):

We have seen the model of the an atom as described by Rutherford earlier, where there is a small massive nucleus and the electrons move around the outside.

Early ideas suggested atoms might be like miniature 'solar systems' where the nucleus = sun; electrons = orbiting planets.

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There are problems with this 'solar system' view:

-
-
-

Summary of Quantum Theory (Chang, Chapter 7):

Quantum theory: electrons in an atom can only exist in certain allowed states.

Each state is fully described by **Four Quantum Numbers:**

n : Principal quantum number can have integral values: 1, 2, 3, ... gives us an indication of how far away from the nucleus the electron is.

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l : Angular momentum quantum number, tells us the 'shape' of the orbital.
The values of l depend on the principal quantum number, n .

Each value of l is given a letter designation:

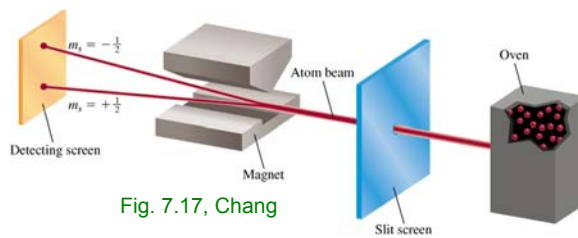
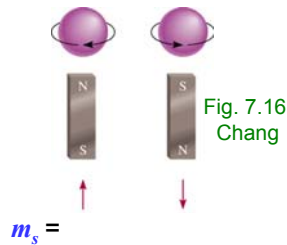
$l = 0$, the electron is in an
 $l = 1$,
 $l = 2$,
 $l = 3$,

m_l : Magnetic quantum number, tells us the orientation of the orbital in space.
The values of m_l depend on the angular momentum quantum number, l .

$(2l + 1)$ possibilities for each l .

if $l = 1$ (p orbital), $m_l =$
 if $l = 2$ (d orbital), $m_l =$

m_s : Electron spin quantum number, tells us the orientation of the spin of the electron. The values of $m_s =$



Schrodinger Wave Equation

$$\Psi = fn(n, l, m_l, m_s)$$

Existence (and energy) of electron in atom is described by its **unique** wave function Ψ . They give the electron a unique 'address'.

Pauli exclusion principle –

Shell – electrons with the same value of

Sub-shell – electrons with the same values of **l** and **m_l**

Orbital – electrons with the same values of **l** and **m_l** are in the same orbital. Each orbital has two electron states: $m_s = \pm \frac{1}{2}$ and therefore can hold two electrons.

These orbitals have different 'shapes', 'size', 'direction' and energy.

Shapes, Size and Direction of Atomic Orbitals (Chang, 7.6 and 7.7):

An electron is more likely to be in some regions of space than others.

Strictly speaking, an orbital does not have a well defined shape as the wave function character of the electron in an orbital stretches to infinity.

We present the 'shape' of an orbital as a probability distribution which is dependent upon

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s Orbitals –

The larger the value of **n**, then the larger the orbital extends into space. s orbitals are spherical.

Where 90% of the e⁻ density is found for the 1s orbital

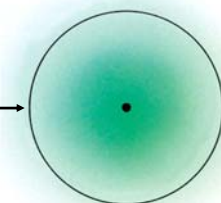


Fig. 7.18, Chang

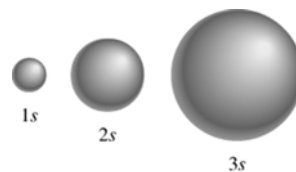
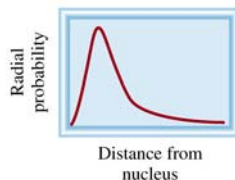
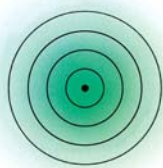


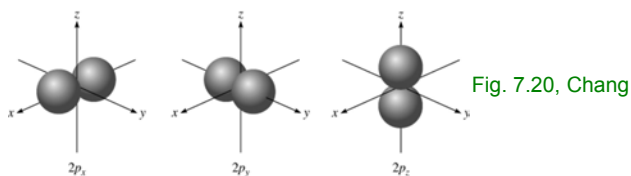
Fig. 7.19, Chang



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p Orbitals –

Each p sub-shell consists of 3 orbitals:



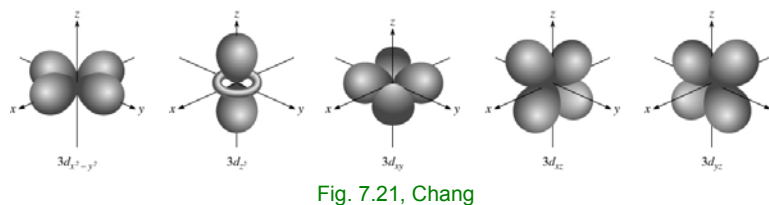
The boundary surface diagrams of the $2p$ orbitals (above) are identical in shape and energy but their orientations (x , y , z) are different.

The 'dumb-bell' shape and orientation are the same in $3p$, $4p$, $5p$... orbitals, although these have larger size.

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d Orbitals –

Each d sub-shell consists of 5 orbitals:



The boundary surface diagrams of the $3d$ orbitals (above) are identical in

The shape and orientation are the same in $4d$, $5d$,... orbitals, although these have larger size.

Orbitals with higher energies than d -orbitals are labelled f , g , and so on are more complicated... we are not concerned with them here.

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TABLE 7.2 Relation Between Quantum Numbers and Atomic Orbitals

n	ℓ	m_ℓ	Number of Orbitals	Atomic Orbital Designations
1	0	0	1	1s
2	0	0	1	2s
	1	-1, 0, 1	3	2p _x , 2p _y , 2p _z
3	0	0	1	3s
	1	-1, 0, 1	3	3p _x , 3p _y , 3p _z
	2	-2, -1, 0, 1, 2	5	3d _{xy} , 3d _{yz} , 3d _{xz} , 3d _{x²-y²} , 3d _{z²}
⋮	⋮	⋮	⋮	⋮
⋮	⋮	⋮	⋮	⋮
⋮	⋮	⋮	⋮	⋮



How many electrons can an orbital hold?



How many 2p orbitals are there in an atom?



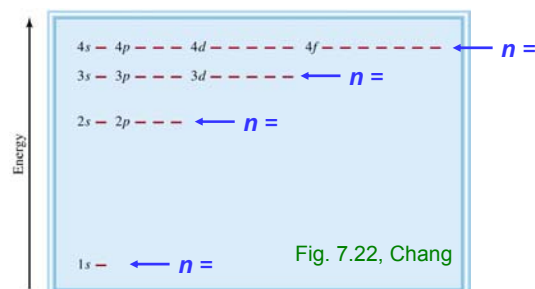
How many electrons can be placed in the 3d sub-shell?

The Order of Orbital Energies (Chang, Chapter 7.7):

Every electron in an orbital has a definite energy (**quantization**). The order of orbital energies changes with the number of electrons.

Case 1: H, a single electron:

The energy of an electron in a hydrogen atom is determined solely by its principal quantum number, n . Thus, the energies of hydrogen orbitals increase as follows:

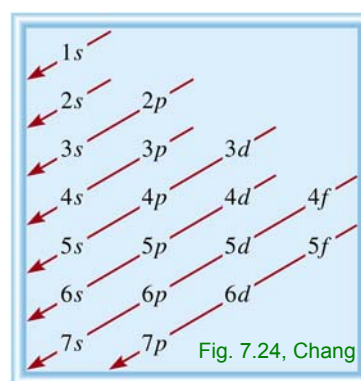
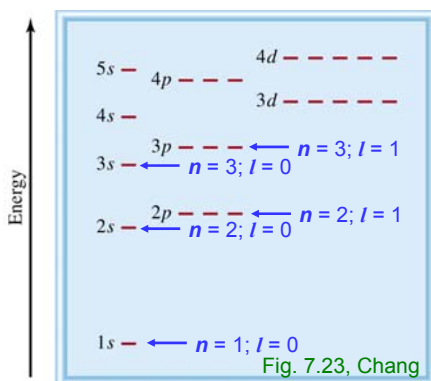


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Case 2: Atoms with many electrons: He, Li, Be, B, C, ...

The energy of an electron in a multi-electron atom is determined by its principal quantum number, n , and its angular momentum quantum number, l .

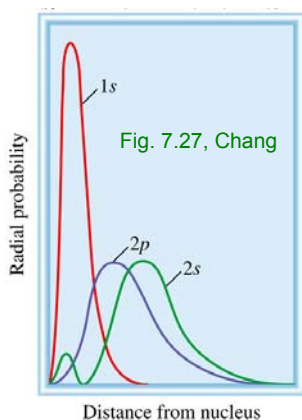
Thus, the lower n is, the lower the energy, but only orbitals in the **same sub-shell** are degenerate:



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Why does the orbital energy change from the 1 electron atom to the many electron atoms?



As a result the energy of the $2p$ is higher than the energy of the $2s$.

In fact, s -orbitals are able to penetrate further toward the nucleus than p -orbitals and therefore feel greater attraction to the nucleus.

Penetrating power: $s > p > d > f > \dots$

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Rules for Assigning Electrons to Atomic Orbitals: Electron Configuration (Chang, Chapter 7.8):

How are the electrons in an atom distributed amongst the various atomic orbitals? e.g. the element, atomic N has 7 electrons, where do they go?

Guiding Principles: 1) **The Pauli Exclusion Principle:**

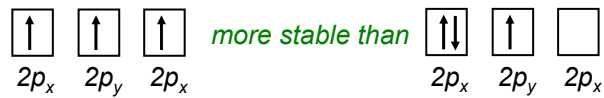
2) The Aufbau ('building up') Principle:

Electrons are added one by one to the atomic orbitals, starting with the lowest energy orbitals. e.g. N ($1s$) ($2s$) ($2p_x$) ($2p_y$) ($2p_z$).

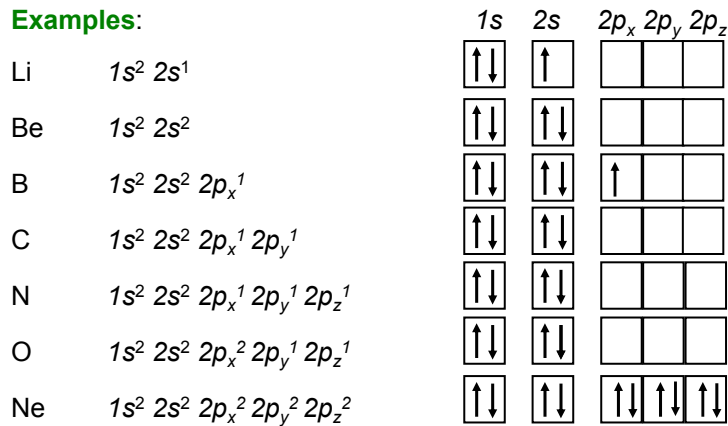
Because each atomic orbital can only have a maximum of 2 electrons, we can represent the electronic configuration as:

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3) Hund's Rule:

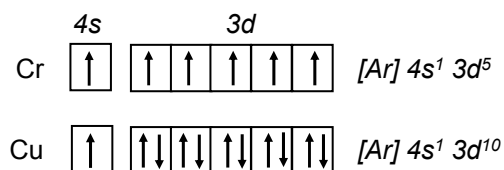


Examples:



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Note:



It is the electronic configuration of the atom that explains a lot of the chemistry, the periodic table and chemical bonding.

Molecules and Bonding (Chang, Chapter 9):

Molecules are made by atoms bonding together.

The two most common types of bond are: **i)** **ii)**

A knowledge of the electron configuration helps understand bonding and chemical bonds.

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Valence electrons: When atoms combine to form molecules, generally only the *outer or valence electrons* are affected.

We use Lewis dot symbols to represent valence electrons:

e.g.

Fig. 9.1, Chang

1 1A	2 2A																			18 8A						
•H																					•He•					
•Li	•Be•																				•B•	•C•	•N•	•O•	•F•	•Ne•
•Na	•Mg•	3 3B	4 4B	5 5B	6 6B	7 7B	8 8B	9 8B	10 8B	11 1B	12 2B									•Al•	•Si•	•P•	•S•	•Cl•	•Ar•	
•K	•Ca•																				•Ga•	•Ge•	•As•	•Se•	•Br•	•Kr•
•Rb	•Sr•																				•In•	•Sn•	•Sb•	•Te•	•I•	•Xe•
•Cs	•Ba•																				•Tl•	•Pb•	•Bi•	•Po•	•At•	•Rn•
•Fr	•Ra•																									

You will need to know the valence electrons of all the elements from the first 2 rows of the periodic table.

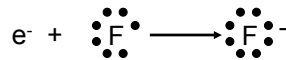
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Ionic Bonds (Chang, Chapter 9.2):

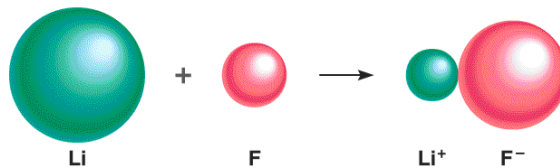
One atom loses an electron, donating it to the other atom to form a bond

Lithium has a low ionization energy, will lose an electron:

Fluorine has a high electron affinity and likes to gain an electron:

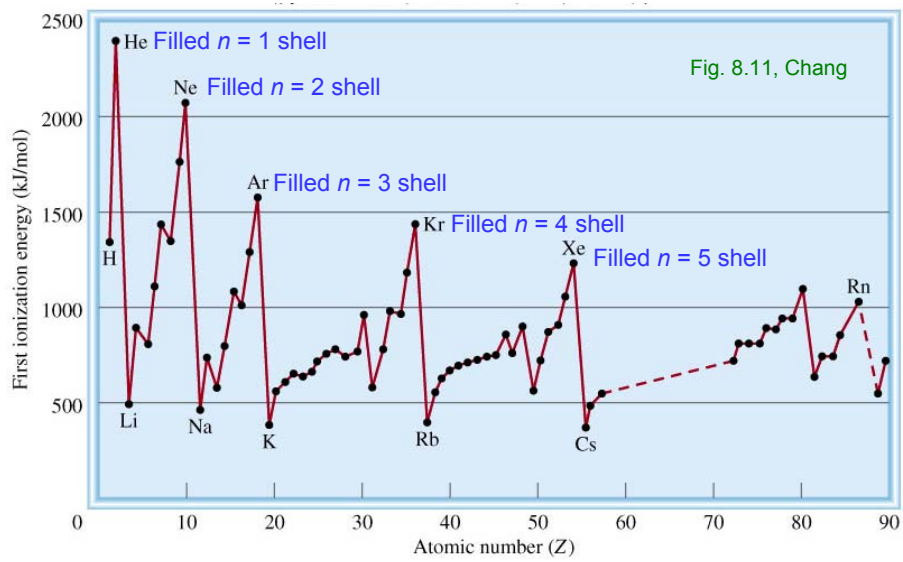


Electrostatic forces hold the compound together: LiF is an



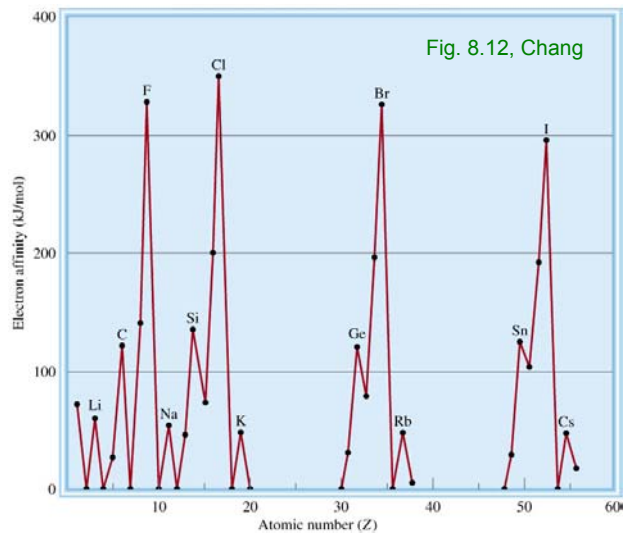
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Variation of the First Ionization Energy with Atomic Number



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Variation of Electron Affinity With Atomic Number (H – Ba)



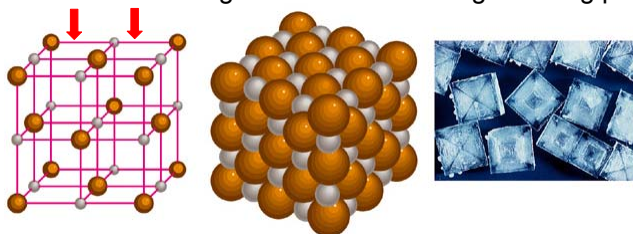
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Nature of Ionic Solids:

NaCl: Extended lattice containing ions rather than distinct molecular entities

Strong electrostatic binding forces.

High melting point.



Covalent Bonds (Chang, Chapter 9.4):

Most elements combine to form covalent bonds.

In covalent bonds,

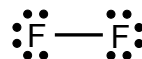
Covalent bonds are not as strong as ionic bonds.

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Why should two atoms share electrons?

Lewis structure of F_2



Lone pairs:

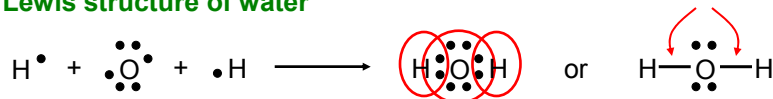
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Guidelines for Sharing Electrons:

H – must attain the electronic configuration of He *i.e.* two electrons.

Other atoms: like to form bonds until they are surrounded by 8 valence electrons (**octet rule**), works particularly well for elements in 2nd period.

Lewis structure of water



Some molecules form multiple bonds: double and triple



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Guidelines for Drawing Lewis Structures:

It is not always easy to decide the Lewis structure (electronic) of molecules.

General Approach:

- 1) Draw skeletal structure of compound showing what atoms are bonded to each other. e.g. *Oxygen*
atoms go round the central atom and *H* and *halogen* atoms go on outside.
- 2)
- 3) Link atoms with single bonds. Add lone pairs to complete an octet for all atoms *except hydrogen*.
- 4)

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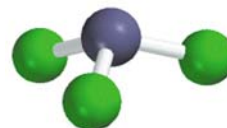
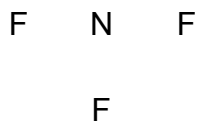
Write the Lewis structure of nitrogen trifluoride (NF₃)

Step 1:

Step 2:

Step 3: Draw single bonds between N and F atoms and complete octets on N and F atoms.

Step 4: Check, are # of e⁻ in structure equal to number of valence e⁻?



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Write the Lewis structure of the carbonate ion (CO₃²⁻)

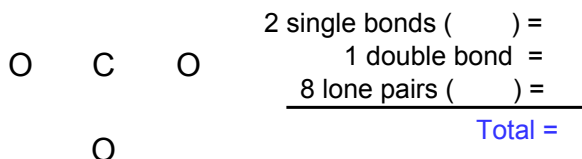
Step 1: C is less electronegative than O, put C in centre

Step 2:

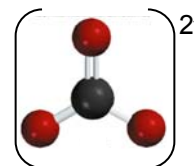
Step 3: Draw single bonds between C and O atoms and complete octet on C and O atoms.

Step 4: Check, are # of e⁻ in structure equal to number of valence e⁻?

Step 5: Too many electrons, form double bond and re-check # of e⁻

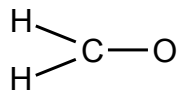
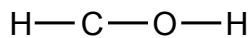


$$\begin{array}{l}
 2 \text{ single bonds (} \quad \quad \text{)} = \\
 1 \text{ double bond } = \\
 8 \text{ lone pairs (} \quad \quad \text{)} = \\
 \hline
 \text{Total} =
 \end{array}$$



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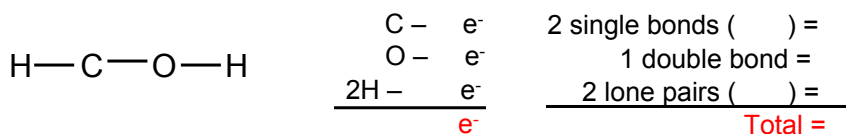
Two possible skeletal structures of formaldehyde (CH₂O)



An atom's **formal charge** is the difference between the number of valence electrons in an isolated atom and the number of electrons assigned to that atom in a Lewis structure.

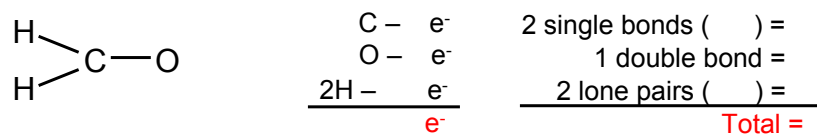
The sum of the formal charges of the atoms in a molecule or ion must equal the charge on the molecule or ion.

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formal charge
on C =

formal charge
on O =



formal charge
on C =

formal charge
on O =

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Formal Charge and Lewis Structures

1. For neutral molecules, a Lewis structure in which there are no formal charges is preferable to one in which formal charges are present.
2. Lewis structures
3. Among Lewis structures having similar distributions of formal charges, the most plausible structure is the one in which negative formal charges are placed on the more electronegative atoms.

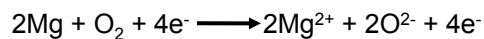
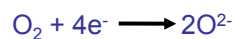
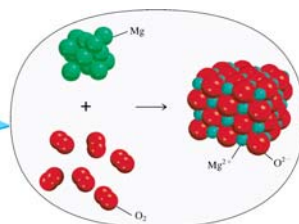


Which is the most likely Lewis structure for CH₂O?

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Oxidation and Reduction Reactions. Electron Transfer Reactions (Chang, Chapter 4.4):

Burning Magnesium Ribbon:



Overall: \longrightarrow

'OILRIG' :

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These reactions are called reduction-oxidation reactions or **REDOX** reactions. The processes of reduction and oxidation always occur together.

Oxidation Numbers (Oxidation States)

An atom's oxidation number (*O.N.*), or oxidation state, is the charge that it would have if the substance were completely ionic – *i.e.* if electrons were transferred completely.

1. The *O.N.* of an atom in an *element* is 0;
2. The *O.N.*
3. The *O.N.* of **O** is usually -2, except for *peroxides* (H_2O_2) where it is -1;
4. The *O.N.* for **H** is +1, except when combined with metals where it is -1;
5. The *O.N.* of **F** is -1;
6. The *O.N.* of other *halogens* is -1 except when combined with **O** or **F**;
7. The sum of the *O.N.*'s
8. The *O.N.* of an atom is usually, but not always an integer.

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What is the oxidation number of **C** in CO_2 ?

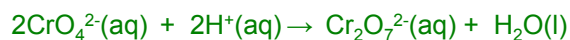


What is the oxidation number of **V** in VO_4^{2-} ?

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Is it a Redox Reaction?



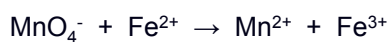
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Balancing Half Reactions

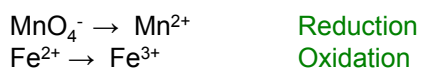
- K** - balance the element;
- O** - balance any loss or gain by adding H_2O to the appropriate side;
- H** - balance by adding H^+ ion to the appropriate side;
- E** - balance the total charge by adding to the appropriate side;
- S** - add the appropriate (phases) to each reaction component in the final, balanced equation (note, electrons do not have a 'state').



Unbalanced redox equation:



Unbalanced half reactions:



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Balance the Reduction Half Reaction:



K:

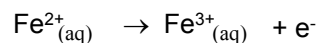
O: Add

H: Add

E: Add

S: Specify the states of the reaction components.

Balance the Oxidation Half Reaction:



K: Fe atoms are already balanced;

O: O atoms already balanced;

H: H atoms are already balanced;

E: Add $3 - 2 = 1 \text{ e}^-$ to the RHS to balance the charges;

S: Specify the states of the reaction components.

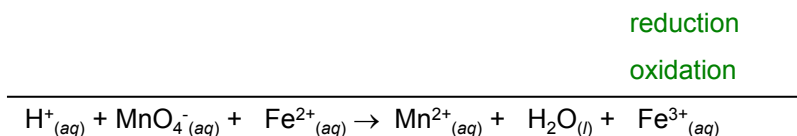
45

Balance the Overall Reaction

Multiply the half reactions to equalise electron counts.

Add the half reactions (cancel out electrons).

Check that elements and charges are balanced.



Multiply the **oxidation reaction** by to equalise the electron counts.

Add the two half reactions (cancel electrons).

Conduct check counts:

	reactants	products
H		
Mn		
O		
Fe		
Charge	$8 - 1 + (5 \times 2) = 17$	$2 + (5 \times 3) = 17$

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Redox Titrations

A titrant (of known concentration) is added, using a burette, to a known volume of an analyte (of unknown concentration).

At the equivalence point, the amount of added titrant is proportional to the amount of analyte.



1. Determine the balanced **redox equations**;
2. Determine the amount of titrant added at the equivalence point;
3. From the stoichiometry of the balanced equations and the amount of the titrant added at the equivalence point, determine the amount of analyte.

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A **16.42 mL** volume of **0.1327 M** KMnO_4 solution is needed to oxidize **25.00 mL** FeSO_4 solution in acidic medium. What is the concentration (i.e. **M**) of FeSO_4 solution?

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An Introduction to the Gas Laws Chang, Chapter 5

Physical Characteristics of Gases:

Gases assume the volume and shape of their containers.

Gases are the most compressible state of matter.

Gases will mix evenly and completely when confined to the same container.

Gases have much lower densities than liquids and solids.

Fig 5.1, Chang: Elements that exist as gases at 25°C and 1 atmosphere

TABLE 5.1 Some Substances Found as Gases at 1 atm and 25°C

Elements	Compounds
H ₂ (molecular hydrogen)	HF (hydrogen fluoride)
N ₂ (molecular nitrogen)	HCl (hydrogen chloride)
O ₂ (molecular oxygen)	HBr (hydrogen bromide)
O ₃ (ozone)	HI (hydrogen iodide)
F ₂ (molecular fluorine)	CO (carbon monoxide)
Cl ₂ (molecular chlorine)	CO ₂ (carbon dioxide)
He (helium)	NH ₃ (ammonia)
Ne (neon)	NO (nitric oxide)
Ar (argon)	NO ₂ (nitrogen dioxide)
Kr (krypton)	N ₂ O (nitrous oxide)
Xe (xenon)	SO ₂ (sulfur dioxide)
Rn (radon)	H ₂ S (hydrogen sulfide)
	HCN (hydrogen cyanide)*

*The boiling point of HCN is 26°C, but it is close enough to qualify as a gas at ordinary atmospheric conditions.

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Experimental Variable: Pressure

$$\text{Pressure} = \frac{\text{Force}}{\text{Area}}$$

(Force = mass x acceleration)

Units of Pressure

1 Pascal (Pa) = 1 N m⁻²

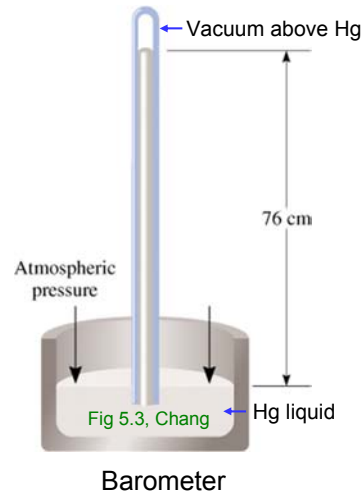
(1 N = 1 J m⁻¹; 1 Pa = 1 J m⁻³)

1 atm = 760 mmHg = 760 torr

1 atm = 101,325 Pa = 101.3 kPa

A bar is defined to be 10⁵ Pa

1 bar = 10⁵ Pa = 100 kPa = 0.98292 atm = 750.06 Torr



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Experimental Variable: *Temperature*

Celsius temperature (θ) is referenced to the freezing point of water and is measured in **degrees Celsius** ($^{\circ}\text{C}$).

Absolute temperature (T) is referenced to the thermodynamic minimum and its unit is the **Kelvin** (K).

Experimental Variable: *Volume*

In chemistry, volume is commonly measured in:

cubic centimetres (cm^3) or millilitres (mL)

cubic decimetres (dm^3) or litres (L)

The SI (*Systeme Internationale*) unit for volume is the cubic metre (m^3).

When volume is used in an equation it should normally be converted to m^3 .

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Experimental Variable: *Amount*

In chemistry, amount (n) is a way of quantifying the number of elementary entities (atoms, molecules, *etc.*)

$$1 \text{ mol} \approx 6.022 \times 10^{23}$$

The actual number of entities (N) is related to the amount by **Avogadro's constant** ($N_{\text{A}} \approx 6.022 \times 10^{23} \text{ mol}^{-1}$):

For a given mass (m) of entities, the amount is given by:

M_{r} is the molar mass, with units of g mol^{-1} .

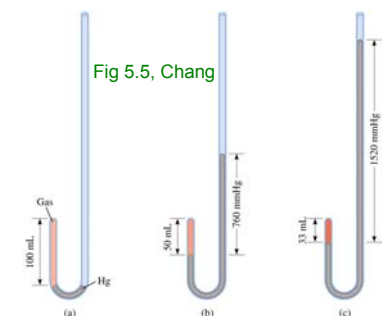
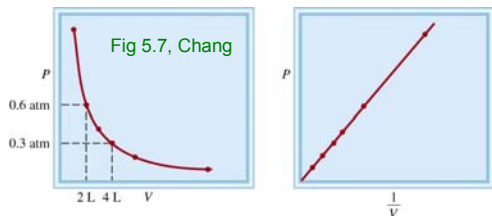
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Boyle's Law: The Pressure-Volume Relationship

Boyle studied the pressure-volume relationship of a *fixed amount* of gas at a *constant temperature*:



Robert Boyle
(1627-1691)
Irish Chemist



Pressure is caused by gas molecules colliding with the container walls:

If the volume is halved:

- there are twice as many molecules per unit volume, so;
-
- the average force they exert is doubled, so;
-

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Boyle's Law:



A sample of chlorine gas occupies a volume of **946 mL** at a pressure of **726 mmHg**. What is the pressure of the gas (in **mmHg**) if the volume is reduced at constant temperature to **154 mL**?

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Gay-Lussac's or Charles's Law: The Temperature-Volume Relationship

Gay-Lussac rationalised Charles's earlier observations on the effects of heating a *fixed amount* of gas at *constant pressure*:



Joseph Gay-Lussac
(1778-1850)



Jacques Charles
(1746-1823)

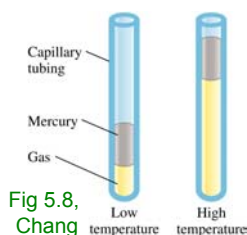
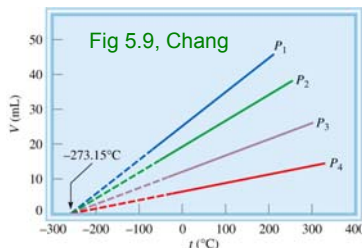


Fig 5.8,
Chang



	Kelvin scale	Celsius scale
Absolute zero:	0 K	-273 °C
Freezing point H ₂ O:	273 K	0 °C
Boiling point H ₂ O:	373 K	100 °C

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Pressure is a measure of kinetic energy density. **If the temperature is doubled:**

- the molecules' average kinetic energy is doubled, so;
- the total kinetic energy of the container is doubled, so;
- the kinetic energy density is doubled, so;
- the pressure is doubled.

Charles's Law:



A sample of carbon monoxide gas occupies **3.20 L** at **125 °C**. At what temperature (in **K**) will the gas occupy a volume of **1.30 L** if the pressure remains constant?

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Avogadro's Law: The Volume-Amount Relationship

'At constant pressure and temperature, the volume of a gas is directly proportional to the number of moles of the gas present'



Amadeo Avogadro
(1776-1856):

At $T = 273\text{ K}$ and $P = 1\text{ atm.}$, 1 mol of gas occupies

The Ideal Gas Equation (IDE):

We can combine all three expressions to form a single master equation for the behaviour of gasses; *the ideal gas equation*:

$$\left. \begin{array}{l} \text{Boyle's law: } V \propto \frac{1}{P} \text{ (at constant } n \text{ and } T) \\ \text{Charles's law: } V \propto T \text{ (at constant } n \text{ and } P) \\ \text{Avogadro's law: } V \propto n \text{ (at constant } P \text{ and } T) \end{array} \right\}$$

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An Ideal Gas

Molecules in an ideal gas do not attract or repel each other.

The total volume of the molecules is negligible in comparison with the volume of the container.

Real gases rarely deviate significantly from ideal behaviour.



What is the volume (in litres) occupied by 49.8 g of HCl at STP?

58



Argon is an inert gas used in lightbulbs to retard the vaporization of the filament. A certain lightbulb containing argon at 1.20 atm and $18 \text{ }^\circ\text{C}$ is heated to $85 \text{ }^\circ\text{C}$ at constant volume. What is the final pressure of argon in the lightbulb (in atm)?



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Calc. the density of $\text{H}_2\text{O}_{(\text{vap})}$ at $100 \text{ }^\circ\text{C}$ and 1 atm pressure. The density of $\text{H}_2\text{O}_{(\text{liq.})}$ at $0 \text{ }^\circ\text{C}$ is 1 g/cm^3 , compare. [$M(\text{H}_2\text{O}) = 18.02 \text{ g/mol}$]

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Dalton's Law of Partial Pressures:

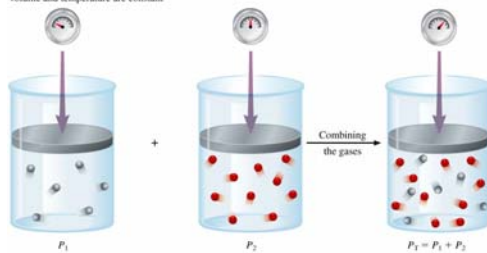
In a mixture of gases **A** and **B**, each gas has a partial pressure dependent on the amount of that gas:



John Dalton
(1766-1844):

Dalton's law says that the *total pressure* is the sum of the partial pressures:

Fig 5.14, Chang
Volume and temperature are constant



Only P_{total} can be determined directly.

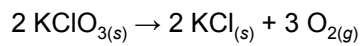
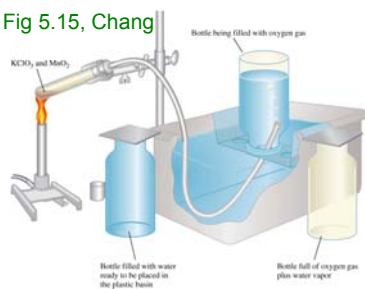
With sufficient information, P_A and P_B can be inferred by using Dalton's law.

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An important application of Dalton's law occurs when a gas is collected over a volatile liquid, like water. In that case:

It can be assumed that the gas is saturated with water. The value of $P_{\text{H}_2\text{O}}$ is equal to the saturated vapour pressure of water at that temperature.

Fig 5.15, Chang



$$P_{\text{total}} = P_{\text{O}_2} + P_{\text{H}_2\text{O}}$$

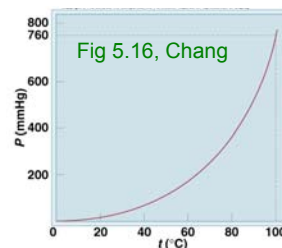


Fig 5.16, Chang

TABLE 5.3
Pressure of Water Vapor at Various Temperatures

Temperature (°C)	Water Vapor Pressure (mmHg)
0	4.58
5	6.54
10	9.21
15	12.79
20	17.54
25	23.76
30	31.82
35	42.18
40	55.32
45	71.88
50	92.51
55	118.04
60	149.38
65	187.54
70	233.7
75	289.1
80	355.1
85	433.6
90	525.76
95	633.90
100	760.00

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O₂ gas generated by the decomposition of KClO_{3(s)} is collected over water (Fig. 5.15). The volume of O₂ collected at 24 °C and a pressure of 762 mmHg is 128 mL. Calculate the mass (in grams) of O₂ gas obtained. The pressure of water vapor at 24 °C is 22.4 mmHg.

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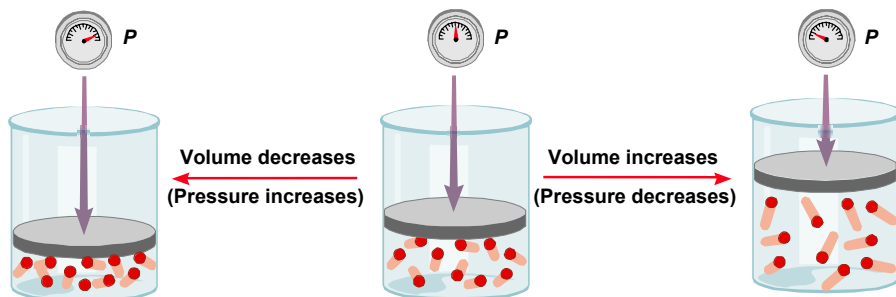
A cylinder of a car engine has a volume of 400 cm³. It takes in air at 1.00 atm and 15 °C and compresses it to 50.0 cm³ (8:1 compression ratio) at 77 °C. What is the final pressure in the cylinder?

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Appendix: Illustration of the Gas Laws, Chang Figure 5.6

Boyle's Law

Increasing or decreasing the volume of a gas at a constant temperature



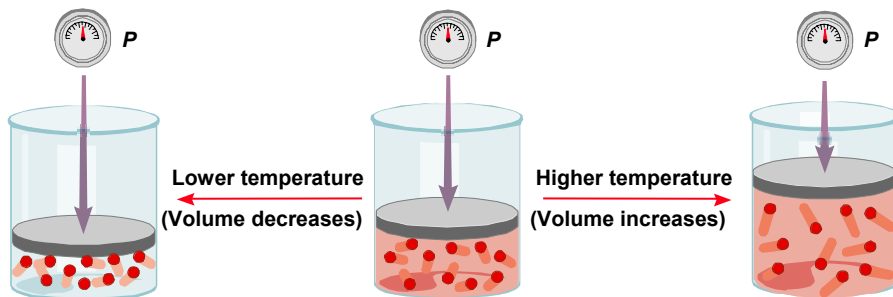
Boyle's Law

$$P = (nRT) \frac{1}{V} \quad (nRT \text{ is constant})$$

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Charles's Law

Heating or cooling a gas at constant pressure



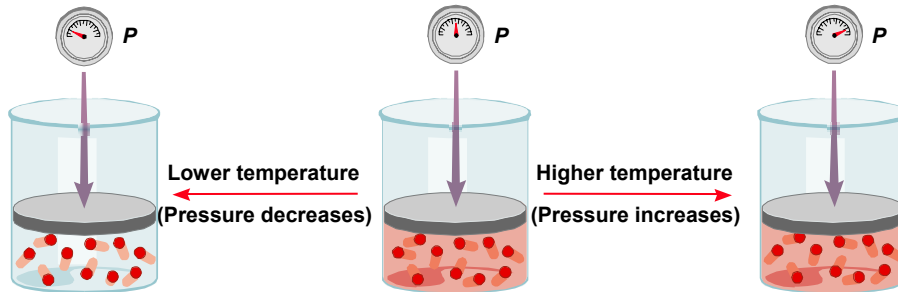
Charles's Law

$$V = \frac{(nR)}{P} T \quad \left[\frac{nR}{P} \text{ is constant} \right]$$

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Charles's Law

Heating or cooling a gas at constant volume



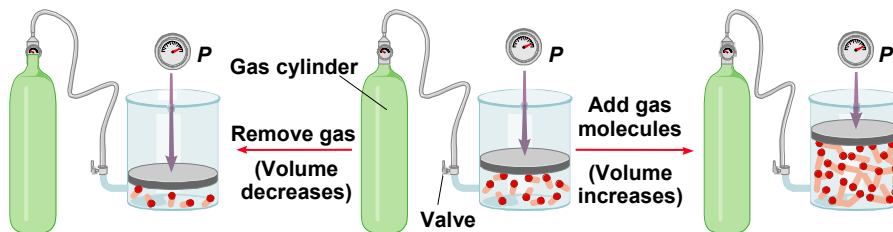
Charles's Law

$$P = \frac{(nR)}{V} T \quad \left[\frac{nR}{V} \text{ is constant} \right]$$

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Avogadro's Law

Dependence of volume on amount of gas at constant temperature and pressure



Avogadro's Law

$$V = \frac{(RT)}{P} n \quad \left[\frac{RT}{P} \text{ is constant} \right]$$

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