# Chapter 11

# Solutions






## Chapter 11 Topics

1. The Composition of solutions
2. The Solution process
3. Factors that affect solubility
4. Measuring concentrations of solutions
5. Quantities for reactions that occur in aqueous solution
6. Colligative properties



## Solubility of Drugs

Figure 11.1



## Solutions are important in the medical field:

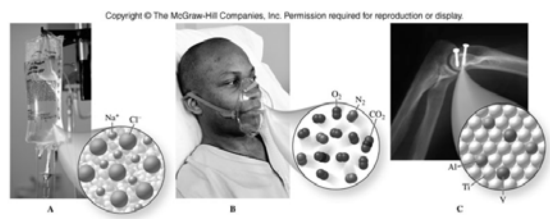



Figure 11.3



## 11.1 The Composition of Solutions

- **Solution** – homogeneous mixture
- **Solute** – substance being dissolved, usually present in the smallest amount
- **Solvent** – substance present in the larger amount
- **Aqueous solution** – a solution where the solvent is water



Electrolytes form ions in solution.
Strong electrolytes dissociate or ionize completely:

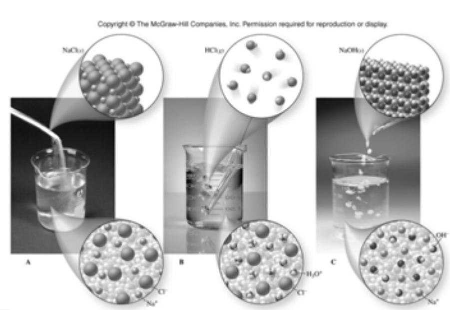



Figure 11.5

$NaCl(s) \rightarrow Na^+(aq) + Cl^-(aq)$ $HCl(g) \rightarrow H^+(aq) + Cl^-(aq)$ $NaOH(s) \rightarrow Na^+(aq) + OH^-(aq)$

# Dissolving NaCl



**$NaCl(s) \rightarrow Na^{+}(aq) + Cl^{-}(aq)$**

**NaCl dissociates completely.**

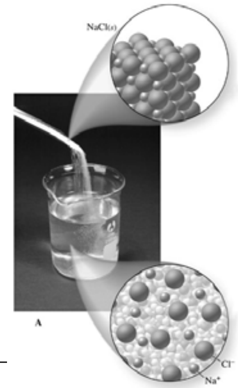



Figure 11.5

# Dissolving HCl



**$HCl(g) \rightarrow H^{+}(aq) + Cl^{-}(aq)$**

**Or**

**$HCl(g) + H_2O(l) \rightarrow H_3O^{+}(aq) + Cl^{-}(aq)$**

**HCl ionizes completely.**


Figure 11.5



# Dissolving NaOH



**$NaOH(s) \rightarrow Na^{+}(aq) + OH^{-}(aq)$**

**NaOH dissociates completely**


Figure 11.5



# $Mg(OH)_2$ is insoluble in water



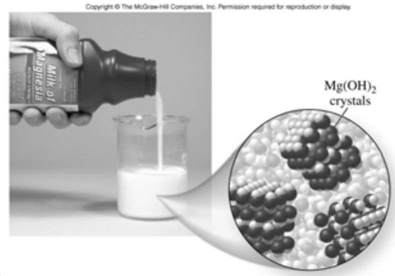



Figure 11.6

Solubility rules are in Table 11.1 and in your lab manual.





**TABLE 11.1 Rules Used to Predict the Solubility of Ionic Salts**

| Ions | Rule |
|---|---|
| $Na^+$, $K^+$, $NH_4^+$ | Most salts of sodium, potassium, and ammonium ions are soluble. |
| $NO_3^-$ | All nitrates are soluble. |
| $SO_4^{2-}$ | Most sulfates are soluble. Exceptions: $BaSO_4$, $SrSO_4$, $PbSO_4$, $CaSO_4$, $Hg_2SO_4$, and $Ag_2SO_4$. |
| $Cl^-$, $Br^-$, $I^-$ | Most chlorides, bromides, and iodides are soluble. Exceptions: AgX, $Hg_2X_2$, $PbX_2$, and $HgI_2$ (X = Cl, Br, or I). |
| $Ag^+$ | Silver salts, except $AgNO_3$, are insoluble. |
| $O^{2-}$, $OH^-$ | Oxides and hydroxides are insoluble. Exceptions: NaOH, KOH, $Ba(OH)_2$, and $Ca(OH)_2$ (somewhat soluble). |
| $S^{2-}$ | Sulfides are insoluble. Exceptions: Salts of $Na^+$, $K^+$, $NH_4^+$, and the alkaline earth metal ions. |
| $CrO_4^{2-}$ | Most chromates are insoluble. Exceptions: Salts of $Na^+$, $K^+$, $NH_4^+$, $Mg^{2+}$, $Ca^{2+}$, $Al^{3+}$, and $Ni^{2+}$. |
| $CO_3^{2-}$, $PO_4^{3-}$, $SO_3^{2-}$, $SiO_3^{2-}$ | Most carbonates, phosphates, sulfites, and silicates are insoluble. Exceptions: Salts of $Na^+$, $K^+$, and $NH_4^+$. |

# Nonelectrolytes



- Most molecular compounds are nonelectrolytes – they retain their molecular structure in aqueous solution.
- $H_2O_2(l) \rightarrow H_2O_2(aq)$


Figure 11.8

## 11.2 The Solution Process

- All substances that dissolve ,whether they're electrolytes or nonelectrolytes, must form new attractive forces with the solvent molecules.


Figure 11.8



## Dissolving NaCl in Water

- What happens when NaCl dissolves in water?

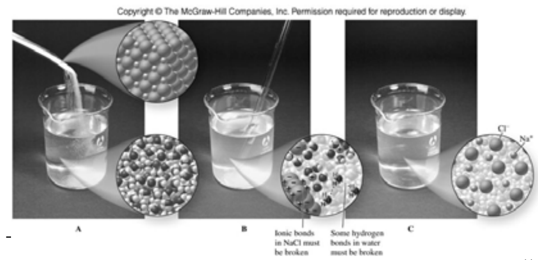



Figure 11.12



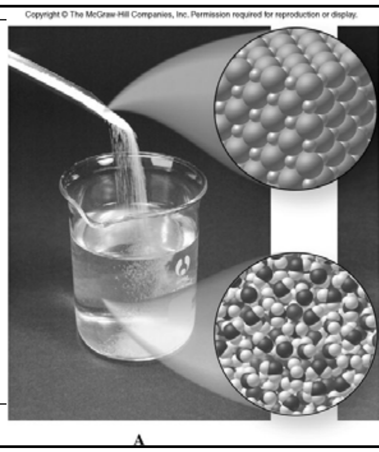



Figure 11.12A



When NaCl dissolves, NaCl ionic bonds break and some H-bonds between water molecules are overcome.

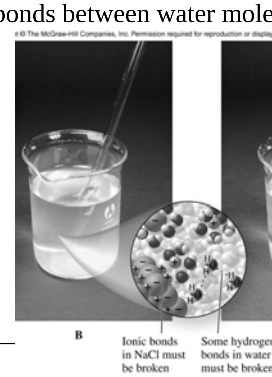



Figure 11.12B



## New forces called ion-dipole forces form.

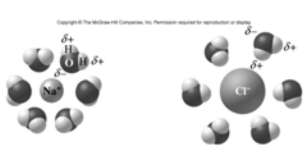



Figure 11.13

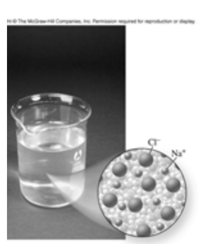



Figure 11.12C



- Some dissolving processes are endothermic (cold pack).
- Some dissolving processes are exothermic (hot pack).


Figure p. 416

## Exothermic Dissolving Process

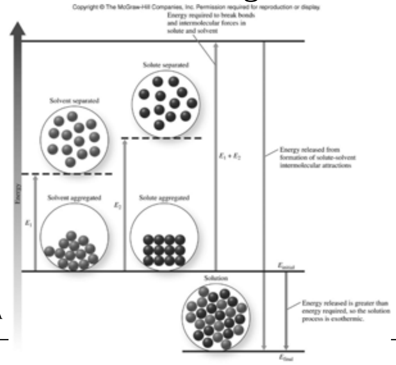



Figure 11.14A



## Endothermic Dissolving Process

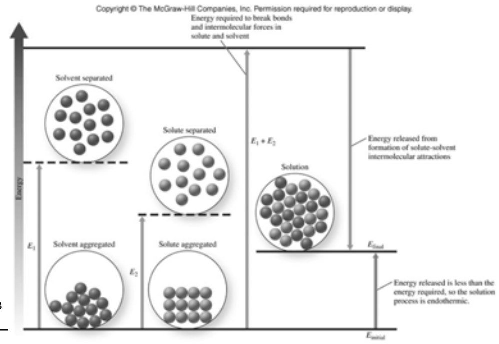



Figure 11.14B



## Driving Forces for Solution Formation

- Decreasing energy is favorable.
- Increasing entropy (disorder) is favorable.

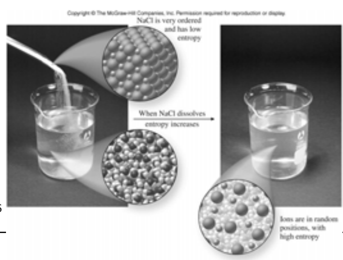



Figure 11.16



## Dissolving of $I_2$ in $CCl_4$

1. What forces are broken?
2. What forces are formed?
3. Does entropy increase or decrease?

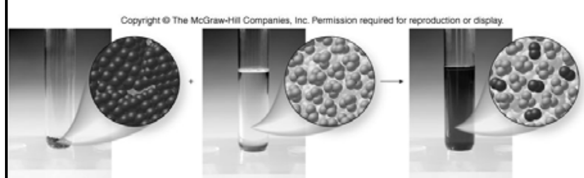



Figure 11.15



## 11.3 Factors that Affect Solubility

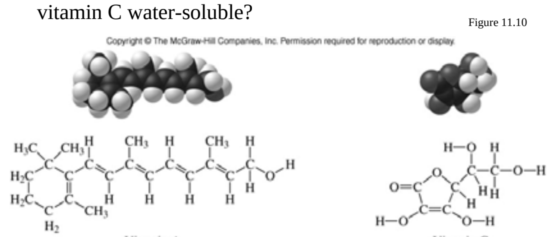
- Structure
- Temperature
- Pressure (gas solutes only)



## Structure and Solubility

- Why do some substances dissolve in one solvent but not another?
- Why is the fat-soluble vitamin A insoluble in water, and vitamin C water-soluble?

Figure 11.10

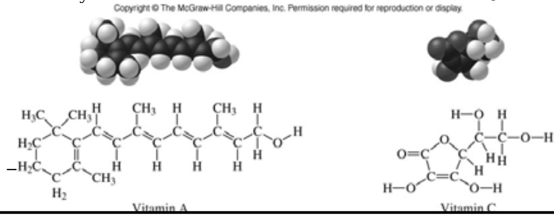


## Structure and Solubility

- A general rule "like-dissolves-like" allows us to predict solubility.
- Vitamin A has a large nonpolar section, so it dissolves more easily in nonpolar solvents such as fat.
- Vitamin C has many –OH groups (very polar) so it dissolves readily in water.

Figure 11.10






## Fat is Mostly Nonpolar



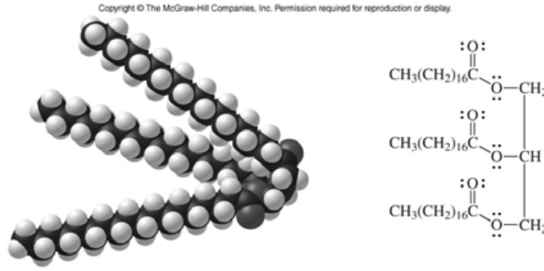



Figure 11.11



## Why are $CCl_4$ and $C_6H_{14}$ soluble in one another?



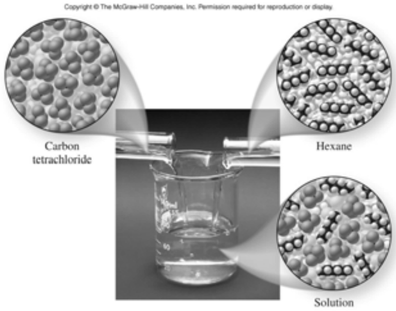



Figure 11.17



## Why do oil (hydrocarbons) and vinegar not mix?



Figure 11.18



## Like-Dissolves-Like

- Which should be more soluble in water?
  - $CH_3OH$ or $I_2$?
- Which should be more soluble in benzene ($C_6H_6$)?
  - $CH_3OH$ or $I_2$?

**TABLE 11.2**

**Like Dissolves Like**

| Solute | Solvent: Polar | Solvent: Nonpolar |
|---|---|---|
| Ionic | Soluble | Insoluble |
| Polar | Soluble | Insoluble |
| Nonpolar | Insoluble | Soluble |

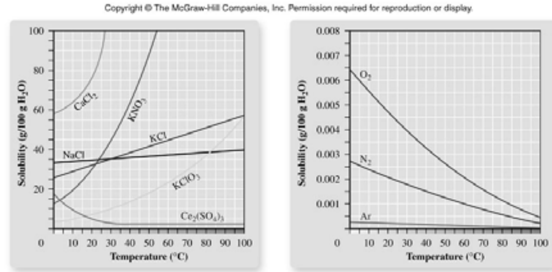
$NH_3$-$H_2O$



## Temperature

Figure 11.19



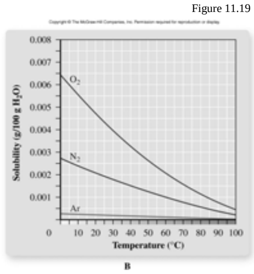



**Solubility of Solids in Water** **Solubility of Gases in Water**

## Gas Solubility vs. Temperature

- The lower $O_2$ solubility at higher temperatures causes fish to die when industries dump hot water into lakes and rivers.

Figure 11.19

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## Gas Pressure

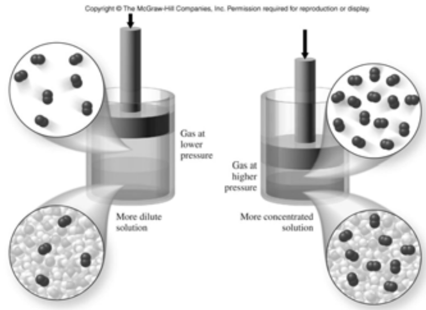


Figure 11.20

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## When <u>pressure</u> is released when the cap is opened, what happens to the solubility of $CO_2$?

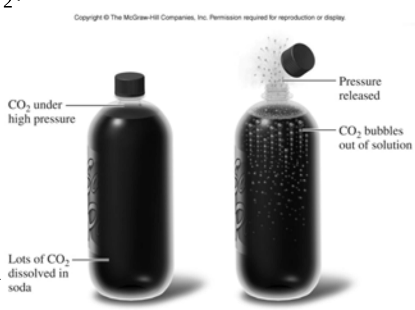


Figure 11.21

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## Gas Solubility vs. Pressure

- What happens to gas solubility in blood when a scuba diver descends to lower depths of the ocean?
- What happens to gas solubility when the scuba diver ascends?
- The “bends” occurs when a scuba diver ascends too quickly. How can the “bends” be cured?

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## 11.4 Measuring Concentrations of Solutions

- Concentration – relative amounts of solute and solvent.
- A **<u>saturated solution</u>** contains the maximum amount of solute in a given amount of solvent.
- An **<u>unsaturated solution</u>** contains less solute than a saturated solution.
- Solubility describes the concentration of a saturated solution. (g solute/100 g solvent)

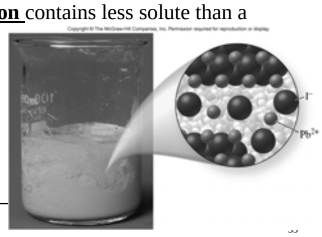


Figure 11.22

## Supersaturated Solution

- Some solutions, when heated and treated carefully, become <u>supersaturated</u>. They hold more solute than a saturated solution. When disturbed by adding a crystal or scratching the container, they <u>precipitate</u> out the excess solute and become saturated.

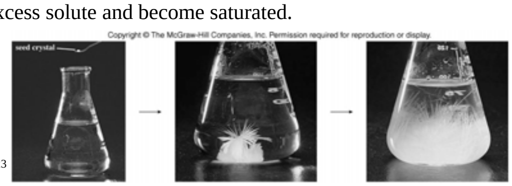


Figure 11.23

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## Solubility  Group Work

- The solubility of NaCl is 38 g/100 grams of water at 25°C.  Describe the resulting solution after 45 g NaCl is added to 150 grams of water.



## Concentrations Expressed Quantitatively

- Percent by Mass
- Percent by Volume
- Parts per Million and Parts per Billion
- Molarity (*M*)
- Molality (*m*)





**TABLE 11.3 Concentration Units**

| Unit | Definition | Unit | Definition |
|---|---|---|---|
| Percent by mass | $\frac{\text{grams of solute}}{\text{grams of solution}} \times 100\%$ | Parts per billion | $\frac{\text{grams of solute}}{\text{grams of solution}} \times 10^9$ |
| Percent by volume | $\frac{\text{volume of solute}}{\text{volume of solution}} \times 100\%$ | Molarity (*M*) | $\frac{\text{moles of solute}}{\text{liters of solution}}$ |
| Mass/volume percent | $\frac{\text{grams of solute}}{\text{volume of solution}} \times 100\%$ | Molality (*m*) | $\frac{\text{moles of solute}}{\text{kilograms of solvent}}$ |
| Parts per million | $\frac{\text{grams of solute}}{\text{grams of solution}} \times 10^6$ | | |



## Percent by Mass of Solute

- Solution concentration is often expressed as the mass percent of solute:

$$\text{Percent Mass Solute} = \frac{\text{mass of solute}}{\text{total mass of solution}} \times 100$$



## Percent by Mass of Solute

- What is the mass percent of NaCl in a solution that is prepared by adding 10.0 g NaCl to 50.0 g water?

$$\text{Percent Mass Solute} = \frac{\text{mass of solute}}{\text{total mass of solution}} \times 100$$



## Molarity (*M*)

- Another common way to express the concentration of a solution is in molarity units:

$$\text{Molarity} = \frac{\text{moles solute}}{\text{liters of solution}}$$

# Colligative Properties

- *A property that does not depend on the identity of a solute in solution*
- **Vary only with the number of solute particles present in a specific quantity of solvent**
- **4 Colligative Properties:**
  - **Osmotic pressure**
  - **Vapor pressure lowering**
  - **Boiling point elevation**
  - **Freezing point depression**



# Osmotic Pressure

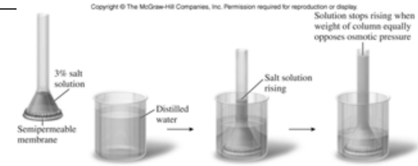



Figure 11.27

- **Osmosis**
  - ***A process in which solvent molecules diffuse through a barrier that does not allow the passage of solute particles***
- **The barrier is called a semipermeable membrane.**
  - ***A membrane that allows the passage of some substances but not others***
- **Osmotic Pressure**
  - ***Pressure that can be exerted on the solution to prevent osmosis***

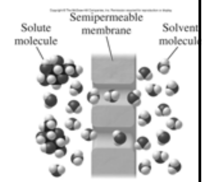



Figure 11.26



# Osmotic Pressure

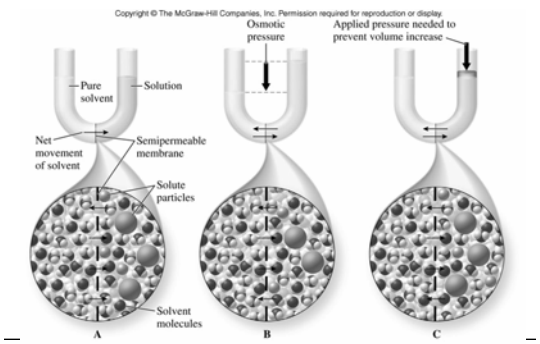
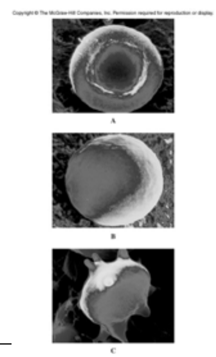



Figure 11.28



# Red Blood Cells

- Isotonic
  - Same concentration inside and out
- Hypotonic
  - Inside is more concentrated than outside
- Hypertonic
  - Outside is more concentrated than inside




cells

Figure 11.30



# Vapor Pressure Lowering

- **Generally, the addition of a nonvolatile solute lowers the vapor pressure of a solution when compared to the pure solvent.**

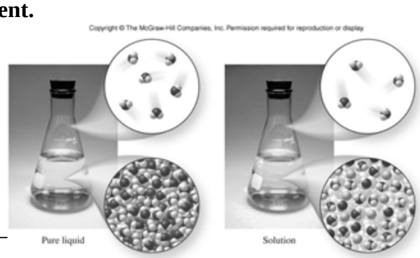



Figure 11.32

# Vapor Pressure Lowering

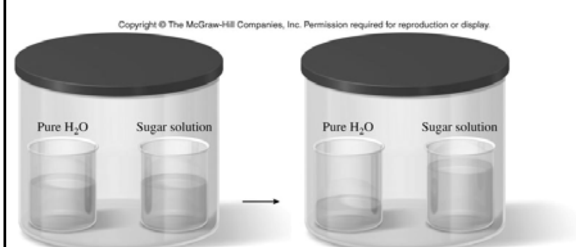



Figure 11.31

## Phase Diagram

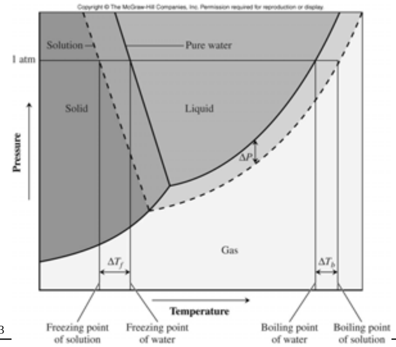



Figure 11.33



## Boiling Point Elevation

- **The boiling point is raised with the addition of solute in comparison to the pure solvent:**

$$\Delta T_b = K_b m$$

**$\Delta T_b$ is the increase in temperature from the pure solvent's boiling point**

**$K_b$ is the boiling point constant, which is characteristic of a particular solvent**

**$m$ is the molality (moles of solute per kg of solution)**



## Freezing Point Depression

- **The freezing point is lowered with the addition of solute in comparison to the pure solvent:**

$$\Delta T_f = K_f m$$

**$\Delta T_f$ is the decrease in temperature from the pure solvent's freezing point**

**$K_f$ is the freezing point constant, which is characteristic of a particular solvent**

**$m$ is the molality (moles of solute per kg of solution)**



## Practice – Freezing Point Depression

- **What is the freezing point for a 1.5 *m* solution of sucrose in water?**



## Colligative Properties and Strong Electrolytes

- **Colligative properties are proportional to the number of particles in solution.**
  - **Strong electrolytes dissociate most of the time into their constituent ions.**
  - **Therefore, the number of particles (in this case ions) increases with the number of ions per formula unit.**
  - **Example:**

$$MgCl_2(s) \rightarrow Mg^{2+}(aq) + 2\,Cl^-(aq)$$

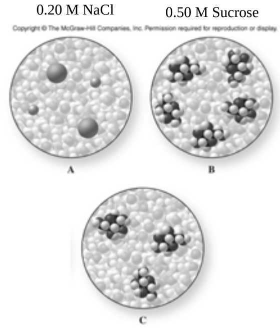


## Group Work

- Which solution should have the highest boiling point?
- Which should have the lowest freezing point?

0.20 M NaCl

0.50 M Sucrose




Question 11.95

0.30 M Sucrose

## Colligative Properties and Strong Electrolytes

$MgCl_2(s) \rightarrow Mg^{2+}(aq) + 2\ Cl^-(aq)$

What is the concentration of ions in a 1.00 *m* $MgCl_2$ solution?

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## Practice – Strong vs. Weak Electrolytes

- Which of the following aqueous solutions is expected to have the lowest freezing point?
  - 0.5 *m* $CH_3CH_2OH$
  - 0.5 *m* $Ca(NO_3)_2$
  - 0.5 *m* KBr

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